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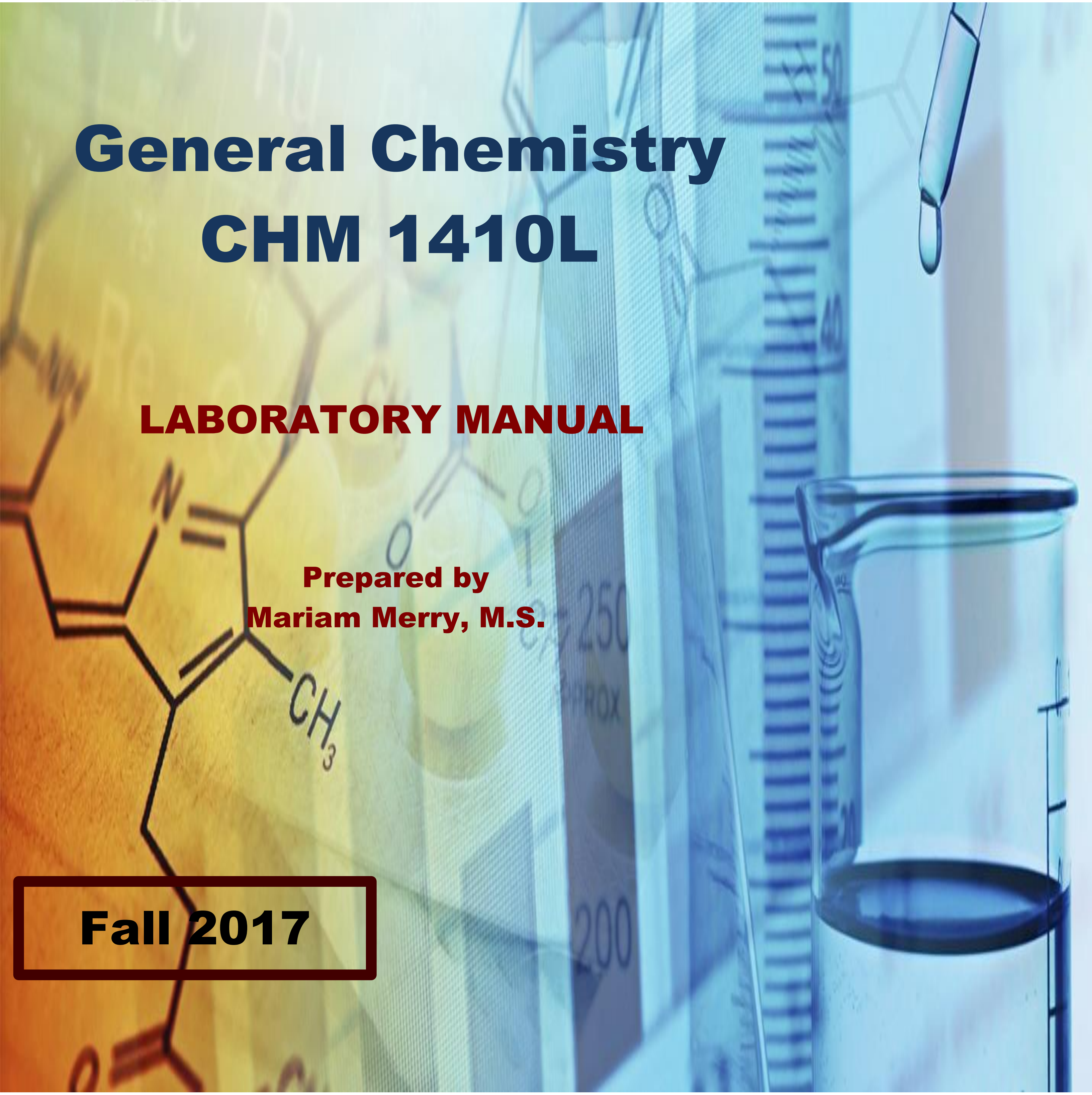

General Chemistry
CHM 1410L

LABORATORY MANUAL

Prepared by
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Fall 2017

Preface and Acknowledgments

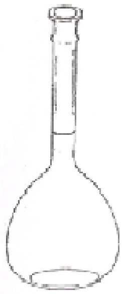
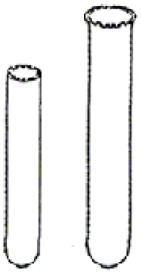
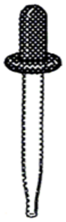
The first version of General Chemistry I Lab Manual was drafted by Ms. Khalida Jaffar and Dr. Khalid Omer. It was used in the academic year, 2012-2013. The manual in your hands, 2015-2016, has been improved based on the first version, experience learned in the first year, and what I think is more affiliated with the theoretical subjects taught in the class. Therefore, the Manual aims to guide students in conducting experiments through the semester. Whenever it is needed, the students need to refer to the textbook, Chemistry: the Central Science, 13th Edition. Theodore E. Brown, H. Eugene H. LeMay, Bruce E. Bursten and Catherine Murphy. Pearson Education Inc. 2014.

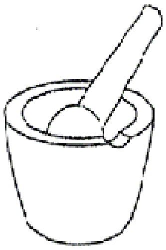
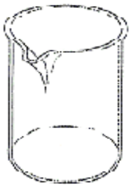
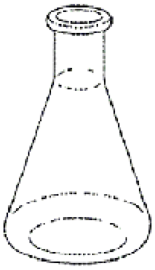
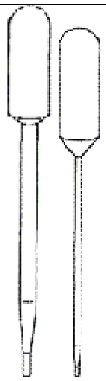
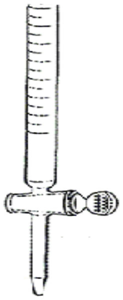
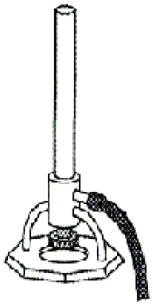
This manual was completed with support of KUST administration, faculty, and staff members. Therefore, I am thankful for their efforts.

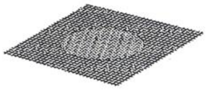
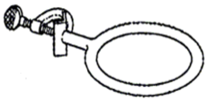
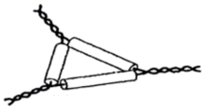
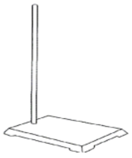
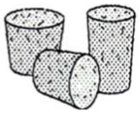
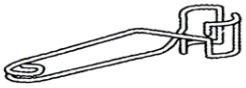
Table of Content

Common Laboratory Apparatus*

Laboratory apparatus and equipment like electronic scales, glassware, burners, graduated cylinders, and more, help the scientist conduct experiments, observe data, heat liquids, make measurements, and more.

Items Name	Picture	Items Name	Picture
Volumetric Flask		Test Tubes	
Watch Glass		Crucible and lid	
Dropper		Evaporating dish	
Funnel		Graduated cylinder	

Items Name	Picture	Items Name	Picture
Mortar and pestle		Beaker	
Stirring rod		Erlenmeyer Flask	
Micropipette		Burette	
Thermometer		Bunsen burner	

Items Name	Picture	Items Name	Picture
Wire gauze		Clamp	
Iron ring		Clay triangle	
Stand		Tong	
Wash bottle		Safety goggles	
Corks		Rubber Stopper	
Brush		Forceps	
Test tube holder		Spatulas	

Items Name	Picture	Items Name	Picture
Centrifuge		Oven	
Water bath		Electronic Lab Scale	
Hot plate		pH-meter	

* Matin J., Martin C. (2012). Catalyst: the Pearson Custom Library for Chemistry: Laboratory Experiments, Pearson Prentice Hall. ISBN13: 978-0-536-93404-8

Chemistry Lab Safety

The chemistry laboratory can be a place of discovery and learning. However, by the very nature of laboratory work, it can be a place of danger if proper common-sense precautions aren't taken. While every effort has been made to eliminate the use of explosive, highly toxic, and carcinogenic substances from the experiments which you will perform, there is a certain unavoidable hazard associated with the use of a variety of chemicals and glassware. You are expected to learn and adhere to the following general safety guidelines to ensure a safe laboratory environment for both yourself and the people you may be working near. Additional safety precautions will be announced in class prior to experiments where a potential danger exists.

Laboratory Attire

1. Safety goggles *must be worn at all times* while in the laboratory. This rule must be followed whether you are actually working on an experiment or simply writing in your lab notebook. You must wear safety goggles provided by the chemistry department.

2. Contact lenses are not allowed. Even when worn under safety goggles, various fumes may accumulate under the lens and cause serious injuries or blindness.

3. Closed toe shoes and long pants must be worn in the lab. Sandals and shorts are not allowed.

4. Long hair must be tied back.

Laboratory Conduct

1. Eating, drinking, and smoking are strictly prohibited in the laboratory.

2. No unauthorized experiments are to be performed. If you are curious about trying a procedure not covered in the experimental procedure, consult with your laboratory instructor.

3. Coats, backpacks, etc., should not be left on the lab benches and stools. Beware that lab chemicals can destroy personal possessions.

4. Always wash your hands before leaving lab.

5. Be especially careful of spills around the balance. These electronic devices are *extremely* sensitive to corrosion. A brush is kept near the balance so you can brush the balance thoroughly after *each* use.

6. Clean any spill near the balance *immediately* and report it to instructor.

7. Notify the instructor immediately in case of an accident.

8. Consider *all* chemicals to be hazardous as a result never taste anything, never directly smell the source of any vapor or gases, by means of your couple hands, bring a small sample to your nose, Fig 2.2.1

9. Know what chemicals you are using. Carefully read the label *twice* before taking anything from a bottle.

Fig 2.2.1 Waft Toward Your Nose

10. Excess reagents are **never** to be returned to stock bottles. If you take too much, dispose of the excess.

11. Many common reagents, for example, alcohols and acetone, are highly flammable. *Do not use them anywhere near open flames.*

12. Always pour acids into water. If you pour water into acid, the heat of reaction will cause the water to explode into steam, sometimes violently, and the acid will splatter.

13. If chemicals come into contact with your skin or eyes, *flush immediately* with copious amounts of water and consult with your instructor.

14. Never point a test tube or any vessel that you are heating at yourself or your neighbor, Fig. 2.2.2.

Fig. 2.2.2 Pointing a Test Tube at Your Neighbor

15. Dispose of chemicals properly. Waste containers will be provided and their use will be explained by your TA. Unless you are explicitly told otherwise, assume that only water may be put in the lab sinks.

16. Clean up all broken glassware immediately and dispose of the broken glass properly.

17. Never leave burners unattended. Turn them off whenever you leave your workstation. Be sure that the gas is shut off at the bench rack when you leave the lab.

18. Beware of hot glass--it looks exactly like cold glass.

Fire Case

1. In the event of fire, *do not panic*.

2. If a small portion of your clothes catches fire, the fire may be extinguished by patting it out.

3. *Never* use a fire extinguisher on a person. Carbon dioxide fire extinguisher is extremely cold and may cause shock to the person, frostbites, or harm to the eyes.

4. If a fire should occur in a breaker or some other container, cover it with a glass dish or other flame-retardant item.

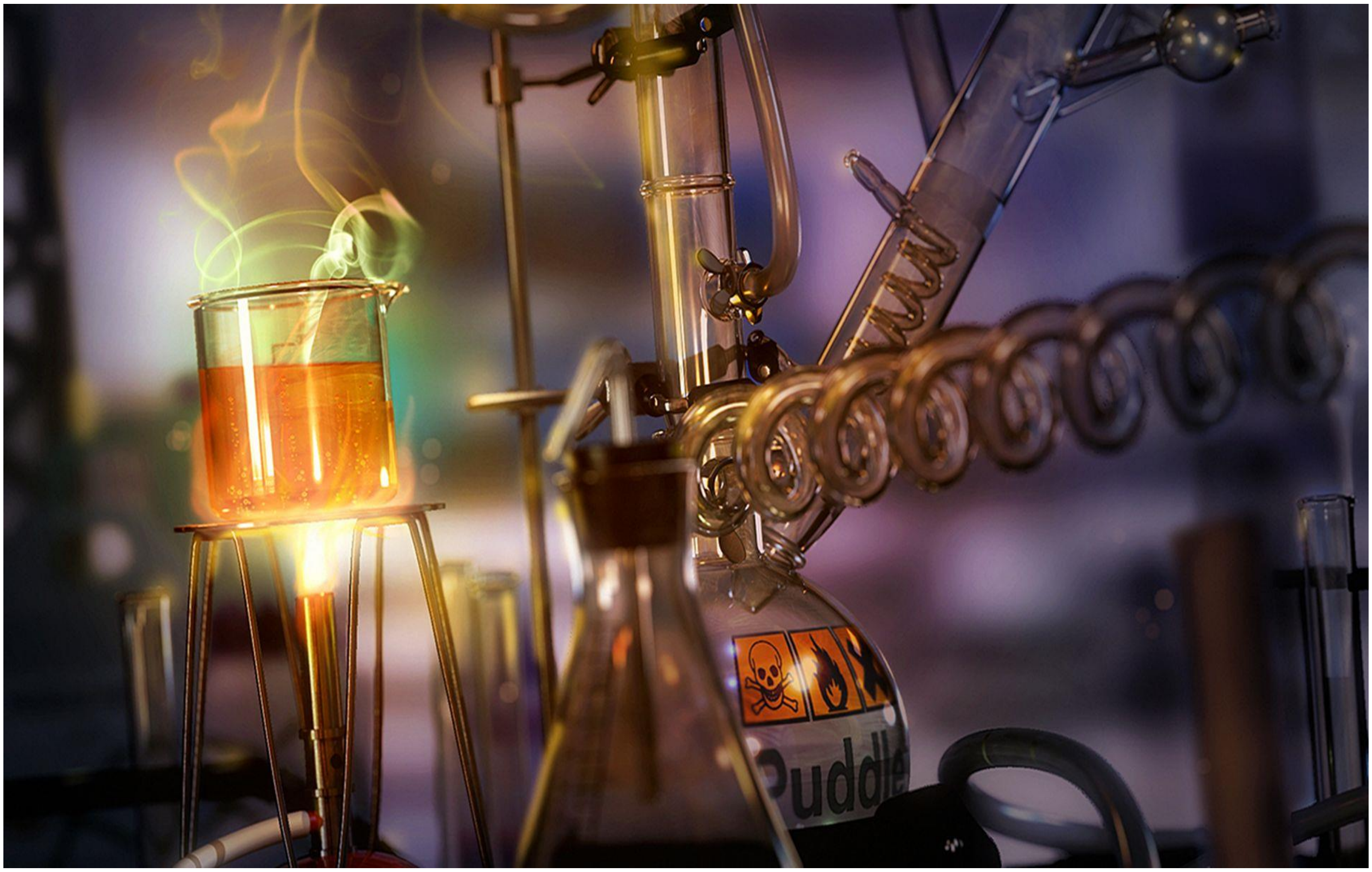
Writing Lab Reports

The lab report for this course will be represented by a designed report sheet prepared by the course instructor and accompanying with each experiment. This sheet will illustrate the understanding of the experiment via answering some questions related to the work and it also summarize the data via tables with simple calculations whenever it is needed and simple discussion .

Submission of Lab Reports

Each group should submitted one report and the lab report of previous week experiment should be submitted to the lab assistant at the beginning of the next experiment lab work. Any delay in submission is not acceptable unless with reasonable cause. Late submission will cause deduction from your grade

Laboratory Experiments



Experiment 1: Measurements and Calculating Density of Liquids and Solids

Note*: *The significant figure rules apply to all the measurements done in the lab.

Objectives

1. To learn how to use some of the common lab supplies.
2. To learn how to use the proper units of the SI system and the use of proper significant figures
3. Determine the densities of different substances in liquid and solid state.

Introduction

Density is a basic physical property of a homogeneous substance; it is an intensive property, which means it depends only on the substance's composition and does not vary with size or amount. The determination of density is a nondestructive physical process for distinguishing one substance from another.

Density is the ratio of a substance's mass to its own volume. In the metric system the unit of density for a liquid or solid is measured in g/mL or g/cm^3.

$$d = \frac{mass}{volume}$$

In this experiment students will determine the density of two liquids and compare the physical properties of those liquids. The volume of a liquid is measured using a graduated cylinder, a pipette, or some other volumetric apparatus. Also Students will determine the density of a solid object. The volume of a regular solid (e.g. a cube or a sphere) may be determined by measuring its dimensions and then calculating it using the correct mathematical formula. The difficulty in determining the volume of an irregular solid in this manner is obvious. The method commonly used is to measure the change in the volume of water when the object is immersed in the water. The object displaces a volume of water equal to its own volume.

If the solid material is soluble in water, another liquid, in which the solid is insoluble, is used (e.g. carbon tetrachloride for salt).

Pre-Lab Questions

1. How would you measure the volume of sample of sand?
2. Why can density be used as a means for identification?

Which tool consider more accurate: graduated beaker, graduated cylinder, pipette, burette.

Table 1.1: Chemicals and Supplies

Chemicals	Supplies
Tap Water	25 mL graduated cylinder (1)
	Piece of wood as a rectangular shape
	Marble
	Piece of metal as a cylindrical shape
	Ruler
	Calipers

Procedure

Part I: Density of Liquids

1. Using the electronic scale, measure the mass of an empty, clean and dry graduated cylinder. Record the result in table 1.2

2. Add 15 mL of distilled water to the graduated cylinder. Precisely measure this volume of water. Record the precise measured volume of water in table 1.2

3. Measure the mass of water and cylinder using the scale and record the result in table 1.2.

Part II: The Density of a solid from the Thickness.

1. Bring a piece of wood from your instructor. Use the ruler to measure the length, width and height of the wood. Record the result in table 1.3.

2. Measure the mass of the wood using the scale and record the result in table 1.3.

3. Use the marble and the calipers to find the diameter of this small ball then repeat step 2 part II.

Part III: The Density of a solid from displacement of water.

1. Bring a piece of a metal from the instructor and measure its mass then record the result in table 1.4.

2. Use the graduated cylinder to measure 15 mL water.

3. Immerse the piece of the metal inside the cylinder and record the new volume.

Report Sheet 1 of Exp. 1: (Attach any extra paper to this sheet)

Group Names:

Part A: Collecting Data

Table 1.2: Part I-Density of Water

Mass of empty graduated cylinder (g)	
Mass of graduated cylinder and water (g)	
Mass of water (g)	
Volume of water (mL)	

Table 1.3: Part II: The Density of a solid from the Thickness

	Wood		Marble
Length		Diameter	
Width		Radius	
Height		Mass of the marble	
Mass of the wood			

Table 1.4: The Density of a solid from displacement of water

	Metal
Mass of the metal	
Volume of water only, v1	
Volume of water with metal, v2	
V of metal = v2 – v1	

Part B: Answer the following:

1. Calculate the density:
   a. Water from part I

   b. Wood from part II

   c. Marble from part II.

   d. Metal from part III.

2. A piece of metal cylinder has a height of 10.0 cm and a diameter of 4 cm. a student measure its mass and found to be 18.715 g. what is the density of this cylinder?

Experiment 2: Separation of the Component Mixture

Objective

To separate the components of a mixture depending on their physical property

Introduction

Mixtures are not unique to chemistry; we encounter them on a daily basis. The food and drinks we consume, the fuel we use in our vehicles, building materials, soil, air, and ordinary water are all mixtures. In fact, very few common materials in the world are pure. *Any material made up of two or more substances that are not chemically combined is a mixture*. It is not difficult to separate the components of a mixture if you take advantage of differences in *physical properties* of the components.

1. *Sublimation.* Some substances such as iodine, caffeine, naphthalene, and dry ice; change directly from a solid to a gas when they are heated. The reverse process, when the vapor goes back to the solid phase without a liquid state in between, is called deposition.

2. *Extraction.* It uses a solvent to selectively dissolve one component of the solid mixture. With this technique, a soluble substance can be separated from an insoluble substance.

3. *Decantation.* It separates a liquid from insoluble solid sediment by carefully pouring the liquid from the solid without disturbing the solid. Fig.2.1

4. *Filtration.* It separates a solid from a liquid through the use of a porous material such as a filter paper, charcoal, or sand can serve as a filter. These materials trap the solid but let the liquid pass through.

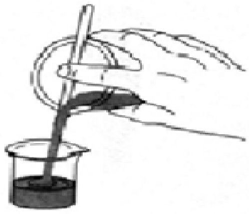


Fig.2.1 Decant Process

5. *Evaporation.* This is the process of heating a mixture in order to remove a liquid that evaporates easily, leaving a dry residue.

In this experiment, the mixture contains three components: naphthalene, $C_{10}H_8$, common table salt, NaCl, and sand, SiO_2. Fig 2.2 is a flow chart summarizes the sequence of steps that you will use:

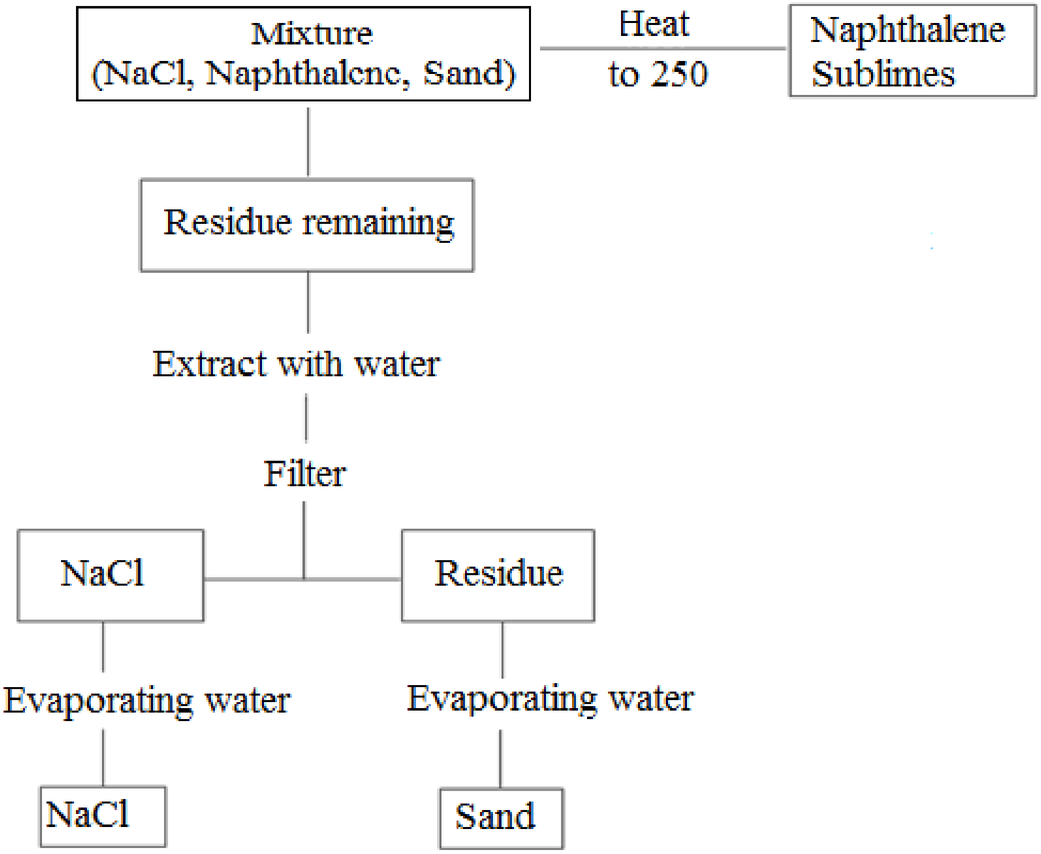


Fig.2.2 Flow Chart Summarizes the procedure

Pre-lab question

How you can separate out the following mixture:

a. Lead filling and iron filling.
b. Sand and gravel.
c. Sand and Styrofoam (both are finely powdered)
d. Salt and sugar.
e. Alcohol and water.

Table 2.1: Chemicals and Supplies

Chemicals	Supplies
Distilled Water	Bunsen burner and its sets
Mixture of (Naphthalene, Sand, Salt) (prepared by TA)	Watch glass
Ice	Filter paper
	Mortar and pestle
	50 mL Beaker (3)
	Funnel
	Dropper
	Tong
	Stirring rod

Procedure

1. Label one of the three given beakers as “beaker #1”.
2. Put the mixture in the mortar for grinding and start grind until your mixture become finely powdered.
3. Empty all the mortar contents in beaker #1.
4. Place a clean, dry watch glass on top of beaker #1 that contains the mixture.

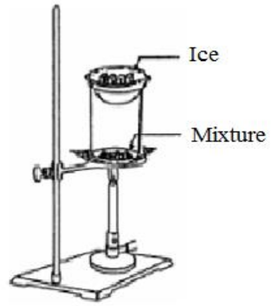


Fig.2.3 The Experiment Setup

5. Set up the burner and put beaker #1 on the wire gauze as Fig.2.3 shows.
6. Put 1 – 2 ice cubes on the top of the watch glass.

Note: ***be careful not to get any water inside the beaker; use the dropper to get rid of the melted water obtained from the ice.***

7. Heat the beaker gently. Vapors will appear in the beaker and solid crystals should collect on the underside of the watch glass. (You may want to wear a mask).

   <u>Caution</u>: **Naphthalene is *flammable*. Be sure to keep the *flame* low. Do not let the flame come into contact with vapor that escapes from the beaker. *Turn off* the burner each time you remove the evaporating dish for scraping.**

8. After 5 minutes of heating, turn off the Bunsen burner.

9. Collect the crystallized solid by scraping it off the glass into a piece of paper with a stirring rod.

10. Stir the content of the beaker with a glass rod then add more ice cubes, in case your previous ice cube is melted, to the watch glass and place it back on the beaker again.

11. Repeat steps 7-10 until no more solid crystals collected on the underside of the watch glass.

12. Let the beaker cools down completely.

13. Measure 25 ml of water using the graduated cylinder and add it to beaker #1. Stir the mixture with a glass rod to thoroughly dissolve the salt.

14. Label another clean beaker as “beaker #2”.

15. Assemble the apparatus for filtration. Fig.2.4.

16. Fold the filter paper. Fig.2.5.

17. Decant the mixture in beaker #2 into the filter paper.

18. Rinse Beaker #2 with water until all its content transferred to the filter paper.

Fig.2.4 Filtering Apparatus

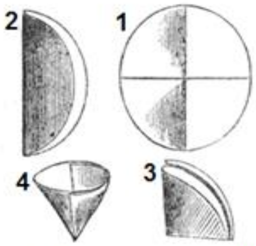


Fig.2.5 ow to old a Filter Paper

19. Turn on the Bunsen burner and start heating beaker #2 gently until all the water will evaporate.

20. When all the water evaporates, turn off the burner and put the beaker aside to cool down.

21. Label the last clean beaker as “beaker #3”.

22. Transfer the sand from the filter paper to beaker #3 using the stirring rod.

23. Turn on the Bunsen burner again and start heating beaker #3 gently until all the water will evaporate.

24. When all the water evaporates, turn off the burner and put the beaker aside to cool down.

25. Clean your mess

Report Sheet of Exp. 2: (Attach any extra paper to this sheet)

Group Names:

Part A: Answer the following questions

1. How could you separate barium sulfate from ammonium chloride?
2. How could you separate tellurium dioxide from silicon dioxide?
3. How could you separate Naphthalene from potassium bromide?
4. Suppose a lab group reports a percent yield of sand of 105%. Is it really possible to collect more sand than was originally present? What is a possible explanation for the “extra” product? Do not consider calculation mistakes; think about the procedure used.
5. Why we put a piece of ice cube on the top of the watch glass?
6. Think about the lab procedure and what you know about types of mixtures.
   a. Name one heterogeneous mixture present in this exp. ____________________
   b. Name one homogeneous mixture present in this exp. ____________________
   c. Name one pure substance present in this exp. _____________

Part B: What makes you able to separate the mixture into its constituent substances?

Experiment 3: Difference between Ionic and Covalent Compounds

Objectives

1. Classify compounds as either primarily ionic or primarily covalent from the name, formula and the behavior of the compounds.
2. Observe and record some properties of several ionic and covalent compounds.
3. Recognize patterns among the properties and distinguish ionic compounds from covalent compounds.

Introduction

A compound is defined as a chemical combination of two or more elements. A chemical bond is the “glue” holding together atoms of different elements. Two types of bonds are *ionic* and *covalent*. Ionic bonds generally occur between a metallic atom and a nonmetallic atom. The bond results from the transfer of one or more electrons from the metallic atom to the nonmetallic atom, resulting in a charge difference. The positively charged metal *ion* is then attracted to the negatively charged nonmetallic *ion*. Covalent bonding generally occurs between two or more nonmetallic atoms. Covalent bonding involves the sharing of electrons.

Properties such as melting point, boiling point, solubility, electrical conductivity, color, and odor can help you distinguish ionic from covalent compounds. As in many areas of chemistry, the distinctions are not always clear, nor do the distinctions apply to all compounds.

The salt and sugar on your kitchen table both dissolve easily in water, but the solutions they form have an important difference. One of those kinds of white crystals is an ionic compound, and when it dissolves, it *dissociates*, or breaks up into ions. The ions are free to move in the solution, and that solution, therefore, conducts electricity. The more ions in solution, the better it conducts electricity. If something produces a large amount of ions it is called a *strong electrolyte*. If something produces a small amount of ions, it is called a *weak electrolyte*. The other kind of crystal, however, is a molecular compound, and its molecules remain whole when they dissolve. With no ions, that solution does not conduct electricity. If something produces no ions, it is called a *non-electrolyte*.

In this experiment, you will observe several properties of some ionic and some covalent compounds and attempt to recognize patterns among the properties. The patterns you may recognize are generalizations and may not necessarily apply to all ionic and covalent compounds.

Pre-lab questions

Read the introduction and procedure before answer the following questions:

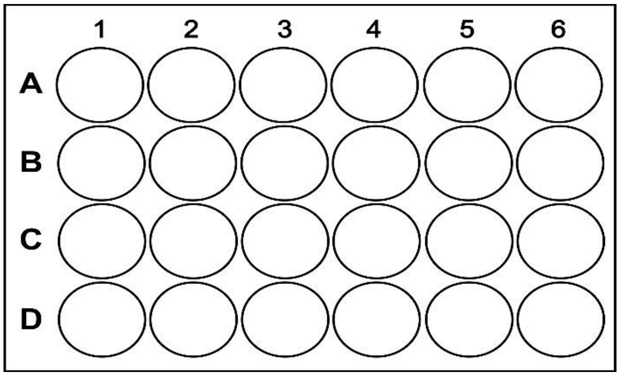
1. What kinds of elements generally form ionic compounds and covalent compounds?
2. Based on the formula, predict whether each of the following compounds is primarily ionic or primarily covalent.
   a. sodium iodide NaI
   b. methane CH_4
   c. calcium chloride $CaCl_2$
   d. ammonia NH_3
   e. glucose $C_6H_{12}O_6$

Table 3.1: Chemicals and Supplies

Chemicals	Supplies
Solid Benzoic acid, C_6H_5COOH	Well plate
Solid Magnesium chloride, $MgCl_2$	Toothpicks
Vegetable oil	Conductivity tester
Solid Potassium iodide, KI	
Solid Sodium sulfate, Na_2SO_4	
Solid Sugar, $C_{12}H_{22}O_{11}$	
Petroleum Oil	
Distilled water	

Procedure

The TA will give the students a well plate labeled as showing below:



1. To the first column, (#1), add a very few grains of benzoic acid to A1, and B1.

2. To the second column, (#2), add very few grains of $MgCl_2$ to A2, and B2

3. To the third column, (#3), add 15 drops of vegetable oil to A3, and B3

4. To the forth column, (#4), add very few grains of KI to A4, and B4

5. To the fifth column, (#5), add very few grains of Na_2SO_4 to A5, and B5,

6.To the sixth column, (#6), add very few grains of sugar to A6, and B6.

Note: be careful not to mix the spatulas to avoid the contamination among substances.

7. To the first row, A, add 20 drops of distilled water to each cell, A1-A6. Stir with a toothpick.

Note: do not stir each solution with the same toothpick

8. To the second row, B, add 20 drops of Petroleum oil to each cell, B1-B6. Stir with a toothpick.

9. Record the observation of the relative solubility of each of the solid using the two different solvents in table 3.2.

10. Use the light bulb as a conductivity tester and test the conductivity of each cell. Record your data in table 3.3

11. Clean your mess

Report Sheet of Exp. 3: (Attach any extra paper to this sheet)

Group Names:

Part A: Fill the tables below

***Table 3.2: Collecting Data for solubility**

Compounds	Distilled Water solvent	Petroleum Oil solvent
Benzoic acid		
Magnesium chloride		
Vegetable Oil		
Potassium iodide		
Sodium sulfate		
Sucrose		

[*] <u>***Note***</u>**:** Describe the solubility of substance by using: *not at all*, *partially*, or *completely*.

***Table 3.3: Collecting Data for conductivity**

Compounds	Distilled Water solvent	Petroleum Oil solvent
Benzoic acid		
Magnesium chloride		
Vegetable Oil		
Potassium iodide		
Sodium sulfate		
Sucrose		

[*] <u>***Note***</u>**:** describe the conducts electricity: *no conductivity*, *low conductivity*, or *high conductivity.*

Part B: Answer the following

1. What do you conclude from this experiment?

Experiment 4: Determine the Limiting Reactant and the Percent Yield

Objectives

1. Obtain hands on experience with the limiting reagent problem.
2. Interpret the meaning of an experimentally measured percent yield.

Introduction

This experiment is designed to illustrate the relationship between quantities of reactants and the amount of product produced by a chemical reaction. The principles of stoichiometry and limiting reactants will be used to predict the amount of product produced by mixing two solutions. The reaction to be studied is as follow:

$$CaCl_2\,_{(aq)} + Na_2CO_3\,_{(aq)} \longrightarrow 2\,NaCl\,_{(aq)} + CaCO_3\,_{(s)}$$

The balanced chemical equation for this reaction can be expressed in a net ionic form as follow:

$$Ca^{2+}\,_{(aq)} + CO_3^{-2}\,_{(aq)} \longrightarrow CaCO_3\,_{(s)}$$

It is often difficult and impractical to combine the right amount of two reactants needed for a particular reaction to occur. This experiment is designed for one reactant to go to completion. The reactant that goes to completion is defined as *limiting reactant* because it limits the amount of product being formed. Since the other reactant will have a quantity remaining, it's called the *excess reactant*.

The two chemicals that are used in this experiment make common uses in everyday life. In a solid state, sodium carbonate is known as "washing soda" and is used to enhance the effectiveness of laundry soap. Calcium chloride acts as a desiccant (drying agent), which is used by recreational vehicle owners to remove moisture from the air in the vehicle during winter season. The formed product, $CaCO_3$, is a component of classroom chalks.

From the given date, you will be asked to calculate the mole of each chemical added for the reaction to occur. Using the principles of stoichiometry, you will be able to determine which chemical is the limiting reactant and thereby predict how much precipitate will be formed. This stoichiometric determination will then be compared to the actual mass of $CaCO_3$ formed.

Chemists are often concerned about the optimal yield of manufacturing certain chemicals. One way to predict the amount of chemical produced, chemists use percent yield of that particular chemical using the following formula:

$$\text{Percent yield} = \frac{\text{Actual weight product in gram}}{\text{Theoretical weight of the same product in gram}} \times 100$$

Pre-lab Questions

1. Why we have to grind calcium chloride before starting the experiment?
2. What's the benefit of adding sodium hydroxide and sulfuric acid?
3. What would be the effect on the yield if your filter paper were not completely dry?

Table 4.1: Chemicals and Supplies

Chemicals	**Supplies**
Sodium carbonate Na_2CO_3	Funnel
Calcium chloride $CaCl_2$	Disposable plate (1)
Distilled Water	25.0 mL volumetric flask (2)
0.1 M NaOH (Prepared by TA)	25 mL graduated cylinder (1)
0.1 M H_2SO_4 (Prepared by TA)	Filter paper (1)
	Small test tubes (2)
	100ml beaker (2)
	Erlenmeyer flask (1)

Procedure

***Note*: show your calculations to your instructor before you start the work. Any mistake the group will be responsible for ignoring consultation.**

1. Measure the mass of 0.210 M of calcium chloride (M.wt from the bottle label) dissolves in 25.0 mL volumetric flask.

2. Add a reasonable amount of water with a gentle shaking to the volumetric and when all the salt dissolve add more water until it reach to the mark. (You may need to use the dropper for adding few drops to the mark).

3. Transfer the solution from the volumetric flask to a labeled beaker (#1).

4. Measure the mass of 0.336 M of sodium carbonate (M.wt from the bottle label) and dissolve it in 25.0 mL volumetric flask and repeat step 2.

5. Transfer the solution in step 3 from the volumetric flask to another labeled beaker (#2).

6. Combine calcium chloride solution and sodium carbonate solutions by adding the solution in beaker #1 to beaker #2. A white precipitate should be formed. You may need to wait several minutes for the precipitate to coagulate sufficiently to filter well.

7. Measure the mass of a clean piece of filter paper and write the mass of the filter paper, the group names, and trial number #1 on the edge of the filter paper using pencil.

8. Fold the filter into quarters and put it into the funnel. Wet the filter with distilled water, and let it drain.

9. Set the filtration apparatus, Fig 4.1.

10. Pour the white mixture you get from step 7 through the filter paper. The filter will stop the calcium carbonate and any extra water, dissolved sodium chloride, and excess reactants will pass through the filter and collected as a filtrate in the beaker, Fig 4.1.

11. Save the first filtrate that you collect in the Erlenmeyer flask. Rinse your beaker out with distilled water to make sure all of the calcium carbonate is washed into the filter.

12. Rinse the filter itself with a small amount of distilled water to remove any sodium chloride from the filter.

13. Store your filter paper on a disposable plat and allow it to "air-dry" over the weekend.

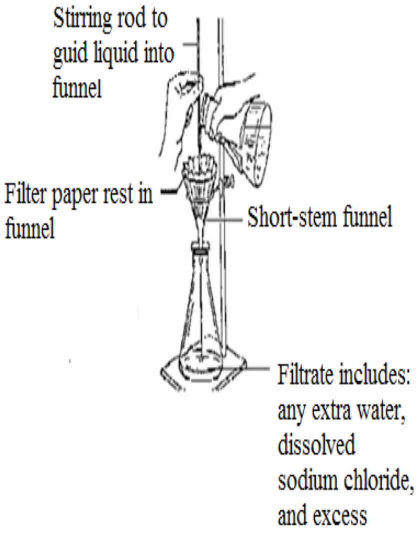


Fig 4.1 Filtration Apparatus

14. Divide the filtrate in the Erlenmeyer flask into two portions in small-size test tubes.

15. Treat one portion with a few milliliters of sodium hydroxide and record any observation.

16. Treat the second portion with a few milliliters of sulfuric and record any observation.

17. Clean your mess.

Report Sheet of Exp. 4: (Attach any extra paper to this sheet)

Group Names:

Part A: Fill the following table

Table 4.2: Collecting Data

Mass of $CaCl_2$	Mass of Na_2CO_3	Theoretical Mass of $CaCO_3$	Actual Mass of $CaCO_3$

Part B: Do the following calculations:

1. Calculate the theoretical weight of $CaCO_3$. Show your calculation steps
2. Calculate the % yield of $CaCO_3$.
3. Find the % error using the formula below:

$$\%\ \text{Error} = \left| \frac{\text{Theoretical Value} - \text{Experimental Value}}{\text{Theoretical Value}} \right| \text{X } 100$$

Part C: Answer the following

1. Which substance is the limited reactant and which one was in excess?
2. Discuss your %error.

Experiment 5: Reaction in Aqueous Solutions

Objectives

1. Determine the species present in aqueous solutions of compounds using the solubility rules.
2. Predict the type of the reaction that will occur when two aqueous solutions are mixed.
3. Write the chemical equation, the ionic equation, and the net ionic equation for reactions taking place between aqueous solutions.

Introduction

One of the most important characteristics of water is its ability to dissolve many compounds. Solutions in which water is the solvent are called aqueous solutions. Many important reactions take place in aqueous solutions. In fact, many of the reactions that take place throughout your body (from your organs down to individual cells) are aqueous reactions. Understanding the most common aqueous reactions and how to correctly write them is one of the most important skills you should master in Chemistry. One must ask the question: what observations indicate that a chemical reaction has occurred? Some indications include: the formation of an insoluble solid (precipitate), or color change, the evolution of a gas, or a temperature change. In this experiment, you will predict what happen when two aqueous solutions are mixed, and then test your prediction in laboratory.

In metathesis or double displacement reactions, cations and anions exchange partners as in the following generic reaction:

$$\mathbf{A}X + \mathbf{B}Y \longrightarrow \mathbf{A}Y + \mathbf{B}X$$

There are three types of metathesis reactions - precipitation reactions, gas-forming reactions, and neutralization reactions. The forth reaction that you will perform is totally different because it is a type of an electron transfer reaction. Students must be familiar with the solubility rules and be able to use then efficiently.

Table 5.1: Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds	**Important Exceptions**
Compounds containing	
NO_3^-	None
CH_3COO^-	None
Cl^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
Br^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
I^-	Compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}
SO_4^{2-}	Compounds of Sr2$^+$, Hg_2^{2+}, $Pb^{2+}$and Ba^{2+}
Insoluble Ionic Compounds	**Important Exceptions**
Compounds containing	
S^{2-}	Compounds of NH_4^+, the alkali metal cations, and Ca^{2+}, Sr^{2+}, and Ba^{2+}
CO_3^{2-}	Compounds of NH_4^+, and the alkali metal cations.
PO_4^{3-}	Compounds of NH_4^+, and the alkali metal cations.
OH^-	Compounds of NH_4^+, the alkali metal cations, and Ca^{2+}, Sr^{2+}, and Ba^{2+}

Pre-lab questions

1. Write the molecular equation, ionic equation, and net ionic equation for the following reactions:
   a. Sodium chloride and ammonium nitrate.
   b. Sodium sulfate and silver nitrate.
2. Write the precipitation reaction equations in (part I) from the procedure and predict your observation.

Table 5.2: Chemical and Supplies

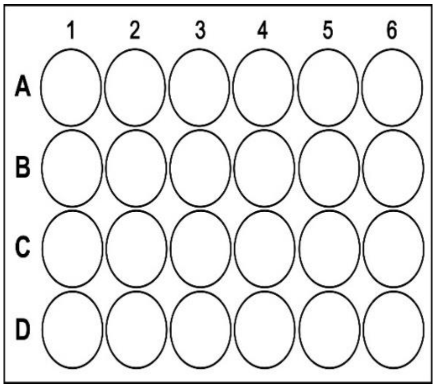
***Chemicals**	**Supplies**
0.1 M Na_2SO_4	Well plate (1)
0.1 M NaOH	droppers
0.1 M Na_2CO_3	
Cu wire	
Conc. HNO_3	
Solid $Mg(OH)_2$	
2 M HCl	
0.1 M $AgNO_3$	
Solid $CaCO_3$	
(ph-ph) indicator	
0.1 M NH_4NO_3	

***Note: All solutions in this experiment should be prepared by TA**

Procedure

Part I: Precipitation Reactions, Fig 5.1.

Fig 5.1 Precipitation Reaction



1. Using the well plate, add 10 drops of Na_2SO_4 to cell A1
2. Add 10 drops of NaOH to cell B1.
3. Add 10 drops of Na_2CO_3 to cell C1.
4. To the first column (#1), add 10 drops of 0.1 M $AgNO_3$ and record your observation.

Part II: Gas-Formation Reaction, Fig 5.2

Fig 5.2 Gas-Formation Reaction

1. Using a new clean cell, add few grams of $CaCO_3$ to cell A2.
2. Add 1 mL of water to step 1 and make sure all the $CaCO_3$ is dissolved.
3. Add 15 drops of 2 M HCl to the cell A2 and record your observation.

Part III: Neutralization Reaction, Fig 5.3

Fig 5.3 Neutralization Reaction

1. In another clean cell, add few grams of $Mg(OH)_2$ to cell A3.
2. Add 1 mL of water to step 1 and make sure all the $Mg(OH)_2$ is dissolved.
3. Add 1-2 drops of phenolphthalein indicator to cell A3 and record your observation.
4. Add 15 drops of 2 M HCl to the cell A3 and record your observation.

Part IV: Oxidation-Reduction Reaction, Fig 5.4

***Note*: this reaction must be performed inside the hood and you should wear mask.**

Fig 5.4 Redox Reaction

1. Put a small piece of copper wire in A4 cell.
2. Add 15 drops of concentrated HNO_3 to A4 cell.
3. Record your observation.

Report Sheet of Exp. 5: (Attach any extra paper to this sheet)

Group Names:

Part A: Collecting Data

Table 5.3: Recording the Reaction Observations

Part I: Precipitation reactions	
	$AgNO_3$
Na_2SO_4	
NaOH	
Na_2CO_3	
Part II: Gas formation reaction	
	HCl
$CaCO_3$	
Part III: Neutralization reaction	
	HCl
$Mg(OH)_2$	
Part IV: Oxidation-reduction reaction	
	HNO_3
Cu	

Part B: Write a balanced chemical equation for each reaction you did in each part (except Part IV) and indicate the state of both reactants and products.

Part C: Write a ***balanced net ionic equation*** for each of the following aqueous metathesis reactions. (*Be sure to include all states for reactants and products*). Then classify each reaction as a neutralization, precipitation, or gas-forming reaction:

a) Chromic acid and cesium hydroxide Classification: ------------------------------
b) Sulfuric acid and sodium carbonate Classification: ------------------------------
c) Cadmium chloride and sodium sulfide Classification: ------------------------------

Experiment 6: Making a Solubility Curve

Objectives

1. Examine the effect of temperature on the solubility of different salts
2. Study the effect of some factors that enhance the solubility.
3. Be able to identify and understand the key terms: solubility, solute, solvent, solvation, saturated, unsaturated and supersaturated solutions

Introduction

Solutions are homogeneous mixtures of **solvents** (the larger volume of the mixture) and **solutes** (the smaller volume of the mixture). For example, a hot chocolate is a solution, in which the solute (the chocolate powder) is dissolved in the solvent (the milk or water). The solute and solvent can be either a solid, liquid or a gas. A solution forms when the attractive forces between the solute and the solvent are similar. For example, the ionic or polar solute, NaCl, dissolves in water, a polar solvent. The phase "like dissolves like" has often been used to explain this.

As the water molecules collide with the ionic compound (NaCl), the charged ends of the water molecule become attracted to the positive sodium ions and negative chloride ions. The water molecules surround the ions and the ions move into solution. This process of attraction between the water molecules (the solvent) and the ionic compound (NaCl, the solute) is called **solvation**. Solvation continues until the entire crystal has dissolved and all ions are distributed throughout the solvent.

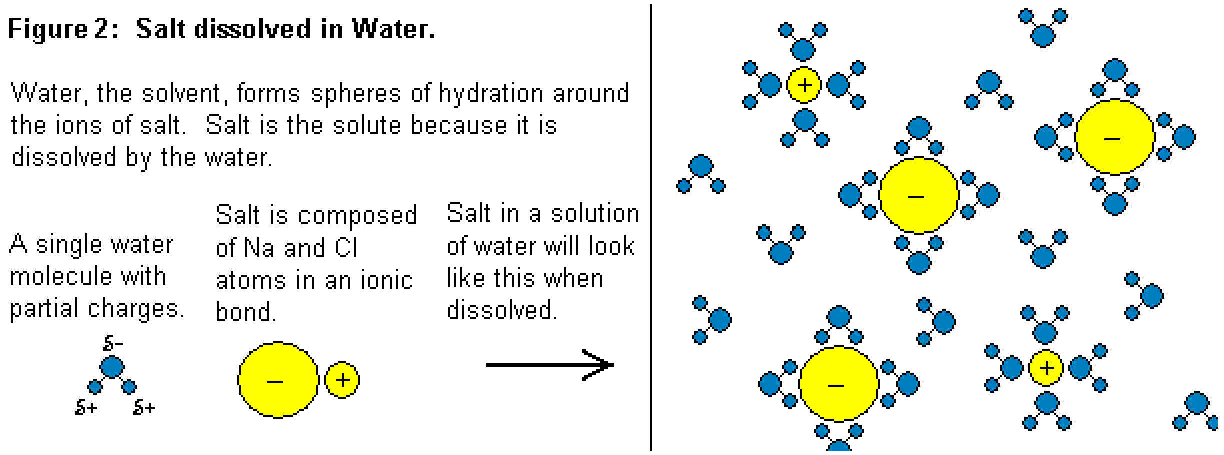


Figure 2: Salt dissolved in Water.

Some solutions form quickly and others form slowly. The rate depends upon several factors, such as, the size of solute, stirring, or heating. When making hot chocolate, we stir chocolate powder into hot milk or water. When a solution holds a maximum amount of solute at a certain temperature, it is said to be **saturated**. If we add too much chocolate powder to the hot milk, the excess solute will settle on the bottom of the container. Generally, the chocolate powder dissolves better in hot milk than cold milk. Thus, heating the solution can increase the amount of solute that dissolves. Most solids are more soluble in water (solvents) at higher temperatures.

Solubility is the quantity of solute that dissolves in a given amount of solvent. The solubility of a solute depends on the nature of the solute and solvent, the amount of solute, the temperature and pressure (for a gas) of the solvent. **Solubility** is expressed as the quantity of solute per 100 g of solvent at a specific temperature.

Pre-lab question

a. Does the solubility of all ionic solids increase as the temperature increases?
b. Give an example of a salt that is less soluble when the temperature increases.
c. Why is a mixture of ice and water used to make the freezing ice-water bath?
d. Define the following: saturated, unsaturated and super-saturated solutions?

Table 6.1: Chemicals and supplies

Chemicals	Supplies
Sodium chloride, NaCl	100 mL Beaker (3)
Potassium chloride, KCl	50 mL graduated cylinder
Ammonium chloride, NH_4Cl	Thermometer
Calcium chloride, $CaCl_2$	funnel
	Stirring rod
	Ice water bath

Procedure

1 Using a graduated cylinder, measure 50 mL of distilled water and pour it into a 100 mL beaker.

2 Add a piece of ice cube to step 1.

3 Use the stirring rod to stir the ice-water mixture for 1 minute. Insert the thermometer in the beaker to measure the temperature of the mixture. Your temperature reading on the thermometer should reach 0.0 ˚C. Add more ice if all the ice melt before the temperature reading reaches to 0.0 ˚C.

Note: Do not use the thermometer to stir the mixture.

4 Remove the thermometer and the stirring rod from the ice-water mixture and place the watch glass over the beaker. Pour the cold water (without ice) back into the 50 mL graduated cylinder to fill it up to the mark. Once done, pour the content into the second 100 mL beaker provided to you.

5 Place the second 100 mL beaker with its content in a pan containing ice cubes. Surround your beaker with additional ice.

6 Insert the thermometer into this second beaker to get a temperature reading

7 Measure 2.0 g of the solute using a clean beaker and add it to the water in the beaker exist in your pan. Stir the mixture until the solid is dissolved.

8 Repeat step 7 until no more dissolution is observed. Make sure to keep track of the total mass of the solute added to the water and record the total amount in table 6.2**.**

Note: Your solution is considered saturated if some of the solute remains on the bottom of the beaker.

9 Remove the beaker from the ice-water bath (when your solution is saturated) and carefully dry the outside of the beaker with a tissue. Place the beaker on a flame source and make sure to provide a gentle heat.

10 Insert the thermometer in the second beaker to get a temperature reading of 20˚C.

11 Remove and replace the fire source accordingly to maintain a constant temperature of 20˚C.

12 When the undissolved solid from the saturated solution dissolves, add another 2.0 g of the solute to the water. Stir until the salt dissolves.

13 Continue adding 2.0 g increments of the solute until no more solid is able to dissolve in the water. Make sure to keep track of the total mass of the solute added to the water and record the total amount in table 6.2. The solution is saturated again when excess solid is observed in the bottom of the beaker.

14 Repeat steps 10 -13 at temperature readings of 40˚C, 60 ˚C and 80˚C.

15 Remove the beaker from the fire source and gently set it on the lab bench to cool.

Caution: be careful, the beaker is hot.

16 Clean up your mess.

Report Sheet of Exp. 6: (Attach any extra paper to this sheet)

Group Names:

Part A: Fill the following tables

Table 6.2: The Salt assign to the group

Temperature (˚C)	**Mass of solute ( ) added to make a saturated solution (g)/50 mL H_2O**
0.0	
20.0	
40.0	
60.0	
80.0	

Part B: Answer the following:

1. Plot a graph, using graph papers, of the **mass/50 mL** of the solute dissolved (Y-axis) versus temperature (X-axis). Draw a best-fit smooth curve through the data points **table 6.2**.

2. You must discuss the behavior of the salt towards the temperature base on your graph. In another word, How the heat affect the solubility of your salt?

Experiment 7: Freezing Point Depression and Molar Mass

Objectives

1. Determine and compare freezing points of a pure solvent and a solution.
2. Use freezing point depression to determine the molar mass of an unknown substance
3. Explore colligative properties of solutions.

Introduction

When a solute is dissolved in a solvent, the solution properties change from those of the pure solvent. Relative to the pure solvent, the solution has a lower vapor pressure and freezing point, and a higher boiling point. These effects are called *colligative properties* of the solution because they are linked together by a common feature; *they all depend primarily on the ratio of the number of solute particles to the number of solvent particles*. The size of the particles whether they are molecules or ions makes little difference.

The freezing point depression has several important uses. It allows us to compare the freezing-point depression of a known and unknown substance in the same solvent in order to determine the molar mass of the unknown. This type of comparison is called *cryoscopy* (cryo-comes from the Greek word meaning "icy" or "cold").

Cryoscopy is used in clinical laboratories to measure the total concentration of solute in the urine, making use of the fact that the freezing point depression is proportional to the total concentration of all dissolved particles. The urine that is initially formed within the kidney has a much lower concentration of solute than the urine that leaves the kidney. This is because the kidney functions as a filter where it excretes water in the urine as the fluid passes through the kidney tubules. The kidney's ability to produce concentrated urine is one of its most important functions, and when the renal tubes are damaged, the ability to produce concentrated urine is thus lost. Therefore, it's of no surprise that cryoscopy measurement of the solute content in the urine is used to predict kidney function.

The freezing-point depression

The equation that describes the freezing point depression is: $\Delta T = i \cdot K_f \cdot m$

Where ΔT is the freezing-point depression ($\Delta T = T^{\circ}_f - T_f$ where T°_f is the freezing point of the pure solvent, T_f is the freezing point of the solution). Freezing point depression is directly proportion to the concentration of particles in the solution, which is expressed in molality (*m*). The proportionality constant, K_f is called molal freezing point depression constant. By using the molar mass of a known solute, we can determine K_f for a particular solvent. The van't Hoff factor or *i* is used to indicate the number of ions present. The molality (and molar mass) of an unknown compound can be determined by measuring the freezing-point depression of a solution of the unknown in the same solvent.

Note that the phrase *particle molality* is used for the molality of ionic solute, such as NaCl. If the solute does not dissociate into two or more particles, the particle molality equals the molality of the solute. The distinction between particle molality and solute molality is important to keep in mind when dealing with aqueous solution.

Pre-lab Questions

1. If you were asked to choose a solvent for this experiment from the list below, which would you choose? Justify your answer.
   - Carbon tetrachloride K_f= 29.8 ˚C/m
   - Chloroform K_f= 4.90 ˚C/m
   - Ethanol K_f= 1.99 ˚C/m
   - Naphthalene K_f= 6.80 ˚C/m
2. Why do we use molality rather than molarity as our concentration unit for this experiment
3. What are the steps you need to do to determine the molar mass of an unknown solute?
4. What is the Van't Hoff factor for glycerol molecule?
5. Use the molality equation along with the freezing-point depression equation to derive an expression for calculating the molar mass of a solute

Table 7.1: Chemicals and supplies

Chemicals	Supplies
Pure water	400 mL beaker (1)
Glycerol	Test tubes (3)+Rack
Chemical NaCl	50 mL beaker (2)
Ice cubes	10 mL Graduated cylinder (1)
NaCl (food Salt) for ice-water-bath	50 mL Graduated cylinder (1)
	Stirring rod
	Modeling clay
	Thermometer

Procedure:

1. Place an empty 50 mL beaker on the scale and tare the scale.

2. Weigh 2.0 g of sodium chloride until you have obtained the correct amount of NaCl, and set the beaker aside.

3. Assemble the freezing point apparatus by getting a 400 mL beaker and fill half of it with ice. Then, pour a few amount of sodium chloride (not the 2 g you weighed out in Step 2) inside the 400 mL beaker and mixing it thoroughly.

4. Place the thermometer in the ice bath and monitor the temperature. The ice bath is ready when the temperature drops to approximately - 8 °C.

5. While you wait for the temperature to drop, label the test tubes 1 – 3.

6. Use a 50 mL graduated cylinder to measure and pour 10 mL of pure water into test tube 1.

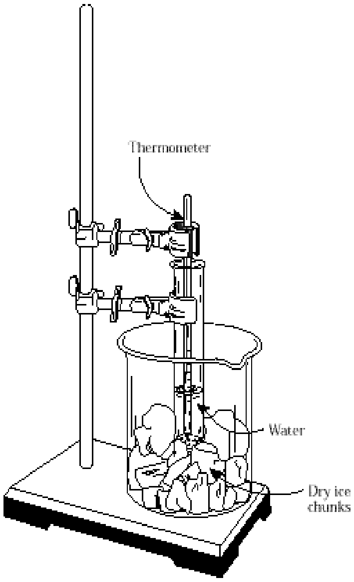
7. After the temperature in the ice bath has reached -10 °C, remove the thermometer from the beaker and rinse it with distilled water.

8. Take the modeling clay and mold it around the top of test tube 1 to create a closed system. Leave an opening for the thermometer. Place the thermometer in the opening and seal the clay around it, making sure that the thermometer is positioned in the water but not touching the bottom of the test tube.

9. Place the test tube in the ice bath, and record the freezing point for the water.

   Note: You will know that the water has hit its freezing point when small ice crystals begin forming in the solution. Record this value in Table 7.2.

10. Next, use the 10 mL graduated cylinder to measure and pour 5 mL of glycerol into another clean 50 mL beaker.



11. Use the 50 mL graduated cylinder to measure and pour 10.0 mL of distilled water into the beaker in step 10 to dilute the glycerol. Gently stir the solution with the stir rod.

12. Pour the solution in step 11 into test tube 2.

13. Remove the clay/thermometer from test tube 1 and remold it on test tube 2. Place test tube 2 in the ice bath, and record the freezing point in Table 7.2.

14. Measure 10 mL of distilled water and add it to the 2.0 grams you measured in step 2 and stir the solution to ensure an even distribution.

15. Pour the solution in step 14 into test tube 3.

16. Repeat Step 13 to determine the freezing point of the NaCl solution.

17. Clean your mess.

Report Sheet of Exp. 7: (Attach any extra paper to this sheet)

Group Names:

Part A: Collecting Data:

Table 7.2: Collecting data

Solution	Freezing point, T_f (˚C)	Van't Hoff factor, i	Experimental Molar mass	Theoretical molar mass	**Information Needed**
H_2O		----------	----------	----------	Density of water = 1.00g/mL
Glycerol		1			Density of Glycerol = 1.26g/cm^3
NaCl		2			K_f of water = 1.86 ˚C/m

Part B: Answer the following:

1. Calculate the theoretical molar mass for each solute.
2. Calculate the experimental molar mass of all the solutes using the derived equation from lecture.
3. Find the % Error of each solute.
4. Discuss your % Error result.

Experiment 8: Enthalpy and Heat Capacity

Objectives

1. Students will be able to define exothermic reaction, endothermic reaction, and the specific heat.
2. Understand the calorimeter principle.

Introduction

This experiment is an introduction to the basic principles of thermochemistry and involves the exchange of energy as heat. The ideas and concepts involved in thermochemistry are illustrated in your everyday experiences. For example, on a hot summer day the hood of a car can get hotter than the sidewalk cement and when cooking, you have probably noticed that a wooden spoon does not heat as fast as metal one. After completing this experiment, you will better understand the reasons behind these and other thermal phenomena.

In the first part of this experiment you will construct a simple "coffee-cup" calorimeter. When used properly, this calorimeter can give very good results. In the next part of the experiment you will measure the specific heat of a calorimeter. In the third part of the experiment, you will determine the enthalpies, ΔHrxn, of exothermic reactions. You will be exploring the factors that cause a reaction to occur.

In order to make sense of your observations for the third part of the experiment you will need to consider an additional concept. In an exothermic reaction, the reaction releases heat, implying that the products are of lower energy than the reactants (ΔHrxn is negative). However, in an endothermic reaction heat is absorbed, indicating that the products are higher in energy (ΔHrxn is positive).

All parts of this experiment require the use of a calorimeter. In the first part of this experiment, you will construct an inexpensive but effective coffee-cup calorimeter. Before you can use this calorimeter you must determine the **heat capacity** of the calorimeter itself. You will do this by adding a weighed sample of hot water to a known amount of cold water in the calorimeter and measuring the temperature change.

The amount of energy required changing the temperature of an object or a sample of a substance by one degree Celsius or Kelvin is called that object's **heat capacity**.

When two substances having different temperatures come into contact, energy in the form of heat is exchanged between them until they reach a common temperature. If they are insulated from their surroundings, the amount of heat lost from the hotter substance equals the heat gained by the colder one. The heat lost or gained is related to the mass, the specific heat of the substance, and the temperature change. This relationship is expressed as:

$$\mathbf{q = m * Csp * \Delta T}$$

Pre-lab questions

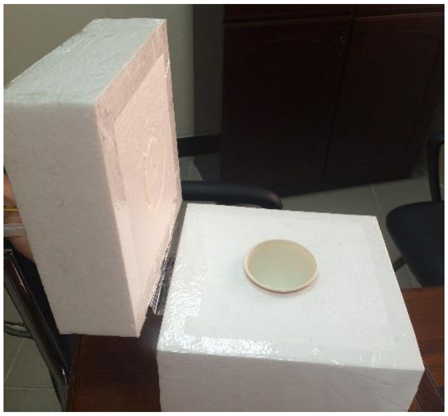
1. In this experiment, why we used two cups of coffee instead of one cup?
2. What makes the temperature reading using the thermometer fluctuate up and down?
3. Calculate the number moles of sodium hydroxide used in part II.

Table 8.1: Chemicals and supplies

Chemicals	Supplies
NaOH	(1) Styrofoam cups with a box
Distilled Water	Boiled water bath (prepared by TA)
	400 mL Beaker (1)
	100mL Beaker (2)
	50 mL graduated cylinder (1)
	Stand and clamp

Procedure

Part I: Determining the Heat Capacity of the Calorimeter



1. Place the coffee-cup (your calorimeter) on the scale and record its mass, this will represent the mass of the calorimeter (mass $_{\text{calorimeter}}$), and calibrate the balance to zero mass. This is called "tarring" the container.

2. Pour distilled water into the cup in step 1 until the scale reads 70 grams. We will refer to this mass as the "mass of the cool water", $m_{\text{cool water}}$.

3. Place the cup back into the calorimeter box as in the right figure and cover the cup by covering the whole box then insert a thermometer through the hole as in the figure below.

4. Make sure the thermometer bulb is covered with water but is not touching the bottom of the cup.

5. Measure and record the temperature of the water in the calorimeter cup. We will refer to this temperature as the initial temperature of the cool water, $Ti_{\text{cool water}}$.

6. Place 100 ml beaker on the scale and tare it then put 30 ml of distilled water in the 100 ml beaker and record the mass of the water. We will refer to this mass as the "mass of hot water, $m_{\text{hot water}}$.

7. Place the 100ml beaker in the boiling water bath for 20 minutes in order for the temperature of water in the 100 ml beaker to equilibrate to the temperature of the boiling water in the water bath.

8. After 20 minutes, measure and record the temperature of the water in the 100 mL beaker with your thermometer. We will refer to this temperature as the initial temperature of hot water, $T_{i\ \text{hot water}}$.

9. Next quickly but carefully pour the 30 mL hot water into the calorimeter cup.

10. Cover the calorimeter with its lid, then adjust the thermometer's height so that it is not touching the bottom or sides of the calorimeter yet the water is covering the thermometer bulb.

11. Gently swirl the water in the calorimeter until the highest temperature is reached. This is the equilibrium temperature. Watch the thermometer closely as it rises. Sometimes it will rise, fall, and rise again due to the initial uneven distribution of heat within the calorimeter. We will refer to this temperature as the final temperature, T_f.

Part II: Calculating the Enthalpy of an Exothermic Reaction

1. Repeat step 1 in part I

2. Weight out about 25 grams distilled water into the calorimeter and record the mass of the water.

3. Repeat step 3 – 4 in part I

4. Measure and record the temperature of the water in the calorimeter. This is the initial temperature of the water, Ti_{water} and calorimeter $Ti_{\text{calorimeter}}$.

5. Tare a clean, dry 100 mL beaker and measure and record the mass of 5 grams of sodium hydroxide.

6. Next, remove the thermometer from the calorimeter box. Carefully, add sodium hydroxide to the calorimeter and cover it again with its lid. Be careful that none of the sodium hydroxide or water spills out of the calorimeter.

7. Replace the thermometer through the hole on the calorimeter box. Adjust the thermometer's height so that it is not touching the bottom or sides of the calorimeter yet the water is covering the thermometer bulb.

8. Gently swirl the water in the calorimeter until the sodium hydroxide is dissolved.

9. Watch the thermometer closely. Stirring the solution in the calorimeter achieves a uniform temperature throughout the calorimeter. Monitor the temperature and record the highest temperature attained. This is the final temperature, T_f.

Report Sheet of Exp. 8: (Attach any extra paper to this sheet)

Group Names:

Part A: Fill the table below

Table 8.2: Collecting Data

Part I: Determining the Heat Capacity of the Calorimeter	
Mass of cool water	
Mass of calorimeter	
Mass of hot water	
Specific heat of water	4.186 J/g.C°
$T_{i\ \text{cool water}}$	
$T_{f\ \text{Cal, cool water, hot water}}$	
$Ti_{\text{hot water}}$	
$\Delta T_{\text{Cal, cool water}} = T_f - T_i$ (for $q_{\text{cool water}}$)	
$\Delta T_{\text{hot water}} = T_f - T_i$ (for $q_{\text{hot water}}$)	
Specific heat of calorimeter	J/g.C°
Part II: Calculating the Enthalpy of an Exothermic Reaction	
Mass of water and NaOH	
$T_{i\ \text{solution}}$	
$T_{f\ \text{solution}}$	
$\Delta T_{\text{solution}} = T_f - T_i$	
$\Delta H = q_{\text{rxn}}$	KJ/mole NaOH

Part B: Do the following calculations

For Part I:

1. Calculate the heat gained by the calorimeter, $q_{\text{calorimeter}}$. This can be done by using the equation: $-q_{\text{hot water}} = (q_{\text{cool water}} + q_{\text{calorimeter}})$
2. Calculate the heat capacity of your calorimeter. Note that the temperature change of the calorimeter is assumed to be the same as the temperature change of the "cool water" in the calorimeter.

For Part II:

1. Calculate the number of moles of sodium hydroxide dissolved.
2. Calculate the heat gained by the chemical system of sodium hydroxide, q_{rxn} using the equation: $-q_{\text{rxn}} = q_{\text{solution}}$ and $\Delta Hrxn = q_{\text{rxn}}$
3. Calculate the enthalpy of the reaction for each mole of sodium hydroxide in unites of Joule/mole.

Experiment 9: Hess's Law

Objective

1. Understand Hess's law definition.
2. Given the heat capacity of calorimeter system, the mole of the reactant participating in the reaction, and the temperature change for the reaction, students should be able to find the values for q and ΔH of the reaction.

Introduction

Energy changes due to the change in the state of the system are associated with chemical processes. A *system* is that region of the universe under consideration. *Surroundings* are everything other than the system.

Calorimetry is a tool used to measure the heat that is generated or consumed by a substance during a chemical reaction or a physical change. When heat is absorbed, the reaction is endothermic and when heat is generated, the reaction is exothermic. Most reactions involve some amount of heat transfer. Therefore, calorimetry has industrial applications in pharmaceutical, chemical, and other biological fields.

Students in the lab will be using a Styrofoam calorimeter due to its ability to minimize the heat exchange with the outside environment and its reliable insulation. The heat capacity (C) of a calorimeter is the amount of heat required to raise the temperature of a calorimeter by one degree Celsius. If we assume that our calorimeter is perfectly insulated. Heat, we assume, will not flow through the walls of the calorimeter. This assumption allows us to restrict the extent of the surroundings. Because heat cannot flow out of or into the calorimeter, we can define the **surrounding** as the complete calorimeter system; this includes the cups, thermometer, and water that made up the solutions that were mixed together.

Systems include any other substances that are contained in the calorimeter. This definition of a system includes substances that dissolve in the water, such as the reactants, and products of a chemical reaction. It also includes water produced in the reaction.

Hess's Law: *If a reaction is carried out in a series of steps, ΔH for the reaction is equal to the sum of the enthalpy changes for the individual steps.*

The overall enthalpy change for the process is independent of the number of steps or the particular nature of the path by which the reaction is carried out. Thus, we can use information tabulated for a relatively small number of reactions to calculate ΔH for a large number of different reactions.

In this experiment, you will measure the enthalpy change for three reactions. The enthalpy change, ΔH, under the lab conditions is equal to the amount of heat lost or gained during the reactions. The First Law of Thermodynamics states that the amount of energy in the universe is constant. If the amount of work done in a process is zero or can be neglected the only form of energy is heat. Thus, the amount of heat in the universe cannot increase or decrease due to the reactions that we performed. One way of saying this mathematically is that the heat lost (or gained) by the system plus the heat lost (or gained) by the surrounding is equal to zero.

q(system) + q(surroundings) = 0 ----------------------------------- equation 1

By convention, heat that is gained is positive and heat that is lost is negative. The equation can be rearranged to give:

q(system) = - q(surroundings) ----------------------------------- equation 2

The heat lost (or gained) by the system is **equal** to the enthalpy change of our process, ΔH. All we need to do is to find the heat lost (or gained) by the surroundings and change the sign. This will be done by multiplying the heat capacity, (c), of the surrounding by the temperature change.

q(surroundings) = m c ΔT --------------------------- equation 3

If we know that the heat capacity of the entire calorimeter system (this includes the cup and the solvent water), we can easily calculate the heat lost (or gained) by the system, q(system), by measuring the temperature increase of the calorimeter and water.

q(system) = - m c ΔT ---------------------------- equation 4

Pre-lab question

1. What is the purpose of this experiment?
2. Using the enthalpy changes given below, calculate the enthalpy change for the following equation:

$$N_2 (g) + 2\ O_2 (g) \rightarrow 2\ NO_2 (g)$$

Using the following equations:

a. $2\ NO (g) \rightarrow N_2 (g) + O_2 (g)$ ΔH = -180.0 kJ

b. $NO_2 (g) \rightarrow NO (g) + 1/2\ O_2 (g)$ ΔH = 112.0 kJ

Table 9.1 Chemicals and Supplies

*Chemicals	Supplies
3.0 M HCl	Styrofoam coffee cups (2)
3.0 M NaOH	Thermometer
3.0 M NH_3	50 mL Graduated cylinder
3.0 M NH_4Cl	Stirring rod

***Note: these solutions should be prepared by the TA.**

Procedure

1 Set up your calorimeter as you did in the previous experiment, see the figures on page 42.

2 Take the Styrofoam cup from the calorimeter box, place it on the balance, record its mass and this will represent the mass of the calorimeter (mass $_{\text{calorimeter}}$).

3 Using a graduated cylinder measure 50.0 mL of 3.0 M HCl.

4 Pour the measured solution in step 3 into your Styrofoam cup, then cover the box and insert a thermometer through the hole in the center of the box.

5 Measure the temperature of the solution in the Styrofoam cup and record this temperature as the initial temperature for the solution, T_i.

6 Rinse the graduated cylinder and measure 50.0 mL of 3.0 M NaOH.

7 Add NaOH solution to the calorimeter cup that contains HCl in step 5.

8 Immediately, cover the box and begin swirling the cup until the thermometer reading stabilized.

9 Read and record the temperature as the final temperature for the solution, T_f.

***Note*: do not stir the cup with the thermometer.**

10 Weigh and record the mass of the calorimeter cup with its contents (100 mL solution).

11 Clean your calorimeter and repeat the steps 3-10 for the following reactions:

Rxn.2: 50.0 mL of 3 M HCl, 50.0 mL of 3.0 M NH_3.
Rxn.3: 50.0 mL of 3 M NH_4Cl, 50.0 mL of 3.0 M NaOH

Report Sheet of Exp. 9: (Attach any extra paper to this sheet)

Part A: Fill the following table

Table 9.2: Collecting Data

	NaOH- HCl	**NH_3-HCl**	**NH_4Cl-NaOH**
Mass of calorimeter cup			
Mass of the calorimeter cup and 100 mL solution			
Initial temperature of solution in calorimeter. Ti (˚C)			
Final temperature of solution in calorimeter. T_f (˚C)			
ΔT (Tf-Ti)			
Heat evolved or absorbed by reaction (KJ/mole)	ΔH_1=	ΔH_2=	ΔH_3=

Part B: Answer the following:

1. Find the mole of any of the reactant.
2. Calculate the ΔH in units of KJ/mole for each reaction. Use 4.186 J/g.C˚ for the heat capacity of solution and the equation: q(rxn) = - q(solution)
1. Compare the sum ΔH of rxn 2 and 3 with ΔH of rxn 1**.** What do you conclude?

Experiment 10: Ideal Gas Law

Objective

Understand the application of ideal gas law

Introduction

Hydrogen peroxide (H_2O_2) is a common household item used to clean minor cuts. It decomposes and forms a gas that is impossible to detect by smell. The bottle of H_2O_2 you purchase from any store contains % H_2O_2. The H_2O_2 slowly decomposes over time to form water and oxygen gas. Oxygen gas is not easily detected when inhaled. Ideal gas law, however, with the use of yeast as a catalyst can help in detecting the amount of oxygen gas decomposed.

Measuring the weight of any gas is difficult. Laws such as the ideal gas law can be used to relate the number of moles to measurements such as mass, volume, pressure, etc.

Since H_2O_2 forms oxygen gas when decomposed, ideal gas law can be used to check the H_2O_2% available in the H_2O_2 commercial product. To calculate the percent H_2O_2, a small sample of the commercial bottle H2O2 is taken to decompose through the use of a catalyst.

$$2\ H_2O_2 \xrightarrow{\text{Catalyst}} 2\ H_2O + O_2$$

Yeast, containing the enzyme catalase, will get activated in warm water. The activated yeast will then be added to a known amount of H_2O_2 in an Erlenmeyer flask. The flask is quickly sealed so that the O_2 gas formed is collected in the graduated cylinder. After measuring the total volume of gas, the temperature and the atmospheric pressure (will be given to you), the ideal gas law can then be utilized to calculate how many moles of O_2 gas is formed.

$$n = \frac{PV}{RT} \qquad \text{-----------------} \quad \text{Ideal gas law}$$

Percent yield H_2O_2 can now be found. To do this, divide (**n**), the actual number of moles you calculated, by the **theoretical moles of O_2** multiplied by 100%.

$$\%\ H_2O_2 = \frac{\text{Actual moles of O2 (n)}}{\text{Theoretical moles f O2}} \times 100$$

This value can now be compared to the% H_2O_2 shown on the label of the commercial bottle to see if any decomposition has occurred.

Pre-lab Questions

1.0 mol of H_2 gas at 22 °C and 748 Torr is trapped in an expandable/compressible container.
Use the ideal gas law to calculate the volume of the container?

Table 10.1 Chemicals and Supplies

Chemicals	Supplies	
Yeast	Flexible tubing (1)	Erlenmeyer flask (1)
3% H_2O_2	10 ml graduated cylinders (1)	Thermometer (1)
Tap H_2O	50 mL graduated cylinders (1)	Stopper with hole (1)
*Warm solution of the yeast	600 mL Beaker (1)	Burette clamp and stand (1)
		Rigid plastic tubing (1)

***Note: the yeast solution should be prepared by the TA.**

Procedure

1. Set up your apparatus as in Fig 11.1. (Insert the smaller rigid tubing into one end of the larger, flexible tubing. Insert the free end of the rigid tubing securely into the rubber stopper hole).

2. Bend the free end of the flexible tubing into a U shape, and use the rubber band to hold this shape in place. This will allow you to more easily insert this end of the flexible tubing into the inverted graduated cylinder.

   <u>Note</u>: Make sure the tubing is not pinched and the gas can flow freely through it.

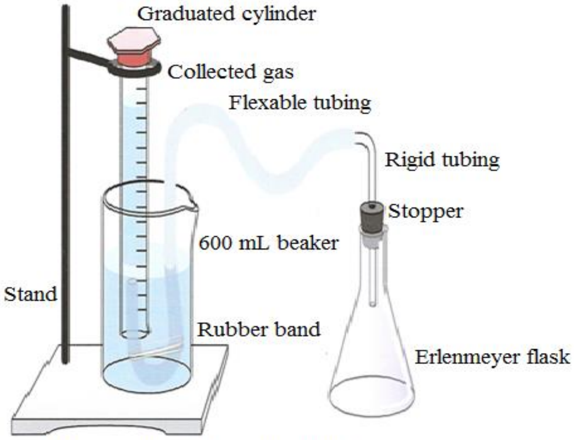


Fig 11.1 Experiment Apparatus Setup

3. Fill the 600 ml beaker with 400 ml tap water.

4. Fill the 50 ml graduated cylinder with tap water slightly over the 50 ml mark.

5. Cover the opening of the graduated cylinder with three or four of your fingers and quickly turn it upside down into the 600 ml beaker already containing the 400 ml of water. Do not remove your fingers from the opening until the graduated cylinder is fully submerged under the water. If the amount of trapped air exceeds 10 ml, refill the cylinder and try again.

6. Attach the burette clamp to the stand to secure the graduated cylinder to the burette clamp.

7. Insert the U shaped side of the flexible tubing into the beaker, and carefully snake it into the submerged opening of the graduated cylinder. You want as little air as possible to be in the graduated cylinder.

8. While the cylinder is vertical, record the volume of air inside (the line at which the water reaches in the cylinder) in table 10.2.

9. Measure the temperature of the water in the 600 ml beaker, and record it in table 10.2.

   Hint: **Assume the lab pressure is 1.00 atm**.

10. Use a 10 ml graduated cylinder to measure 3.0 ml of H_2O_2. Pour the measured H_2O_2 into the Erlenmeyer flask, and place the stopper with stopper tube over the top. Then rinse your cylinder and put it aside.

11. Using the pipette, measure out 3.0 ml of the warm yeast solution that's prepared by the TA into the rinsed 10 ml graduated cylinder.

12. Quickly pour the 3.0 ml of yeast solution into the Erlenmeyer flask. Immediately place the stopper securely in the opening of the Erlenmeyer flask by twisting it down into the flask gently.

13. Swirl the Erlenmeyer flask to mix the two solutions together.

14. You will begin to see bubbles coming up into the 50 ml graduated cylinder.

    Hint: **if gas bubbles are not immediately visible, make sure the stopper is on tight enough and the tubing is not leaking. You will need to start over after correcting any problem.**

15. Continue to swirl the Erlenmeyer flask and let the reaction run until no more bubbles form to ensure the reaction has gone to completion. This should take approximately 3-6 min.

    Hint: **Catalase enzyme works best around the temperature of the human body (37˚C). You can speed the reaction up by covering the Erlenmeyer flask with your hands.**

16. Record the time when the reaction is finished in table 10.2 of the data section, along with the final volume of air. Remember to read it at eye-level and measure from the bottom of the meniscus.

17. Pour all other liquids down the drain and take apart your apparatus when you finish your experiment

18. Clean your mess.

Report Sheet of Exp. 10: (Attach any extra paper to this sheet)

Group Names:

Part A: Fill the table below:

Table 10.2: Collecting Data

Temperature of Water	Initial volume of air (ml)	Final volume of air after reaction (ml)	Volume of O_2 collected (final volume-Initial volume)

Part B: Do the following Calculations:

1. Convert the temperature of the water from ˚C to Kelvin (K). Use the equation K= ˚C + 273.
2. Convert the volume of oxygen from ml to liter (L).
3. Rearrange the ideal gas law to solve for n. and be sure the units cancel so that you end up with only the moles of O_2 left which is going to represent the actual number of moles of O_2.
4. Calculate the theoretical number of moles of O_2. **Note**: the density of H_2O_2 is 1.02 g/mL.
5. Find the %yield of H_2O_2 using the equation below and compare it to the % H_2O_2 shown on the label of the commercial bottle to see if any decomposition has occurred

$$\% \, H_2O_2 = \frac{\text{Actual moles of O2 (n)}}{\text{Theoretical moles f O2}} \times 100$$

Experiment 11: Hydrolysis of Starch a Polysaccharide

Objective

To experience the enzyme-catalyzed and acid-catalyzed hydrolysis of polysaccharides

Introduction

Starch is a complex carbohydrate of high molecular weight having the composition (C6H10O5)n. It is known to be composed of glucose units, for this is the sole monosaccharide obtained on complete hydrolysis of the material by means of hydrochloric acid. Under the influence of certain enzymes hydrolysis proceeds just as easily as with a mineral acid but not as far. The enzyme diastase, which is present in malt, brings about hydrolysis to maltose. The saliva contains an enzyme, ptyalin, which likewise converts starch largely into maltose. This disaccharide is susceptible to hydrolysis by hydrochloric acid and hence cannot be isolated in the reaction of starch with this acid although it undoubtedly is an intermediate product. It is probable that starch is made up of a series of maltose units linked together by oxygen bridges by a process akin to a polymerization.

Starch occurs in plants in the form of granules and it is not directly soluble in water. When it is boiled with water the granules swell and the material gradually passes into solution, possibly as the result of a partial depolymerization

Even soluble starch must consist of fairly large molecules, for the solution is colloidal. This can be shown clearly by a simple dialysis experiment. If a solution containing both starch and glucose is placed in a parchment bag and this is left suspended in pure water for some hours, tests will show that glucose has diffused through the parchment and that the starch has not.

Starch is easily recognized, and at the same time distinguished from the other carbohydrates, by the characteristic blue color of some product of unknown nature which is formed by reaction with iodine. The test is very useful in following the progress of a dialysis or-a hydrolysis. In the latter case an intermediate stage is reached at which a red color is produced with iodine.

At this point the solution contains a certain amount of substances known as dextrin. These are probably tetra- or penta-saccharides, but their exact nature is unknown; a mixture of such products (" dextrin ") is used as a mucilage.

Pre-lab questions

1. How does the activity of the enzyme change as its concentration increases?
2. What are some reasons for this change in activity?
3. What’s the difference between starch and cellulose structure?

Table 11.1: Chemicals and Supplies

Chemicals	Supplies
Starch	Boiling water bath and Ice-water bath
I_2/KI solution	1 small test tube
5 mL Saliva (Volunteer from each group)	100 mL Beaker (2) and 250 mL Beaker
Conc. HCl	10 mL Graduated cylinder
	Thermometer and glass rod
	Well-plate and dropper

***** ***TA must prepare*** I_2/KI solution by dissolving 1 g of iodine in a solution of 2 g KI in 8 mL of H_2O in a beaker

Procedure

Part A: Solubility Test

1. Measure 100 mL of distilled water and transfer it to a 250 mL beaker and heat to boiling.

2. Prepare colloidal solution of starch by mixing thoroughly 1 gram of starch with 10 mL of distilled water in a 100 mL beaker until a uniform paste is obtain

3. Transfer the paste with stirring to the beaker in step 1 to make starch solution.

4. Take 1 mL from solution in step 3 and transfer it to a small test tube, cool it using the ice water bath then add 1 drop of I_2/KI solution.

   Note: step 4 is a test to make sure that there is no reducing sugars present prior to the reaction.

5. Divide the starch solution from step 3 into two parts. Cool the first part to 40°C and add 5 mL of saliva, stir it well.

6. Record the time and test for starch at 1 minute intervals by transferring a few drops (≈ 5-10 drops) of the starch solution from the first part to a small test tube and add 1 drop of I_2/KI solution.

   Note: observe any change in the appearance of the solution and record the time of noting such change, this will represent the time of a complete hydrolysis

7. To the second part of the starch solution from step 5, add 1 mL of conc. HCl and heat the solution using the boiling water bath.

   Note: observe any change in the appearance of the solution for every 5 minutes interval and record the time of noting such change, this will represent the time of a complete hydrolysis.

Report Sheet-1 of Exp.11: (Attach any extra paper to this sheet)

Group Names:

Part A: Collecting Data

Table 11.2: Hydrolysis of Starch

enzyme catalyzed hydrolysis of starch		
Part 1	**Condition of hydrolysis**	**Iodine test (positive or negative)**
	Enzymatic (saliva)	
Sample	**Cooling time (min.)**	**Iodine test (positive or negative)**
1	2	
2	4	
3	5	
Acid catalyzed hydrolysis of starch		
Part 2	**Condition of hydrolysis**	**Iodine test (positive or negative)**
	Acidic (HCl)	
Sample	**Heating time (min.)**	**Iodine test (positive or negative)**
1	5	
2	10	
3	15	

Part B: Answer the following questions:

1. The hydrolysis of starch was stopped when the iodine test no longer gave a blue color. Does this mean that the starch solution was completely hydrolyzed to glucose? Explain.

2. Which hydrolysis of the starch is faster? On the basis of this experiment estimate what will happen to the digestion of a piece of bread (containing starch) when you chew it thoroughly?

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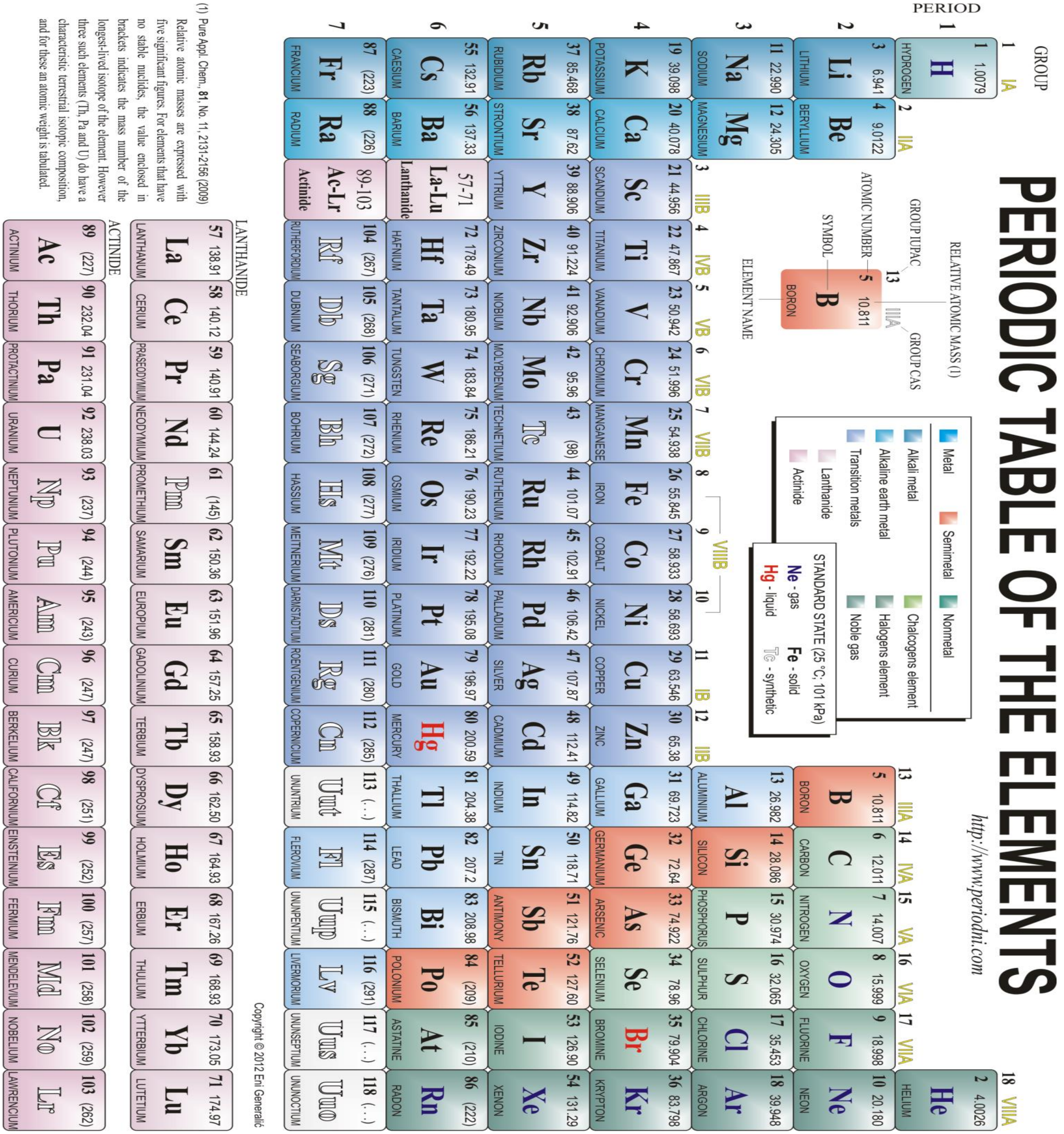
Periodic Table of the Elements



PERIODIC TABLE OF THE ELEMENTS

http://www.periodni.com

PERIOD \ GROUP	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 VIIIA
1	1 1.0079 **H** HYDROGEN																	2 4.0026 **He** HELIUM
2	3 6.941 **Li** LITHIUM	4 9.0122 **Be** BERYLLIUM											5 10.811 **B** BORON	6 12.011 **C** CARBON	7 14.007 **N** NITROGEN	8 15.999 **O** OXYGEN	9 18.998 **F** FLUORINE	10 20.180 **Ne** NEON
3	11 22.990 **Na** SODIUM	12 24.305 **Mg** MAGNESIUM											13 26.982 **Al** ALUMINIUM	14 28.086 **Si** SILICON	15 30.974 **P** PHOSPHORUS	16 32.065 **S** SULPHUR	17 35.453 **Cl** CHLORINE	18 39.948 **Ar** ARGON
4	19 39.098 **K** POTASSIUM	20 40.078 **Ca** CALCIUM	21 44.956 **Sc** SCANDIUM	22 47.867 **Ti** TITANIUM	23 50.942 **V** VANADIUM	24 51.996 **Cr** CHROMIUM	25 54.938 **Mn** MANGANESE	26 55.845 **Fe** IRON	27 58.933 **Co** COBALT	28 58.693 **Ni** NICKEL	29 63.546 **Cu** COPPER	30 65.38 **Zn** ZINC	31 69.723 **Ga** GALLIUM	32 72.64 **Ge** GERMANIUM	33 74.922 **As** ARSENIC	34 78.96 **Se** SELENIUM	35 79.904 **Br** BROMINE	36 83.798 **Kr** KRYPTON
5	37 85.468 **Rb** RUBIDIUM	38 87.62 **Sr** STRONTIUM	39 88.906 **Y** YTTRIUM	40 91.224 **Zr** ZIRCONIUM	41 92.906 **Nb** NIOBIUM	42 95.96 **Mo** MOLYBDENUM	43 (98) **Tc** TECHNETIUM	44 101.07 **Ru** RUTHENIUM	45 102.91 **Rh** RHODIUM	46 106.42 **Pd** PALLADIUM	47 107.87 **Ag** SILVER	48 112.41 **Cd** CADMIUM	49 114.82 **In** INDIUM	50 118.71 **Sn** TIN	51 121.76 **Sb** ANTIMONY	52 127.60 **Te** TELLURIUM	53 126.90 **I** IODINE	54 131.29 **Xe** XENON
6	55 132.91 **Cs** CAESIUM	56 137.33 **Ba** BARIUM	57-71 **La-Lu** Lanthanide	72 178.49 **Hf** HAFNIUM	73 180.95 **Ta** TANTALUM	74 183.84 **W** TUNGSTEN	75 186.21 **Re** RHENIUM	76 190.23 **Os** OSMIUM	77 192.22 **Ir** IRIDIUM	78 195.08 **Pt** PLATINUM	79 196.97 **Au** GOLD	80 200.59 **Hg** MERCURY	81 204.38 **Tl** THALLIUM	82 207.2 **Pb** LEAD	83 208.98 **Bi** BISMUTH	84 (209) **Po** POLONIUM	85 (210) **At** ASTATINE	86 (222) **Rn** RADON
7	87 (223) **Fr** FRANCIUM	88 (226) **Ra** RADIUM	89-103 **Ac-Lr** Actinide	104 (267) **Rf** RUTHERFORDIUM	105 (268) **Db** DUBNIUM	106 (271) **Sg** SEABORGIUM	107 (272) **Bh** BOHRIUM	108 (277) **Hs** HASSIUM	109 (276) **Mt** MEITNERIUM	110 (281) **Ds** DARMSTADTIUM	111 (280) **Rg** ROENTGENIUM	112 (285) **Cn** COPERNICIUM	113 (...) **Uut** UNUNTRIUM	114 (287) **Fl** FLEROVIUM	115 (...) **Uup** UNUNPENTIUM	116 (291) **Lv** LIVERMORIUM	117 (...) **Uus** UNUNSEPTIUM	118 (...) **Uuo** UNUNOCTIUM

LANTHANIDE

57 138.91 **La** LANTHANUM	58 140.12 **Ce** CERIUM	59 140.91 **Pr** PRASEODYMIUM	60 144.24 **Nd** NEODYMIUM	61 (145) **Pm** PROMETHIUM	62 150.36 **Sm** SAMARIUM	63 151.96 **Eu** EUROPIUM	64 157.25 **Gd** GADOLINIUM	65 158.93 **Tb** TERBIUM	66 162.50 **Dy** DYSPROSIUM	67 164.93 **Ho** HOLMIUM	68 167.26 **Er** ERBIUM	69 168.93 **Tm** THULIUM	70 173.05 **Yb** YTTERBIUM	71 174.97 **Lu** LUTETIUM

ACTINIDE

89 (227) **Ac** ACTINIUM	90 232.04 **Th** THORIUM	91 231.04 **Pa** PROTACTINIUM	92 238.03 **U** URANIUM	93 (237) **Np** NEPTUNIUM	94 (244) **Pu** PLUTONIUM	95 (243) **Am** AMERICIUM	96 (247) **Cm** CURIUM	97 (247) **Bk** BERKELIUM	98 (251) **Cf** CALIFORNIUM	99 (252) **Es** EINSTEINIUM	100 (257) **Fm** FERMIUM	101 (258) **Md** MENDELEVIUM	102 (259) **No** NOBELIUM	103 (262) **Lr** LAWRENCIUM

(1) Pure Appl. Chem., **81**, No. 11, 2131-2156 (2009)
Relative atomic masses are expressed with five significant figures. For elements that have no stable nuclides, the value enclosed in brackets indicates the mass number of the longest-lived isotope of the element. However three such elements (Th, Pa and U) do have a characteristic terrestrial isotopic composition, and for these an atomic weight is tabulated.