Chemistry 30S

Mrs. Kornelsen

Lab Booklet
Lab Safety
&
Outcomes Checklist

Name ______________________________

States of Matter Lab

Objectives:

- Identify the three principle states that water can exist at on Earth.
- Determine the melting and boiling points of water.
- Understand that energy is either released or absorbed during a change of state.

Materials:

Goggles
Ice cubes
400ml beaker
Hot plate
Ring stand with ring
Thermometer
String
Beaker Tongs

Procedure:

1. Fill you beaker just over half full with ice.
2. Record the temperature of your ice (ASAP) in the table below, this will be your time zero temperature.
3. Place your beaker on your hot plate (not turned on yet). Attach a string to your thermometer and hang your thermometer from the ring on the ring stand so that it suspends in your beaker about 2cm above the bottom.
4. Turn on your hot plate.
5. Record the temperature of your ice/water in your beaker **every minute** and record in the table below.
6. Watch your beaker closely. Place asterisks next to the temperatures where you initially se melting (the melting point) and where you initially see boiling (the boiling point).
7. Once the water is boiling continue for another 4 minutes.
8. Make a graph showing temperature vs time. Temperature should be plotted on the vertical y axis and time on the horizontal x axis.

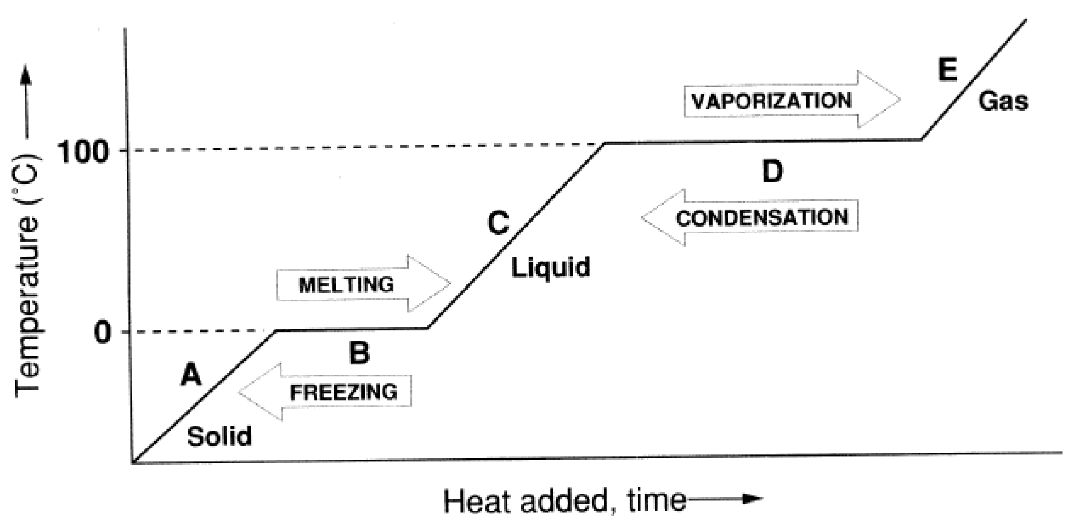
Data Table:

Time (minutes)	Temperature (^oC)	Time (minutes)	Temperature (^oC)
0			

Graphing Directions:

1. Create appropriate scales along your x and y-axis so that the graph takes up the majority of the page.
2. Your scale must be uniform; the scale does not have to be the same for both axis.
3. Connect your dots to show the Heating Curve Trend.
4. Be sure to title your graph and your axis.
5. Your graph should show a similar trend to the one on this page.

The Heating Curve of Water



Concluding Questions

1. What was the melting point? ____________________
2. What was the boiling point? _____________________

3. Describe the states of water as it changes from an ice cube to water vapour. ____

4. Draw and label three diagrams showing the molecules in motion in each of the three states.

5. What is required to change the state of water? ___________________

6. What is happening during the plateau portions of the graph?

7. Why does the temperature stop going up during melting and boiling when we are still adding heat?

8. When ice melts is it absorbing heat or giving heat off? _____________________

Boyle's Law - Simulation Investigation

http://www.chem.iastate.edu/group/Greenbowe/sections/projectfolder/flashfiles/gaslaw/boyles_law.html

Use the Flash simulation of a syringe and a pressure gauge. This simulation uses a moveable syringe with a maximum volume of 30 mL. The pressure is measured in psi (pounds per square inch). Please note that a pressure of 14.7 psi = 1 atm. The starting volume of 30.0 mL is at standard pressure, 1 atm or 14.7 psi. We will examine how pressure changes as the volume in the syringe is decreased.

Procedure:

1. Start collecting data drag the plunger to various volumes. The pressure will read on the gauge and be recorded in the accompanying table.
2. Transfer the results from the simulation to your booklet.
3. Collect the pressure values for at least 6 distinctly different volumes.
4. Do the same for all gases.

Data:

Air		**Hydrogen**		**Helium**		**Oxygen**	
Volume	Pressure	Volume	Pressure	Volume	Pressure	Volume	Pressure

Method of Evaluation:

This assignment will be worth 10 marks. Evaluation will be based on presentation of data, interpretation of the data and accuracy of the answers to questions.

Directions for Hand-In:

1. Draw a graph of pressure vs. volume (pressure on y-axis and volume on the x-axis). Include a title and labels (including units) on the axis. Include four curves on your graph; one for each gas.

Questions

1. How does the pressure change when the volume of the gas was decreased? Is this a direct or inverse relationship?

2. Interpolate on the graph to answer the following:
   a. If the volume is doubled from 10.0 mL to 20.0 mL, what does your data show happens to the pressure? Show the pressure values in your answer.

   b. If the volume is halved from 10.0 mL to 5.0 mL, what does your data show happens to the pressure? Show the pressure values in your answer.

   c. If the volume is tripled from 5.0 mL to 15.0 mL, what does your data show happens to the pressure? Show the pressure values in your answer.

3. One way to determine if a relationship is inverse or direct is to find a proportionality constant, k, from the data. If this relationship is direct, k = P/V. If it is inverse, k = P•V. Based on your answer to Question #2, choose one of these formulas and calculate k for the pressure-volume pairs in your data table (divide or multiply the P and V values for one gas). If you choose correctly, the calculated values should be similar.

Pressure	Volume	K - constant

Charles Law Investigation

Procedure:

1. Grease a 60cc syringe with oil to allow the syringe plunger to move freely.
2. Pick up a thermometer.
3. Travel around the classroom to the different hot water baths. There should be two per bench; each with temperatures of approximately 0^oC, 20^oC, 50^oC, 75^oC, and 95^oC. Try to have only one group and one syringe per water bath (large beaker).
4. Open your syringe and fill it with gas to the 30cc increment. Close your syringe end. Now you have a volume of 30cc of air.
5. Submerge you syringe in the water. You may have to hold it down because we know that gas is less dense than liquids and will float.
6. Take the temperature of the liquid. This will be the temperature of your gas once you let it rest in the liquid.
7. Leave the syringe under water for 3-5 minutes.
8. Take the syringe out and record the new volume and associated temperature in the chart below.

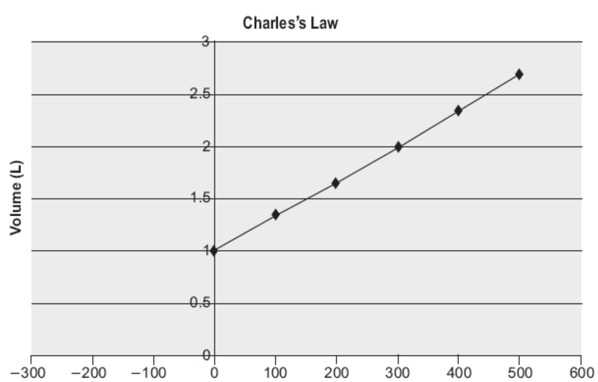
Observations		
Volume (cc)	Temperature (°C)	Temperature (K)

Analysis:

1. Plot a graph of volume vs. temperature (temperature on the x-axis).

Based on your graph, what is the relationship between gas volume and temperature? ____
____

2. The graph below shows Charles Law, but with different volume units. Extend the curve to determine the temperature that would be needed to reduce a volume of gas to zero. This temperature is called absolute zero. ____________________



Why do you think you saw what you did (use the KMT)? __________________________________
____
____
____

The Beanium Lab: Isotopes and Average Atomic Mass

Thanks to Harvey Peltz & RETSD

Introduction:

Within any sample of an element there are millions upon millions of atoms. There are also likely to be various **isotopes** for this element. That is, although all of the atoms may be for the element magnesium with 12 protons and electrons, the sample may have atoms with 12, 13, 14 or more neutrons. Thus, although the atoms all have the **same atomic number** (e.g. 12 protons for magnesium atoms), the atoms may **have variable atomic masses** (e.g., 12 protons + 12 neutrons for an atomic mass of 24 or 12 protons + 13 neutrons for an atomic mass of 25). Because the neutron is such a significant atomic mass unit and can cause such variability in atomic masses, we need to know what amount as a percent of the sample is what type of isotope. The activity that follows will help you to understand this rather complex idea.

Problem:

You are given a sample containing several different types of beans. Each **type** of bean represents an atom for an **isotope** of the element Beanium (Bm). That is, each bean type has the same atomic number; but varies in atomic mass because it contains a variable number of neutrons.

Your job is to determine:

-the relative abundance (%) of each isotope
-the average mass of each bean
-the average atomic mass, in grams, of the element Beanium.

Procedure:

1. Spread out your Beanium sample in front of you. Note that there are three Bm isotopes. Note that the three differ in size and mass. If these were real atoms of three isotopes of the same element, would the three atom types differ in BOTH size and mass? Explain.

2. Look at the beans (atoms) of one bean type (isotope). Are they all the same size and mass? Would they be in a real sample? Explain.

3. Count the **number of each type of bean** (isotope) in your sample and record in the appropriate data table. Remember each bean type is the same Bean element – just different isotopes varying because of different atomic mass.

4. Determine the **average mass** of each type of bean and record this in the chart.

5. Calculate **the average atomic mass, in grams, of your sample of Beanium.** What method will you use?

6. What is the total number of beans in your sample?

Table 1: Beanium Type (Isotope) 1:

Mass of all this type of bean in sample	
Number of beans in this bean type sample	
Average mass of one bean	
Percentage abundance of entire sample with this isotope	

Table 2: Beanium Type (Isotope) 2:

Mass of all this type of bean in sample	
Number of beans of this type sample	
Average mass of one bean	
Percentage abundance of entire sample with this isotope	

Table 3: Beanium Type (Isotope) 3:

Mass of all this type of bean in sample	
Number of beans of this bean type of sample	
Average mass of one bean	
Percentage abundance of entire sample with this isotope	

Calculations and Questions:

1. Percent abundance of each Beanium "isotope" in sample (From table above)

____________________Isotope 1

____________________Isotope 2

____________________Isotope 3

2. Relative (Average) atomic mass of Beanium ____________________

You need to check the % abundance AND average mass of each Beanium isotope in performing this calculation

3. In this sample there was considerable variability in the mass of each bean type. That is all of the same type of bean did not have the same mass. In a sample of an element would all of the atoms for one isotope have the same mass or would there be considerable variability like in each bean type? Why or why not.

4. In your own words explain the meaning of the terms:

   a) isotope:

   b) relative (average) atomic mass of an element:

5. A sample of Magnesium has three isotopes in the following percentages: 12.4% of Mg 24, 46.2 % of Mg 25 and 41.4 % of Mg 26.

   a) how do these three isotopes differ?

   b) Finally, calculate the relative atomic mass of Mg for this sample.

It is important to note that for our studies in chemistry the relative atomic mass provided by **the Periodic Table** is used in determining the atomic mass of individual atoms and the molecular mass of compounds. Although the **atomic mass unit** provides us an approximate atomic and molecular mass, it fails to take into account the presence of isotopes in a sample. **Thus, when calculating the atomic or molar mass of a substance, use the relative atomic masses provided on the Periodic Table rather than adding the number of neutrons and protons in the atoms.**

Conservation of Mass Lab

2 Laboratory Investigation

Conservation of Mass

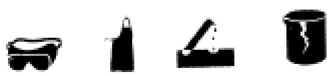
Problem

Is mass conserved in a chemical reaction?

Materials (per group)

Erlenmeyer flask
small test tube
rubber stopper to fit flask
balance
graduated cylinder
copper sulfate solution ($CuSO_4$)
sodium hydroxide solution (NaOH)

Procedure



1. Find the mass of the empty Erlenmeyer flask, test tube, and rubber stopper. Record this combined mass in Data Table 1.
2. Using the graduated cylinder, measure 20 mL of copper sulfate solution. Pour the solution into the Erlenmeyer flask.
3. Measure 10 mL of sodium hydroxide solution and pour this solution into the test tube. CAUTION: Sodium hydroxide is extremely caustic. Follow all appropriate safety precautions.
4. Carefully place the test tube with the sodium hydroxide into the flask containing the copper sulfate, being careful not to mix the two solutions. Stopper the flask and record your observations of the two solutions in Data Table 2.
5. Find the mass of the flask, test tube, rubber stopper, and the two solutions. Record this mass in Data Table 1.
6. Carefully tip the flask so that the solution in the test tube mixes with the solution in the flask. Be careful not to spill the solutions.
7. Record any evidence of a reaction in Data Table 2.
8. Find the mass of the flask, test tube, rubber stopper, and the products of the reaction. Record this mass in Data Table 1.

DATA TABLE 1

Mass of Flask, Test tube, and Stopper (g)	
Empty flask	
Before mixing	
After mixing	

DATA TABLE 2

Appearance of Solutions	
Copper sulfate	
Sodium hydroxide	
Products after reaction	

Observations

1. Find the mass of the solutions (the reactants) by subtracting the mass of the empty flask (step 1) from the mass before mixing (step 5). What is the mass of the two reactants?

2. Find the mass of the products by subtracting the mass of the empty flask (step 1) from the mass after mixing (step 8). What is the mass of the products?

Analysis and Conclusions

1. What changes did you observe to indicate that a reaction took place?

   ____

   ____

2. How did the mass of the reactants (the two solutions) compare with the mass of the products in this reaction?

   ____

   ____

3. Was mass conserved in this reaction? ______________________

4. Was this reaction a chemical or a physical change? Explain.

   ____

   ____

   ____

Reaction

Write the equation for the reaction that took place.

Include the states of all reactants and products.

What type of reaction was this?

Equation Writing and Predicting Products

Introduction

Chemists observe what is happening in a chemical reaction and try to describe it in language that is simple and clear. A chemical equation uses formulas and symbols to describe the substances involved in a reaction, the physical state of the substance, and relative proportions. The general form of an equation is:

Reactants → Products

In this investigation you will perform seven reactions and make careful observations of the changes that occur. Using simple tests and your knowledge of chemistry you will determine the identity of the products. With this information, you will write chemical equations to describe the reactions.

Pre-Lab Discussion

Before we get started we need to know how to test for certain products.

1. What constitutes a positive test for each of the following gases?
   a. Oxygen

   b. Hydrogen

   c. Carbon dioxide

   d. Water vapour

   e. Ammonia (NH_3)

2. What is the role of a catalyst in a reaction? Does a catalyst get included in the equation?

Safety

Wear your goggles at all times during the investigation. Avoid looking at the burning magnesium. The bright light could seriously damage your eyes. Tie back loose hair and clothing when working with a flame. Clean up all spills immediately. Lead and copper compounds should be collected in designated waste containers. The glass tubing breaks easily. Exercise caution when working with it.

General

This lab consists of seven mini experiments designed to demonstrate types of reactions. For each of the reactions, record data in the Observation Table. Observations should include appearance of the reactants; evidence that a chemical reaction has taken place; the results of tests performed or any gases produced; and the appearance of the products.

Procedure

Reaction 1: Hold a piece of Magnesium metal ribbon with tongs, carefully place it in a burner flame. Hold the burning magnesium over a watch glass to catch the debris. CAUTION: *The end of the tongs will be hot. Do not look directly at the magnesium while it burns.* When the magnesium is finished burning, place the remains on the watch glass.

Reaction 2: Place a test tube in the test-tube rack. Have a second test tube ready in a test-tube holder. Add 5-10 ml of 3.0 M HCl to the first test tube. Drop a 2-cm piece of magnesium ribbon into the acid. CAUTION: *Hydrochloric acid is corrosive. Avoid spills and splashes. If you do spill acid, immediately rinse the area with plenty of cold water and report the spill to your teacher.*
Invert the second test tube over the mouth of the first test tube. When the reaction appears to have ended, light a wood splint and quickly test the collected gas for flammability by holding the burning wood splint near the mouth of the second test tube. CAUTION: *The gas in the tube will make a popping sound. Do not be startled.*

Reaction 3: Grasp a small piece of copper foil with your tongs and heat it in the burner flame until it is red hot. Remove it from the flame and allow it to cool. Scratch the surface of the metal with a sharp object. Hint when we burn something we are in fact adding oxygen to it.

Reaction 4: Place one spatula of ammonium carbonate ($(NH_4)_2CO_3$) into a test tube. Place the test tube in a test tube clamp attached to a support stand. Heat the solid gently for a few seconds at a time. Continue heating in this manner for 1 minute. Carefully waft the air towards your nose to detect any odor. *CAUTION: When heating the test tube, point the open end away from yourself and anyone nearby. The gas coming from the tube is a skin and respiratory irritant, so avoid inhaling it deeply.*
Continuing to heat the solid, place a burning splint at the mouth of the test tube. Finally, as heating continues, place a piece of blue cobalt paper just inside the mouth of the test tube.
When you go to write this equation think about how many products you detected. This is a tricky decomposition reaction where the products are not elements.

Reaction 5: Place 10 – 15 ml of hydrogen peroxide, H_2O_2, into a test tube. Have a wooden splint and matches ready. Add a very small amount (about the tip of a spatula) of manganese(IV) oxide, MnO_2 (catalyst) to the hydrogen peroxide. As the reaction occurs, light the splint and allow it to burn freely for 5 seconds. Blow the flame out and place the glowing splint halfway into the test tube.

Reaction 6: Place a drop of potassium iodide solution, KI, in one well of a well plate. Add a drop of lead nitrate solution, $Pb(NO_3)_2$. CAUTION: *Lead compounds are poisonous. If contact occurs, immediately wash with plenty of water and inform your teacher. Clean up all spills immediately.*

Reaction 7: Place a small amount (about one spatula) of copper carbonate, $CuCO_3$, in a test tube. Place the stopper assembly into the test tube. Prepare another test tube with about 10-15 ml limewater. Place the tube with the copper carbonate in a test tube clamp attached to a stand. Heat the test tube and place the end of the rubber tubing all the way to the bottom of the limewater in the other test tube. CAUTION: *Before removing the heat source you must pull the rubber tubing out of the lime water.*

Observations:

Reaction	Observations
1. Burning Mg	
2. Mg and HCl	
3. Heating Cu	
4.Heating $NH_4)_2CO_3$	
5. H_2O_2 and MnO_2	
6. KI and $Pb(NO_3)_2$	
7. Heating $CuCO_3$	

Analysis:

1. Write a balanced equation for each of the reactions performed. Include the physical state of each substance.

Reaction 1

Reaction 2

Reaction 3

Reaction 4

Reaction 5

Reaction 6

Reaction 7

2. Classify each of the reactions as synthesis, decomposition, single replacement, double replacement.

Reaction 1 _______________ Reaction 2 ______________ Reaction 3 _________________

Reaction 4 ________________ Reaction 5 _______________ Reaction 6 ________________

Reaction 7 _________________

3. When potassium bromate ($KBrO_3$) is heated, it decomposes into potassium bromide (KBr) and a gas that supports the combustion of a glowing splint. Write a balanced chemical equation for the reaction.

Lab 5 Stoichiometry of a Reaction

In this experiment you will study the reaction that occurs when solid copper wire is placed in a solution of silver nitrate. You will look at the amounts of reactants used and products formed, and compare those results to what is predicted by the balanced equation.

Purpose: To determine the mass of reactants used and products formed in a chemical reaction and to compare those results with the balanced equation.

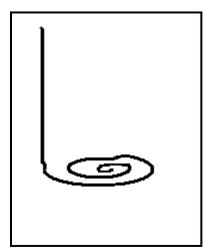
Materials:	30 cm piece of copper wire	250 ml beaker
	Vial containing silver nitrate	stir rod
	Distilled water in squeeze bottle	goggles
	100 ml beaker	balance

Safety: Silver nitrate, as a solid and in solution, can harm your eyes and skin, and damage your clothing. Wear goggles at all times. Rinse skin promptly and thoroughly. Be careful not to get any on your clothing.

Procedure:

Part I

1. Put a piece of tape on the 100 ml beaker and write your names on it. Then mass the beaker to the nearest 0.01 g.
2. Mass the vial containing the silver nitrate to the nearest 0.01 g.
3. Put about 20 ml of distilled water in the beaker. Add the silver nitrate to the water. Do not wash the vial out. Stir the solution gently with the rod to dissolve the crystals. Rinse the solution off the stirring rod into the beaker with a small amount of distilled water from the squeeze bottle.
4. Mass the empty vial and cap to the nearest 0.01 g.
5. Coil the copper wire so that it looks like the drawing . Mass the wire to the nearest 0.01 g.



6. Place the coiled wire in the silver nitrate solution. Do not shake or disturb the contents of the beaker as the reaction proceeds for the next 30 minutes.
7. Observe the reaction and record your observations in point form.

Part II

8. After 30 min. Record any further observations.
9. Shake the newly formed crystals from the copper wire. Leave the wire in the solution for a few minutes to make sure the reaction has finished.
10. Rinse the copper wire off into the beaker with distilled water from the squeeze bottle.
11. Wave the wire around for a minute to help it dry. Set it aside to finish drying.

12. After the crystals have settled to the bottom of the beaker. Decant the solution. (Decant means to pour off the liquid, leaving the solid behind.)
13. Wash the crystals by squirting some distilled water into the beaker and decanting again. Do this several times. There should be no blue- grren color remaining in the crystals.

14. Put the beaker of crystals in the warming oven to dry overnight.

15. Mass the dry copper wire.

Part III

16. The next class, mass the beaker containing the dried crystals.

Observations:

What I saw during the reaction.

____
____
____
____

Data: for lab 5

	Mass in grams
100 ml beaker (dry)	
Vial, cap and silver nitrate	
Empty vial plus cap	
Copper wire before reaction	
Copper wire after reaction	
Beaker plus dried crystals after reaction	

Questions:

1. What type of reaction occurred?

2. Write the balanced equation for the reaction.

3. What is the blue solution formed?

4. Calculate the theoretical yield (what we expect to make) of Ag. Make sure to determine what your limiting reactant is first

5. How much silver did you actually obtain? ___________________________

6. Calculate your percent yield.

7. Some of the copper did not react, even though the reaction appeared to stop. Why?

8. Determine the mass of copper wire used up in the reaction. What do you notice about this value of Cu? (Hint: check your stoichiometric coefficients)

<u>Solutions Lab: Classifying Different Types of Solutions</u>

Solutions are composed of a **<u>solute</u>** and a **<u>solvent</u>** combined together.

The item in solution that is present in *smaller* amounts is usually classified as the **<u>solute</u>**.
The item in solution that is present in *larger* amounts is usually classified as the **<u>solvent</u>**.

i.e. Consider making lemonade by mixing lemonade crystals into a glass of water. Here the lemonade crystals are the **<u>solute</u>** and the water is the

There are nine different types of solutions each classified based on the **<u>state</u>** of the solute and solvent used to create the solution. The lemonade is a solid-liquid solution because it has a solid solute (lemonade crystals) and a liquid solvent (water).

- In groups of 2-3 perform the following mini experiments.
- **Classify the types of solutions you have made by filling in the attached chart.**
- **Be prepared to show off your solutions.**

Equipment: (pick up from the front of the class)

- 100 ml volumetric flask
- 1 large balloon
- 2 Erlenmeyer flasks
- 2 straws
- 1 wash bottle of distilled H_20
- 1 bottle Universal Indicator
- 1 test tube rack
- 3 test tubes & stoppers
- 250 ml beaker

1. Measure out 4.0g of $K_2Cr_2O_7$ (potassium dichromate) on a weigh scale.
   Transfer your 4.0g of $K_2Cr_2O_7$ into a 100ml volumetric flask.
   Fill the volumetric flask with distilled H_2O to the calibration line.

   Name this solution and place it in the appropriate section of your chart.
   In your chart also list an everyday example of this solution.

2. Blow up a large balloon. **What are the solvent and solute in this solution? Where does this solution fit in your chart?**

   ____
   ____

3. Imagine blowing on a shelf covered in dust. **What type of solution would you have created? Add it to your chart.**

4. Expiration mini lab
   - You need 2 flasks, 2 straws, and universal indicator.
   - Put 100ml of H_20 into each flask add 5 drops of universal indicator to both.
   - Watching the clock - time one group member while he/she blows through a straw into the water of one flask.
   - Note the time it takes to observe any change. **What did you see?**
   - After have another group member do 15 jumping jacks and then blow through a straw into the other flask. **Was there any difference?**
   - **What type of solution is this? List it in your chart.**
   - **What is your solute? Your solvent?**
   - **In the chart list another example of this type of solution. Think hard!!**

   ____
   ____

5. There are three liquid chemicals at the front of the class. In 3 test tubes combine them as listed. Add approximately 5 ml of each liquid. Stopper and shake the liquids together.
   - H_2O + vinegar
   - H_2O + oil
   - vinegar + oil

   Which of the above combinations yielded solutions?

   ____
   ____

6. Boiling mini lab
   - You need 1 250 ml beaker filled ¼ full with H_2O, and a hot plate.
   - Place beaker on hot plate and wait for the water to boil.
   - Once your water has started to boil turn off your hot plate immediately.
   - **What did you see?**
   - **What type of solution is this? List it in your chart.**
   - **What is your solute? Your solvent?**
   - **In the chart list another example of this type of solution. Hint: You might wear this on a date!**

   ____
   ____
   ____

***Solutions using solid as a solvent are difficult to reproduce in the lab. We will discuss examples of this as a class. There will be questions on your unit test directly from this lab.

$Na_2S_2O_3 \cdot 5H_2O$ Saturation Lab

Purpose: To distinguish between and define saturated, unsaturated, and supersaturated solutions.

Materials:

- 3 test tubes
- 10 ml graduated cylinder
- Wash bottle
- Glass marking pencil
- Test tube holders
- $Na_2S_2O_3 \cdot H_2O$ (sodium thiosulphate)
- Glass stir rod
- Thermometer
- Scoopula

Procedure:

1. Using scales weigh 0.1 g of crushed $Na_2S_2O_3 \cdot H_2O$ and place it in your first test tube.
2. Label this test tube A with a pencil.
3. Add 3 ml of distilled water to test tube A.
4. Stir and shake the test tube until all the solid dissolves.
5. Add one more crystal of $Na_2S_2O_3 \cdot H_2O$. Did it dissolve? _______________
6. Record the temperature of your solution. _________________
7. Weigh out 0.5 g of $Na_2S_2O_3 \cdot H_2O$ and place it in your second test tube.
8. Label this test tube B with a pencil.
9. Add 3 ml of distilled water to test tube B.
10. Stir and shake the test tube until all the solid dissolves.
11. Add one more crystal of $Na_2S_2O_3 \cdot H_2O$. Did it dissolve? _______________
12. Record the temperature of your solution. _______________
13. Weigh out 5.0 g of $Na_2S_2O_3 \cdot H_2O$ and place it in your third test tube.
14. Label this test tube C and also mark one person's initials on it so we can identify your test tube from others' in the class.
15. Add 1 ml of distilled water.

16. Stir and shake the test tube to dissolve as much solid as possible.

17. If you are unable to dissolve all of the solid bring your test tube and place it in the class hot water bath.

18. Leave your test tube in the heat until all the $Na_2S_2O_3 \cdot H_2O$ dissolves (approximately 5 minutes).

19. While you are waiting take another look at test tubes A & B; make sure all your observations are recorded in the chart.

20. Then poor test tubes A & B down the sink and return them to the equipment cart.

21. Once all the solute has dissolved in test tube C take your test tube back to your bench; carry it in your test tube rack.

22. Record the temperature of your warm solution. _____________________

23. Leave your solution until it has cooled to room temperature. **While you are waiting fill in the bold statements on the next page of your booklet.**

24. Record the temperature of your cooled solution. ___________________________

25. Add one more crystal of $Na_2S_2O_3 \cdot H_2O$. Did it dissolve? ___________________

26. Record any unusual observations after one more crystal was added.

	Contents	**Temperature (^0C)**	**Observations**	**Saturated? Unsaturated? Supersaturated?**
Test Tube A	0.1 g $Na_2S_2O_3 \cdot H_2O$ 3 ml H_2O			
Test Tube B	0.5 g $Na_2S_2O_3 \cdot H_2O$ 3 ml H_2O			
Test Tube C	5 g $Na_2S_2O_3 \cdot H_2O$ 1 ml H_2O	Heated – Cooled -		

<u>Solubility Curve Lab: How Does Temperature Effect Solubility?</u>

Mass of KNO_3 assigned: _______________________________

Introduction

For all solids there is a limiting amount that will dissolve in a given amount of solvent. Some solids are very soluble in water, while others are nearly completely insoluble. About 200 grams of ordinary sugar will dissolve in 100 grams of water at 25°C, but only about 2×10^{-4} g AgCl will dissolve under those conditions.

The solubility of a solid in a given solvent depends on the temperature of the solution. Usually the solubility of a solid in water increases with increasing temperature. A graph, which plots the solubility as a function of temperature, is called the solubility curve of the substance. Given such a graph, the solubility of the solid at any temperature may be determined.

Purpose

This lab is intended to help you understand solubility curves by constructing a solubility curve for KNO_3 ourselves. Each lab group must determine the temperature at which their assigned mass of KNO_3 is soluble in 20ml H_2O. Then we will combine the class results to construct our solubility curve (g KNO_3/100ml H_2O).

Precautions

Handle hot equipment and solutions cautiously.

The thermometer bulb is fragile; handle gently.

Goggles must be worn at ALL times.

Materials

- 1 large test tube
- 1 test tube clamp
- ring stand
- thermometer
- Distilled water
- 10 ml graduated cylinder
- Stirring rod
- Hot plate & 400 ml water bath (shared)

Procedure

1. Identify where your hot water bath is. There should only be 2 groups per hot water bath.

2. Weigh out your assigned amount of potassium nitrate, KNO_3. Your sample will range from 8.00 to 35.00 grams. Transfer the solid to a large test tube and add 20 mL of distilled water.

3. Stir the H_2O + KNO_3 mixture with a stirring rod to dissolve as much potassium nitrate as possible. Carefully insert a thermometer. Record the initial temperature.

4. Clamp the test tube to your ring stand and take your ring stand to the hot water bath. Lower the test tube so that your solution sits in the heated water.

5. Use a stirring rod to stir the mixture gently during heating until all of the potassium nitrate has dissolved. The temperature may vary from 25 °C to 85 °C depending on the amount of potassium nitrate in the sample.

6. When all of the potassium nitrate has dissolved, loosen the clamp from the support stand and raise the test tube out of the water bath. Re-clamp the tube higher on your ring stand and return to your lab bench.

7. Using the stirring rod, stir the solution gently and observe it as it cools.

8. As soon as crystallization begins (you see any precipitate in your solution), record the temperature of the solution. It is easiest to see crystals by looking at the bottom of the test tube where they will collect as soon as they begin to form.

Observations:

9. After the crystallization temperature has been recorded, put the test tube back into the water bath and warm the solution until all of the crystals have re-dissolved.

10. Repeat the cooling procedure to check the crystallization temperature again. The two readings should be within 1 degree of each other. If the two readings do not agree within 2 degrees, re-warm the solution in the water bath, cool it, and continue until a stable temperature is established.

11. Record your sample mass and solubility temperature on the white board in grams of potassium nitrate dissolved per 100 grams of water.

12. When the data from all of the student samples are recorded on the board, copy them to draw a solubility curve for grams of potassium nitrate dissolved per 100 grams of water.

Data Table

g KNO_3/20g H_2O	g KNO_3/100g H_2O	Temperature (^oC)
8.0 g		
10.0 g		
12.0 g		
14.0 g		
17.0 g		
20.0 g		
24.0 g		
27.0 g		
30.0 g		
35.0 g		

Analysis

Construct the solubility curve for KNO_3 .

- Use the class data recorded in your table.
- Plot solubility in grams per 100 grams of water on the y-axis and temperature on the x-axis.
- Connect the points with a smooth curve.
- **Extend your curve to include solubility from 10 °C to 100°C.**
- Attach your graph to your lab report.

From **your** solubility curve predict:

1. solubility of KNO_3 at 100 °C
2. solubility of KNO_3 at 10 °C
3. solubility of KNO_3 at 25 °C
4. solubility of KNO_3 at 75 °C
5. amount of KNO_3 in its saturated solution at 30°C
6. amount of KNO_3 in its saturated solution at 50°C
7. amount of KNO_3 in its saturated solution at 80°C
8. amount of KNO_3 that will crystallize out when its saturated solution at 80 °C is cooled to 50 °C

Discussion

In your own words state in how the solubility of KNO_3 varies with temperature.

Title:

The Effects of Salt & Antifreeze on the Melting and Boiling Points of Water

In this investigation you will study the effects of salt and antifreeze on the melting and boiling points of water.

Materials:

- 2 x 250 ml beaker
- Thermometer
- Salt (7 packages)
- Ice
- 100 ml Antifreeze (ethylene glycol)
- Graduated cylinder
- Stirring rod

Procedure A:

1. Fill one of your beakers with 150 ml of water.
2. Place it on a hot plate and leave it to heat up.
3. Move on to procedures B & C (step 9 below) but keep an eye on your water. Come back to it when your water has begun to boil.
4. Record the temperature of your boiling water in table 1.
5. Add 10 ml of antifreeze to the water. Stir with your stir rod while you continue heating the mixture. Record the new temperature.
6. Repeat step 5 five times until a total of 50 ml of antifreeze has been added.
7. Record the temperature after each addition.
8. Empty your beaker and rinse it.

Boiling water	**Temperature (^oC)**
Water	
Water + 10 ml ethylene glycol	
Water + 20 ml ethylene glycol	
Water + 30 ml ethylene glycol	
Water + 40 ml ethylene glycol	
Water + 50 ml ethylene glycol	

Procedure B:

9. Fill a beaker half full with ice.
10. Add 10 ml of water.
11. Take the temperature of the ice water while stirring with a stir rod.
12. Once a stable temperature has been reached (the thermometer stops moving) record it as the initial temperature of your ice water.
13. Stir in one package of salt. Record the temperature in your chart.
14. Continue adding one package of salt at a time, with continuous stirring and temperature readings, until you have added 7 packages.
15. Record your results in the table.
16. Empty your beaker and rinse it.
17. Check on your boiling water.

	Temperature (°C)
Initial	
1 salt package	
2 salt packages	
3 salt packages	
4 salt packages	
5 salt packages	
6 salt packages	
7 salt packages	
Temperature Range	

Discussion Questions

1. What effect does salt have on the melting point of ice? Why?

2. Why might salt be poured on ice-covered roads?

3. Why might antifreeze be added to the water in radiators of cars in Winnipeg?

4. What effect does antifreeze have on the boiling point of water?

Concentration and Dilution Lab

Objectives

1. To mix a solution and determine its concentration.
2. To perform four dilutions with your prepared solution.

Materials

- $CuCl_2 \cdot 2H_2O$
- Scoopula
- 200 mL volumetric flask
- Distilled water bottle
- 250 ml beaker
- 10 ml transfer pipette
- Pipette bulbs
- 20 ml graduated cylinder
- 3 x 125 ml Erlenmeyer flasks
- 3 test tubes
- Test tube rack

Concentration Procedure

1. Obtain a sample of $CuCl_2$ from the front bench.
2. Using your scale weigh out 12.0g of $CuCl_2$.
3. Dissolve the 12.0g of $CuCl_2$ in 100ml of distilled water.
4. Stir until you have reached a homogeneous solution with all the solute dissolved in the solvent.
5. Calculate the concentration of your solution. **M = mol $CuCl_2$ / L**
   Show your work in the space provided below.

Hint: Copper chloride is typically found in its hydrated form. This means two water molecules are bound to each formula unit of $CuCl_2$. To calculate the molar mass of the $CuCl_2 \cdot 2H_2O$ you must include the molar mass of copper chloride plus the molar masses of two water molecules.

Dilution Procedure

6. Pour your copper chloride solution into a 250 ml beaker to begin your dilutions.
7. Rinse your volumetric flask with a few ml of distilled water.
8. Using a 10 ml pipette transfer 20ml of your solution back into your volumetric flask.
9. Fill your volumetric flask to the calibration line with distilled water. Using the equation **$C_1V_1=C_2V_2$** calculate the concentration of your diluted solution?

10. Using a 10 ml pipette transfer 20ml of your original solution from your beaker into a clean125 ml Erlenmeyer flask.
11. Using the equation $\mathbf{C_1V_1=C_2V_2}$ calculate the volume of the total solution (V_2) required to make a **0.3 *M*** solution. C_1 is the concentration of your initial solution. Show your work.

12. Add enough water to your 20ml of original solution to reach the new volume (V_2). Measure your volume of water with a graduated cylinder. The volume of water (Vw) that needs to be added to dilute the solution is $Vw=V_2-V_1$. What is Vw?

13. Using a 10ml pipette transfer 10ml of your 0.8 *M* solution into a large test tube and put it aside. Make sure you remember which test tube contains which concentration. At the end of the lab you will have 3 different concentrations of aqueous copper chloride in 3 different test tubes.

14. Repeat steps 10-13 using a clean 125 ml Erlenmeyer flask to obtain a **0.2 *M*** solution.
    Solve for V_2 and Vw. Show your work.

15. Repeat steps 10-13 using a clean 125 ml Erlenmeyer flask to obtain a **0.15 *M*** solution.
    Solve for V_2 and Vw. Show your work.

16. Bring your 3 test tubes to the front of the room to compare the colours to previously prepared diluted solutions. Have Mrs. Kornelsen check your dilution.

17. Fill in the blanks. We diluted our solutions by adding water. The more water we added the ________________ the concentration was and the _________________ dilute the solution was.

18. Please clean up your lab bench and complete this lab.

Lab Safety

Chemistry Lab Safety Checklist

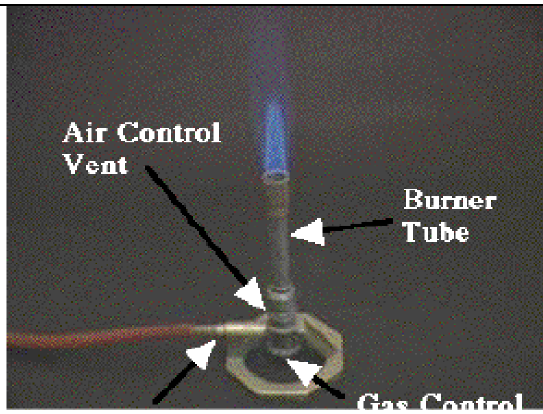
- ✓ Follow your teacher’s directions.
- ✓ Notify your teacher of problems.
- ✓ Know how to use the safety equipment in the lab.
- ✓ Wear approved safety goggles.
- ✓ If you have long hair, tie it back.
- ✓ Avoid awkward transfers of chemicals.
- ✓ If it’s hot, let it cool.
- ✓ Carry chemicals with caution.
- ✓ Dispose of chemical wastes properly.
- ✓ Clean up afterward.

Lab Safety Sheets: Bunsen Burner Use

PERSONAL PROTECTION EQUIPMENT (PPE)	
Safety glasses must be worn at all times in work areas.	Long and loose hair must be tied back
Appropriate footwear must be worn. Shoe must be fully enclosed. No open toed shoes.	Lab coat or apron (chemical resistant)
Rings and jewelry (long necklaces / bracelets, etc.) must not be worn.	Chemical resistant disposable gloves may be required.

HAZARDS PRESENT	ADDITIONAL REQUIREMENTS
✓ Chemical spill ✓ Broken glass ✓ Burns ✓ Explosion	✓ PPE as per MSDS ✓ WHMIS training ✓ Emergency spill kit, sand , kitty, litter ✓ Training in the use of a Bunsen Burner

SAFE WORK PROCEDURE
PRE JOB STEPS: Ensure that the gas tap is in the "Off" position Secure the rubber tubing from the Bunsen burner to the gas tap



Natural gas is fed into the burner through the gas inlet. The gas control needle valve controls the rate at which methane enters the burner. The rate at which air enters the burner is adjusted with the air control vent. Methane and oxygen mix in the burner tube and, when ignited, produce a flame

Procedure:

Lighting a Burner

1. Ensure that the air holes are closed

   Warning: Burns are the most common form of laboratory accident. After heating an object be extremely careful to let it cool before grasping it. Temperatures in the hottest region of the burner flame approach ~1500°C.

2. Before using a Bunsen burner, be certain that no flammable materials are present in the laboratory. Also be careful to make sure that your face, clothing and hair are not above or near the opening of the burner tube.

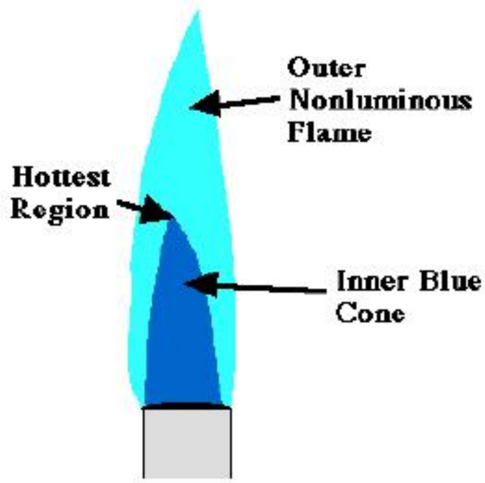
3. Light a match by striking it away from your body.

4. Turn on the gas supply line

5. Place the lighted match about 5 centimeters directly above the Bunsen burner. The Bunsen burner will light to give a yellow safety flame

Adjusting the Burner

After lighting the burner, it must be properly adjusted. In a properly adjusted Bunsen burner, there will be a blue flame containing two or more cones.

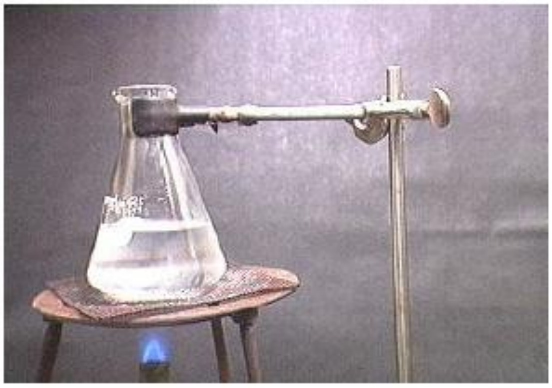


When adjusting the air vent, be careful not to extinguish the flame or disassemble the burner. Increasing the air supply will produce a hotter heat. With the air supply at the half position the flame will be difficult to see. With the air holes fully open there is a lighter visible cone of unburnt gas and the flame will make a "roaring" sound.

Heating substances

Never use your fingers to hold an object in the Bunsen burner flame. Use a pair of tongs to hold a small, solid object.

If the object is large or a liquid in a flask, use a ring stand and a triangle or wire gauze to hold the object in the flame.



Extinguishing the Burner

After you are finished using the burner, turn the gas completely off to extinguish the flame.

Be sure to completely close the gas supply to prevent accumulation of methane in the laboratory- a fire and explosion hazard.

Leave anything that has been heated to cool down before it is put away.

REGULATORY REQUIREMENTS

- WS&H Act W210, Section 4, 5, 7, 7.1
- Mb. **Workplace Safety & Health** Regulations 217/2010,
  - ✓ 2.1 Safe Work Procedures
  - ✓ 6.1 Personal Protective Equipment
  - ✓ 8.0 Musculoskeletal Injuries
  - ✓ 35.0 WHMIS
  - ✓ 36.0 Chemical Biological Substances
- Safe Work Bulletin #164 PPE
- Safe Work Bulletin #246 Safe Lifting
- Safe Work Bulletin #104 Emergency Eye Wash

Lab Safety Sheets: Diluting Acids

LOCATION	WRITTEN BY:	APPROVED BY:	DATE CREATED	LAST REVISION
Schools in PTSD	Lorie Carriere	School Principal S&H Committee	January 26, 2012	New

PERSONAL PROTECTION EQUIPMENT (PPE)	
Safety goggles / face shield must be worn at all times in work areas.	Long and loose hair must be tied back
Appropriate footwear must be worn. Shoe must be fully enclosed. No open toed shoes.	Lab coat or apron (chemical resistant)
Rings and jewelry (long necklaces / bracelets, etc.) must not be worn.	Chemical resistant disposable gloves are required.

HAZARDS PRESENT	ADDITIONAL REQUIREMENTS
✓ Chemical spill ✓ Chemical burns ✓ Corrosive substances ✓ Generates heat when in contact with water ✓ May splatter with water ✓ Strong fumes	✓ PPE as per MSDS ✓ WHMIS training ✓ Emergency spill kit, sand , kitty, litter ✓ Chemical handling

SAFE WORK PROCEDURE

PRE JOB STEPS:

- Supervisors must ensure that employees required to work with chemicals are in possession of or have access to the material data sheet prior to being required to work with the chemical.
- Supervisors must ensure that employees required to work with chemicals are provided with the appropriate PPE i.e. chemical gloves goggles, masks, etc.
- Supervisors must ensure that employees are provided with the proper PPE and are trained in its safe use, inspection, replacement and are encouraged to use the appropriate PPE at all times.
- Workers are to inspect the condition of their PPE prior to use and to store their PPE safely when not in use and to immediately replace their PPE as soon as it's damaged or lost.

- Employees must ensure they do not mix chemicals unless they have read & understood the data which will inform them if chemicals can be mixed and if so how to carry out the operation safely

- Employees shall not decant and store chemicals in unmarked/unsuitable containers

- Supervisors must ensure that employees have a procedure in place for dealing with a chemical spillage.

- Supervisors must ensure that once a spillage is contained and cleaned up, that there is a procedure in place for the safe disposal of the contents /material ,chemical or other contaminates

Ensure that a workplace label has been placed on the decanted container into which the chemical will be decanted prior to starting the decanting process.

Procedure:

1. Turn on the fume hood 5 minutes prior to starting work.
2. Don all personal protective equipment prior to starting work.
3. Place all equipment and materials into the hood.
4. In fume hood measure out 2/3 of distilled water into volumetric flask.
5. Using a graduated cylinder, measure amount of acid needed.
6. Slowly pour acid into water.
7. Add water to fill up to measuring line.
8. Pour solution into storage bottle.
9. Place identification label onto storage bottle including the strength of the new solution.

REGULATORY REQUIREMENTS

- WS&H Act W210, Section 4, 5, 7, 7.1
- Mb. **Workplace Safety & Health** Regulations 217/2010,

  - ✓ 2.1 Safe Work Procedures
  - ✓ 6.1 Personal Protective Equipment
  - ✓ 8.0 Musculoskeletal Injuries
  - ✓ 35.0 WHMIS
  - ✓ 36.0 Chemical Biological Substances

- Safe Work Bulletin #164 PPE
- Safe Work Bulletin #246 Safe Lifting
- Safe Work Bulletin #104 Emergency Eye Wash

Lab Safety Sheets: Fume Hood

PERSONAL PROTECTION EQUIPMENT (PPE)	
Safety glasses or face shield must be worn at all times in work areas.	Long and loose hair must be tied back
Appropriate footwear must be worn. Shoe must be fully enclosed. No open toed shoes.	Lab coat or apron (chemical resistant)
Rings and jewelry (long necklaces / bracelets, etc.) must not be worn.	Chemical resistant disposable gloves are required.

HAZARDS PRESENT	PERSONAL PROTECTIVE EQUIPMEN REQUIRED	ADDITIONAL REQUIREMENTS
✓ Chemical spill ✓ Chemical burn ✓ Chemical contact with eyes ✓ Inhalation of chemical fumes ✓ Ingestion ✓ Fire / explosion	✓ PPE as per MSDS ✓ Lab coat or apron ✓ Disposable gloves (based on chemical) ✓ Face shield ✓ Safety goggles emergency spill kit	✓ Equipment orientation ✓ WHMIS training ✓ Fume Hood Maintenance ✓ Emergency spill clean up

SAFE WORK PROCEDURE

OPERATIONAL SAFETY CHECKS:

1. Do not overload the work surface with apparatus or work material. Safe operation of the hood is based on having proper air flow through the structure.
2. Do not store containers or supplies against the baffle; this will affect airflow through the hood. Blocking the bottom of the baffle will change the airflow pattern in the hood causing turbulence and possible leakage at the face of the hood.
3. Do not use flammable or explosive materials in the hood unless it is equipped with explosion proof components.
4. Do not work with or store chemicals in the hood without operating the exhaust system
5. Perchloric acid must not be used in the hood.
6. Radioisotope materials are not recommended for use in the hood.
7. The use of heat generating equipment in the hood without operating the exhaust system may cause damage to the hood liner.

A fume hood is a laboratory safety device designed to protect the operator from exposure to hazardous chemicals (ex. toxic, corrosive, flammable, malodorous, fumes, vapours, aerosols or dusts) used in the workplace. This is accomplished by creating a shield from hazardous chemical work and exhausting the contaminants, thus providing containment. It is important that all potentially harmful chemical work be conducted inside a properly functioning fume hood with the sash set at the proper level.

GENERAL USE:

1. Workers must perform all work that involves hazardous and noxious materials in the fume hood.

2. Have specific MSDS readily available, read all precautions listed.
   a. If the MSDS states that the chemical is volatile or produces harmful vapours, all work must be done in an approved fume hood.
   b. If corrosive gases or solids are involved use the fume hood.
   c. If using substances with toxic vapours or dusts use the fume hood.
   d. Whenever possible substitute hazardous chemicals with those that are less hazardous.

3. Don all personal protective equipment: safety glasses, footwear (closed toe / heal shoes), and disposable gloves. Remove all jewelry; loose clothing and tie back hair.

4. Do not use a fume hood that has been clearly labeled as out-of-service.

5. Verify that the fume hood is working by turning on the blower fan. If the hood is equipped with a Phoenix control monitor it should read 80 – 120 fpm.

6. Turn the hood blower fan on at least 5 minutes prior to working in the hood.

7. Always work with the sash as low as possible, ~12" (certification sticker height) but never above the sash stop or the posted maximum safe working height (18"). This ensures proper air flow and protection of the user. The sash can be raised above the operating height of 18 inches for experiment set up ONLY and only if absolutely necessary.

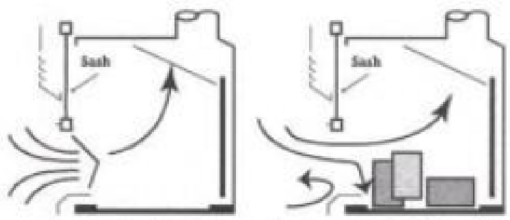
8. The sash should be in the fully closed or lowered positions whenever the teacher is not actively performing work inside the fume hood. Vapors and other contaminants are contained in the hood with greater efficiency at lower sash heights. The sash is also considered a physical barrier providing additional protection to the user from splash and other adverse conditions (e.g., fires, implosions, etc.).

9. Work extending arms under the sash, with the glass between the worker and the chemical source. Place hands and arms into a fume hood only when absolutely necessary. Use slow movements to avoid creating unnecessary air turbulence within the hood. **Do not place your head in the fume hood**.

10. Put the minimum amount of materials in the hood required for the current operation. Each additional item in the hood creates additional turbulence and potential for gas/vapour escape.

11. To ensure adequate containment of vapors, position all work and equipment at least 6 to 8 inches behind the fume hood sash/opening. Vapors and other contaminants can "spill out" of the hood if the generating activity is placed closer than 6" to the sash opening.

12. The airfoil below the sash and the baffles at the back of the fume hood are critical to effective air flow and the containment efficiency of the hood. Be sure to keep these areas clear. Do not block them with equipment, paper towels, absorbent pads, etc. Do not adjust hood baffles.



Demonstration of Airflow Patterns through empty & cluttered fume hoods.

13. Do not put large equipment in the fume hood cupboard, as they block the baffles and produce regions of zero or low flow in the work space.

14. Ensure the room ventilation is on whenever the fume hood is in use.

15. Minimize foot traffic in front of hoods while chemical operations are in process, particularly if the sash is not fully closed. The air current created when walking past a fume is at least 2x the typical face velocity of air entering the hood, thus momentarily causing "spill out" from the hood. Spill out can also occur as you create air currents when walking away from a hood, making it important to close the sash before you walk away.

16. Move with caution, do not make quick motions into or out of the hood (can cause airflow disturbances).

17. When moving the sash up or down, use slow and deliberate motions to minimize air turbulence and the potential for "spill out."

18. Keep items in the fume hood to an absolute minimum. Items within the fume hood will impede air flow and compromise capture efficiency.

19. Do not place electrical apparatus or other ignition sources inside the hood when flammable liquids or gases are present.

20. Do not use equipment that prevents the sash from closing.

21. Keep laboratory doors and windows closed. Open windows and nearby fans or other sources of air turbulence can also negatively impact fume hood performance.

22. Do not compromise the structural integrity of the hood by drilling holes into the side walls or sash; removing sash pieces or other equipment, etc. **Reminder the hoods are PTSD property!**

23. DO NOT use fume hoods as general storage locations for chemicals or equipment. If the hood is not being used during the school year **DO NOT** use it for storage as this can damage the hood, ventilation system and sash.

24. When no processes are being conducted in the fume hood, shut-off the exhaust fan and close the sash. However, observe the following conditions and precautions when shutting-off a fume hood:

Conditions:

- Remove chemicals that present physical and health hazards from the hood and store safely and in their designated storage location.
- No chemicals should remain in the hood. Only if absolutely necessary, and only for short time periods, should any chemicals remain in the hood. When this occurrence is required, they must be of absolute minimum quantities, stored in a tightly closed container, and in observance of chemical compatibility considerations with respect to other materials in the cabinet;
- DO NOT shut off the ventilation at any time. A fume hood must not be used for storage as the airflow is disrupted by the containers and this causes inadequate containment of fumes and increases personal exposure.

25. If any fume hood is suspected of not operating properly, discontinue its use and contact the PTSD Safety and Health Officer to arrange for testing.

26. Fume Hood must be cleaned & waste containers disposed of after each use.

GENERAL MAINTENANCE:

All fume hoods shall be inspected and certified annually to determine a proper face velocity of 80-120 fpm.

All fume hoods functioning properly shall have a certification label affixed to the sash height at which the fume hood was certified.

If a fume hood fails certification, a warning sign shall be placed at a prominent location on the sash of the fume hood indicating that the fume hood should not be used until it has been serviced and is working properly.

This sign shall ONLY be removed by the Safety & Health Officer once the fume hood has passed certification requirements.

REGULATORY REQUIREMENTS

- WS&H Act W210, Section 4, 5, 7, 7.1
- Mb. **Workplace Safety & Health** Regulations 217/2006, Part 16, Sections 16.1 – 16.18
  - ✓ 2.1 Safe Work Procedures
  - ✓ 6.1 Personal Protective Equipment
  - ✓ 35.0 WHMIS
- CSA Z316.5-04 Fume Hoods & Associated Exhaust Systems
- Manitoba Science Safety - Chapter 6 Chemical Spills Clean Up.
- MB. Fire Code 2005
- Lab. Fume Hoods Recommended Practices SEFA 1
- Safe Work Bulletin #164 PPE

Outcomes Checklist

Introduction

This outcomes checklist will be a tool that you can use to help in your self assessment as well. The checklist will be an excellent way for you to keep track of areas you are strong in and others where you need more practice or extra help. At the end of the unit check off the outcomes you feel you have mastered. Mastered means that you can complete the outcome correctly most of the time. By the end of the semester there may be outcomes that you feel you have not mastered yet and we will know that we need to really focus on them before the exam.

You will be given some time in class to complete this checklist and I will be collecting and assessing them at the end of the each unit. You will be marked on the accuracy by which you identify each outcome and show your understanding of it. Please see the example outcome below and the two examples of ways a student could show their having accomplished that outcome. Both responses would receive full marks.

Outcome	Review Example	Mastered
C11-1-01 Describe the properties of gases, liquids, solids, and plasma. Include: density, compressibility, diffusion	In the stations lab I demonstrated that I clearly understood the difference between the compressibility of different states of matter.	

OR

C11-1-01 Describe the properties of gases, liquids, solids, and plasma. Include: density, compressibility, diffusion	Solids are the most dense while liquids have a medium density and gases are not dense at all.	

States of Matter

Outcome	Review Example	Mastered
C11-1-01 Describe the properties of gases, liquids, solids, and plasma. Include: density, compressibility, diffusion		
C11-1-02 Use the Kinetic Molecular Theory to explain properties of gases. Include: random motion, intermolecular forces, elastic collisions		
C11-1-03 Explain the properties of liquids and solids using the Kinetic Molecular Theory.		
C11-1-04 Explain the process of melting, solidification, sublimation, and deposition. Include: freezing point, exothermic, endothermic		
C11-1-05 Use the Kinetic Molecular Theory to explain the processes of evaporation and condensation.		
C11-1-06 Define vapour pressure. .		
C11-1-07 Define normal boiling point temperature in terms of vapour pressure.		
C11-1-08 Interpolate and extrapolate the vapour pressure and boiling temperature of various substances from pressure vs temperature graphs.		

Gases

Outcome	Review Example	Mastered
C11-2-01 Identify the abundances of the naturally occurring gases in the atmosphere and examine how these abundances have changed over geologic time.		
C11-2-02 Research Canadian and global initiatives to improve air quality, and effects on air quality.		
C11-2-03 Examine the historical development of the measurement of pressure.		
C11-2-04 Describe the various units used to measure pressure and be able to convert between them.		
C11-2-05 Experiment to develop the relationship between the pressure and volume of a gas.		

C11-2-06 Experiment to develop the relationship between the volume and temperature. Include: the determination of absolute zero, and the Kelvin temperature scale.		
C11-2-07 Experiment to develop the relationship between the pressure and temperature of a gas.		
C11-2-08 Solve quantitative problems involving the relationships among the pressure, temperature, and volume of a gas.		
C11-2-09 Identify various industrial, environmental, and recreational applications of gases. *Examples: self-contained underwater breathing apparatus (scuba), air bags, propane appliances, hyperbaric chambers…*		

Reactions

Outcome	Review Example	Mastered
C11-3-01 Determine average atomic mass using isotopes and their relative abundance. Include: atomic mass unit (amu)		
C11-3-02 Research the importance and applications of isotopes. *Examples: nuclear medicine, stable isotopes in climatology, dating techniques…*		
C11-3-03 Write formulas and names for polyatomic compounds.		
C11-3-05 Write and classify balanced chemical equations from written descriptions of reactions. Include: polyatomic ions		

C11-3-06 Predict the products of chemical reactions, given the reactants and type of reaction. Include: polyatomic ions		
C11-3-04 Calculate the mass of compounds in atomic mass units.		
C11-3-11 Determine empirical and molecular formulas from percent composition or mass data.		

Stoichiometry

Outcome	Review Example	Mastered
C11-3-07 Describe the concept of the mole and its importance to measurement in chemistry.		
C11-3-08 Calculate the molar mass of various substances.		
C11-3-10 Solve problems requiring interconversions between moles, mass, volume, and number of particles.		
C11-3-12 Interpret a balanced equation in terms of moles, mass, and volumes of gases.		

C11-3-13 Solve stoichiometric problems involving moles, mass, and volume, given the reactants and products in a balanced chemical reaction.		
C11-3-14 Identify the limiting reactant and calculate the mass of a product, given the reaction equation and reactant data.		
C11-3-15 Perform a lab involving mass-mass or mass-volume relations, identifying the limiting reactant and calculating the mole ratio. Include: theoretical yield, experimental yield		

Solutions

Outcome	Review Example	Mastered
C11-4-01 Describe and give examples of various types of solutions. Include: all nine possible types		
C11-4-07 Differentiate among saturated, unsaturated, and supersaturated solutions.		
C11-4-02 Describe the structure of water in terms of electronegativity and the polarity of its chemical bonds.		
C11-4-03 Explain the solution process of simple ionic compounds, using visual, particulate representations and chemical equations (dissociations).		
C11-4-04 Explain heat of solution with reference to specific applications. *Examples: cold packs, hot packs…*		
C11-4-06 Construct, from		

experimental data, a solubility curve of a pure substance in water.		
C11-4-08 Use a graph of solubility data to solve problems.		
C11-4-09 Explain how a change in temperature affects the solubility of gases.		
C11-4-10 Explain how a change in pressure affects the solubility of gases.		
C11-4-11 Perform a lab to demonstrate freezing-point depression and boiling-point elevation.		

C11-4-12 Explain freezing-point depression and boiling-point elevation at the molecular level. *Examples: antifreeze, road salt…*		
C11-4-14 Solve problems involving calculation for concentration, moles, mass, and volume.		
C11-4-15 Prepare a solution, given the amount of solute (in grams) and the volume of solution (in millilitres), and determine the concentration in moles/litre.		
C11-4-16 Solve problems involving the dilution of solutions.		
C11-4-17 Perform a dilution from a solution of known concentration.		

Organic Chemistry

Outcome	Review Example	Mastered
C11-5-01 Compare and contrast inorganic and organic chemistry.		
C11-5-02 Identify the origins and major sources of hydrocarbons and other organic compounds. Include: natural and synthetic sources		
C11-5-03 Describe the structural characteristics of carbon. Include: bonding characteristics of the carbon atom in hydrocarbons (single, double, triple bonds)		
C11-5-04 Compare and contrast the molecular structures of alkanes, alkenes, and alkynes. Include: trends in melting points and boiling points of alkanes only		
C11-5-05 Name, draw, and construct structural models of the first 10 alkanes.		
C11-5-06 Name, draw, and construct structural models of branched alkanes.		

C11-5-07 Name, draw, and construct structural models of isomers for alkanes up to six-carbon atoms. Include: condensed structural formulas		
C11-5-08 Outline the transformation of alkanes to alkenes and vice versa. Include: dehydrogenation/hydrogenation, molecular models C11-5-10 Differentiate between saturated and unsaturated hydrocarbons.		
C11-5-09 Name, draw, and construct molecular models of alkenes and branched alkenes.		
C11-5-12 Name, draw, and construct structural models of alkynes and branched alkynes.		
C11-5-15 Write condensed structural formulas for and name common alcohols. Include: maximum of six-carbon parent chain, IUPAC nomenclature		

C11-5-13/18/21/22 Compare and contrast the structures of aromatic hydrocarbons, organic acids, esters, and polymers.		
C11-5-23 Describe how the products of organic chemistry have influenced quality of life. *Examples: synthetic rubber, nylon, medicines, polytetrafluoroethylene (Teflon®)…*		