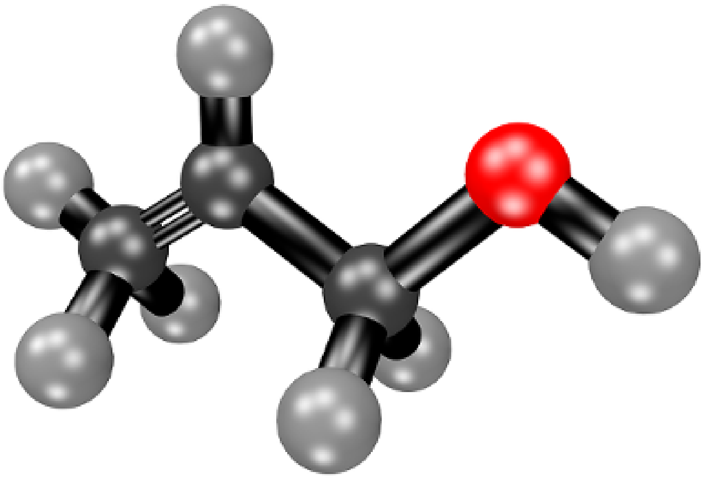
Experiments in College Chemistry

AN OPENTEXTBOOK

Version 2.7 - 2021



Science Department
B.M.C.C.
The City University of New York

THIS BOOK
NEEDS TO BE
SPIRAL BOUNDED
+
PRINTED DOUBLESIDED

1 Letter from the authors

Dear Colleagues

It is with great pleasure that we offer the alpha edition of *Experiments in College Chemistry I: an Opentextbook*. This is, to the best of the author's knowledge, the first opentextbook of experimental chemistry at a entry-college level.

Chemistry embraces everything material around us such as the clothes we wear, the air we breath or the batteries that power our cellphone, and today's chemistry is built on centuries of exploration and discovery. Most of the College Chemistry textbooks in the market contain useful information about the basic principles of chemistry thats is needed for the majority of the science and engineering student, even after graduation. Whether it's a glossary of chemical terms, or basic information about the chemistry principles, it is extremely handy to have those books around for further reference. There are also many good college chemistry laboratory manuals in the market result of the work of devoted faculty. They normally contain a series of laboratory procedures or inquiry-based experiments developed and carefully tested throughout the years. Nevertheless, these are rarely reference material as they have instead a more limited use. Students extensively use these manuals, sometimes writing their measurements in worksheets attached, what makes impossible to reuse them. But more importantly: laboratory manuals are expensive. With an average price of $50, they represent a considerable investment for a student just for single-use purchase. With all these in mind *Experiments in College Chemistry I: an Opentextbook* aims to alleviate these burdens by providing a set of standard chemistry experiments freely available for faculty and students.

The main part of the experiments in this book deal with classical chemistry techniques such as reading a meniscus, making dilutions or performing a filtration or a titration. These are standard techniques and experiments such as the acetic acid titration have been extensively validated in the literature. The reactants employed in the experiment are standard chemicals easily accessible in any general chemistry laboratory and the safety measures are minimal.

I like to think that this this opentexbook is designed to encourage students to think and to develop a solid understanding of chemistry by first building a qualitative understanding and then practicing basic chemistry techniques. Because college students often have forgotten much of their high school chemistry, each experiment begins with a background section that reviews the basic ideas behind the experiment, providing sometimes worked examples to illustrate the concepts.

We are aware of the difficulty that students have with math, as well as with plotting and graphing. With that in mind, we have included several plotting exercises in the labs that will help students practice these techniques.

Last, but by no means least, the author acknowledges the presence of many typos and mistakes and that is the reason this represents the *alpha* edition of the manual. Any critical input from the reader will be well-received.

The authors.

NYC, January 2017

Contents

2 Safety

1. Flammable materials, like alcohol, should never be dispensed or used near. Chose the most correct answer.
   (a) an open door.
   (b) an open flame.
   (c) another student.
   (d) a sink.
2. If a laboratory fire erupts, immediately
   (a) notify your instructor.
   (b) run for the fire extinguisher.
   (c) throw water on the fire.
   (d) open the windows.
3. Approved eye protection devices (such as goggles) are worn in the laboratory
   (a) to avoid eye strain.
   (b) to improve your vision.
   (c) only if you do not have corrective glasses.
   (d) any time chemicals, heat or glassware are used.
4. If you wear contact lenses in the school laboratory,
   (a) take them out before starting the lab.
   (b) you do not have to wear protective goggles.
   (c) advise your science instructor that you wear contact lenses.
   (d) keep the information to yourself.
5. After completing an experiment, all chemical waste should be
   (a) left at your lab station for the next class.
   (b) disposed of according to your instructor's directions.
   (c) dumped in the sink.
   (d) taken home.
6. You are heating a substance in a test tube. Always point the open end of the tube
   (a) toward yourself.
   (b) toward your lab partner.
   (c) toward another classmate.
   (d) away from all people.
7. Personal eyeglasses provide as much protection as
   (a) a face shield.
   (b) safety glasses.
   (c) splash proof chemical goggles.
   (d) none of the above.
8. Long hair in the laboratory must be
   (a) cut short.
   (b) held away from the experiment with one hand.
   (c) always neatly groomed.
   (d) tied back or kept entirely out of the way with a hair band, hairpins, or other confining device.
9. In a laboratory, the following should not be worn.
   (a) Loose clothing.
   (b) Dangling jewelry.
   (c) Sandals.
   (d) All of the above.
10. If an acid is splashed on your skin, wash at once with
    (a) soap.
    (b) oil.
    (c) a weak base.
    (d) plenty of water.
11. Draw a diagram of your science room and label the locations of the following:
    ☐ Fire extinguisher(s)
    ☐ Exits
    ☐ Eyewash station
    ☐ Emergency shower
    ☐ Closest fire alarm station
    ☐ Waste disposal container(s)
12. Store your books and bags:
    (a) on your lab bench.
    (b) in the aisle between lab benches.
    (c) in a cabinet in the lab.
    (d) under your bench.
    (e) on the top of the instructor's desk.
13. You are excused from wearing goggles
    (a) if you have permission from your doctor.
    (b) if they are uncomfortable to wear.
    (c) during the lecture in the lab.
    (d) if you are finished with the experiment.
    (e) None of the above.
14. If you do not wear goggles during an experiment
    (a) the instructor will suggest you to wear them.
    (b) your instructor will not care.

   (c) you are not required to wear goggles.
   (d) you will invited to leave the lab for 15 mins.
   (e) None of the above.

15. Drinks and food in the lab?

   (a) Are always allowed.
   (b) Never allowed.
   (c) Are allowed if you are hungry or thirsty.
   (d) Are allowed if you are on a diet.
   (e) None of the above.

16. If you feel ill in the lab

   (a) call your family.
   (b) ask your coworker to help you do the experiment.
   (c) tell the instructor.
   (d) leave immediately.
   (e) none of the above.

3 Class policies

1. You should arrive to the lab
   - (a) always late.
   - (b) always on time.
   - (c) within the first 20 mins.
   - (d) None of the above.
2. Thinking about missing a lab session?
   - (a) You are not allowed to miss any lab sessions.
   - (b) You are allowed to miss a single lab session.
   - (c) You can miss two labs if its justified.
   - (d) None of the above.
3. When you are done with the experiment
   - (a) you should stay until the end of the session.
   - (b) you are allowed to leave.
   - (c) you should help your coworkers.
   - (d) none of the above.
4. The lab counts for the final grade?
   - (a) Yes. It is 10% of the grade.
   - (b) Yes. It is 25% of the grade.
   - (c) Nope, it does not count for your grade.
   - (d) None of the above.
5. Regarding the lab manual,
   - (a) you should not bring it to each lab session.
   - (b) you can bring loose sheets to the lab.
   - (c) you can bring a photocopy or an e-book.
   - (d) none of the above.
6. Regarding the prelab,
   - (a) you should do it before the lab.
   - (b) you should do it during the lab.
   - (c) you can skip it.
   - (d) None of the above.
7. Regarding the lab results,
   - (a) it should be handed the day of the experiment.
   - (b) It should include two signatures from the instructor.
   - (c) It can not be handed late.
   - (d) All of the above.
8. Can I bring my own googles instead of using the ones provided by you
   - (a) True
   - (b) False
9. Where can you spiral-bind and double-side print the lab manual
   - (a) Mcdonalds
   - (b) Staples
   - (c) Fedex
   - (d) Staples and Fedex
10. This lab manual needs to be spiral-bind and double-side print because
    - (a) we need to save paper and trees
    - (b) You will be tearing out some pages
    - (c) All of the above
11. The two-signature rule says;
    - (a) The pre-lab should be signed by your instructor in the lab.
    - (b) The results of the experiment should be signed by your instructor in in the lab.
    - (c) When handing the lab results it has to include two signatures from the instructor.
    - (d) if you don't have two signatures you won't receive a grade as it means you did not do the work.
    - (e) All of the above.

4 Hazard Symbols

Hazard symbols are recognizable symbols designed to warn about hazardous materials, locations, or objects. Identify the following signs and classify them according to the labels below. Identify any non-real hazard symbol. All Images CC-BY

× Flammable, very flammable
× Toxic
× Very toxic
× Irritant
× Very irritant
× Corrosive
× Oxidizing
× Explosive
× Dangerous for the environment

A:

B:

C:

D:

E:

F:

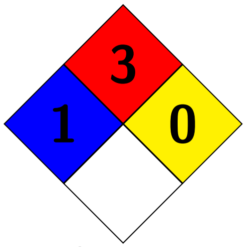
G:

H:

I:

J:

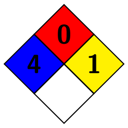
NFPA 704 is a standard system for the identification of the hazards of materials for emergency response, This is a standard maintained by the U.S.-based National Fire Protection Association. It is commonly known as "Safety Square" or "Fire Diamond" and is used by emergency personnel to quickly and easily identify the risks posed by hazardous materials.



Gasoline

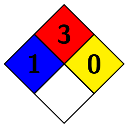
NFPA Code meaning

Rating Number	Health Hazard	Flammability Hazard	Instability Hazard	Rating Symbol	Special Hazard
4	Can be lethal	Will vaporize and readily burn at room temperature	Might explode at normal conditions	OX	Oxidizing
3	Can cause serious or permanent injury	Can be ignited under almost all ambient temperatures	Might explode at high temperatures or shock	SA	Simple asphyxiants
2	Can cause temporary incapacitation or residual injury	Must be heated or high ambient temperature to burn	Violent chemical change at high temperature or pressure	₩	Reacts violently with water
1	Can cause significant irritation	Can be preheated before ignition can occur	Normally stable hight temperatures make unstable		
0	No hazard	Will not burn	Stable		



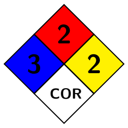
HF

This chemical will burn. True or False?



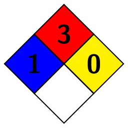
H$_2$CO$_3$

This chemical will burn your skin. True or False?



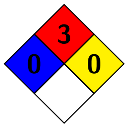
Acetic acid

This chemical will burn your skin. True or False?



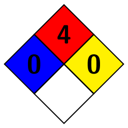
Acetone

This chemical will intoxicate you. True or False?



C2H6O

This chemical will burn. True or False?



H$_{2g}$

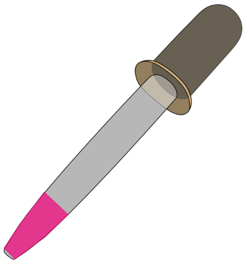
This chemical will be reactive. True or False?

5 Some Lab Material

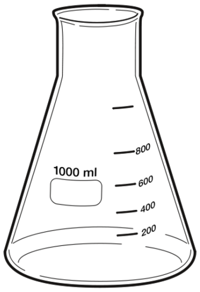
Identify the following laboratory equipment. Think about the basic use of each piece.

- Dropper
- Test tube
- Measuring cylinder
- Conical flask
- Heating plate
- buret
- Watch glass
- Beaker
- Crucible

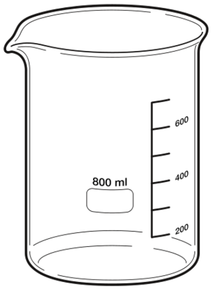
All Images CC-BY



A



B

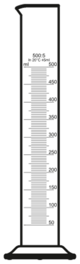


C

D

E

F



G

H

I

Search the lab and gather the following lab equipment. Show your instructor. Next, return each item to their original location:

- Test tube rack
- Crucible tongs
- Clay triangle
- Wire gauze
- Buret
- Suction Flask
- Safety googles
- Beaker

EXPERIMENT

The scientific method

Goal

The goal of this experiment is to understand how the scientific method works. In order to do this you will make observations, look for patterns and come with a hypothesis in order to *identify an unknown compound*. You will do this by comparing the reaction of the unknown compound with a set of known–reference–compounds, also called blanks.

In this experiment you will obtain some blank chemicals and an unknown. Your unknown is the same as one of the blanks. You will test each of these chemicals with a set of reagents. You have to observe the result of the test and record any change, such as *gas evolution*, the formation of a *precipitate*, a *change of color* or maybe *no reaction*. Simultaneously, you will run similar tests for the unknown sample that is identical to one of the blanks. You will determine the unknown's identity by matching its characteristic reactions with one of the blanks.

Materials

- ☐ Six test tubes.
- ☐ The blanks: NaCl(s), NaI(s), Na_2CO_3(s), Na_2HPO_4(s) and Na_2SO_4(s).
- ☐ Unknown compound.
- ☐ The reagents: HNO_3(aq), $Ba(NO_3)_2$(aq), $AgNO_3$(aq) and thymol blue.

Background

Chemistry involves the in-depth study of matter and its transformations. The *macroscopic level* refers to the level at which we can actually see the changes happening with our own eyes, without a microscope. This level deals with the properties of large objects. On the other hand, the *microscopic level* refers to a deeper level that is inaccessible to the eye.

The scientific method: how science works

Examples of 🌍 *gas evolution* and a 🌍 *precipitation reaction.*

Visual

Scientists practice the *scientific method*–a set of steps–in order to make discoveries. The first step is to collect data by *making observations and measurements*. These measurements are usually made on a sample. The next step consists on *looking for patterns and trends* in the data. When a pattern is observed, scientists *develop a hypothesis*, that is, a feasible explanation of the observations. After formulating a hypothesis, scientists think of further experiments to verify the hypothesis. If the results of repeated experiments support the hypothesis, scientists *formulate a theory* that explains the observations. One of the main goals of this experiment, adapted from the literature [1, 2], is to show you how chemists build models, turn them into a testable form, and then refine them in the light of additional evidence.

Reaction signals: gas evolution and precipitate reactions

Certain signs tell us that a chemical reaction has taken place. Oftentimes, the formation of bubbles is a sign of reaction. In general, when gases are formed from a solution, you will see a *large amount* of bubbles moving upward through the liquid.

A color change might also be a sign of a chemical reaction.

Another sign of reaction is the formation of a precipitate, such as when lead(II) iodide (PbI_2), which is an insoluble yellow solid, precipitates result of mixing lead(II) nitrate ($Pb(NO_3)_2$) and potassium iodide (KI) solutions. Precipitation, on the other hand, produces solids that can set at the bottom of the test tube, or when they are finely divided, a solid can stay in suspension, producing a milky appearance.

Chemical and physical properties

A *physical property* of a substance refers to a property that can be measured without altering the identity of the substance. Examples of physical properties are mass, temperature, melting point, hardness, color, state of matter (solid, liquid, or gas), and density. On the other hand, a *chemical property* is the ability of a chemical to transform into a different chemical. For example, certain acids in the presence of zinc change its nature producing hydrogen.

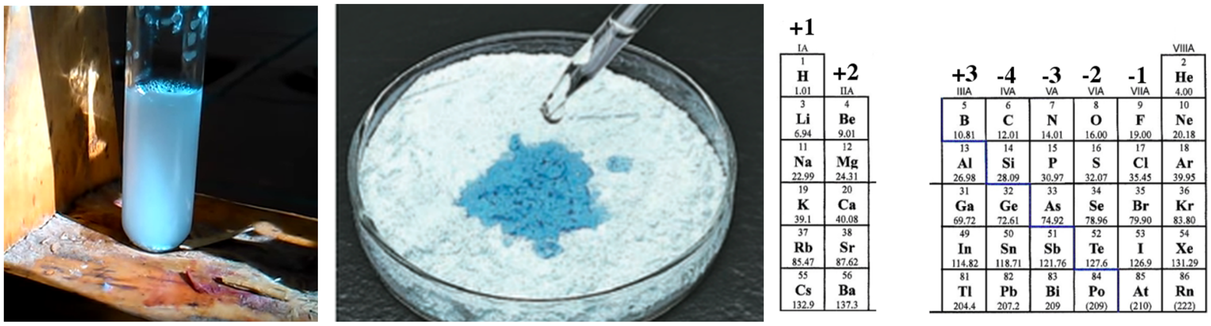


Figure 1: (Left panel) Gas evolution in a chemical reaction. (Center panel) Color change due to a chemical reaction. (Right panel) The valences on the periodic table.

Compounds

A *compound* is a substance that consists of two or more different elements. A *binary compound* consists of only two elements, for example, water (H_2O) is a binary compound made of hydrogen (H) and oxygen (O). A *ternary compound*, on the other hand, is made of three different elements. For example, nitric acid (HNO_3) is a ternary compound made of hydrogen (H), nitrogen (N) and oxygen (O).

Ionic and covalent compounds

Compounds can be classified as *covalent* or *ionic* compounds. Ionic compounds are made of a mixture of a metal with a nonmetal. For example, sodium chloride, NaCl, is an ionic compound, knowing that sodium (Na) is a metal and chlorine (Cl) is a nonmetal. On the other hand, covalent compounds are made of combinations of nonmetals. As an example, water, H_2O, is a covalent compound, because both H and O are nonmetallic elements.

Compounds and chemical valence

The elements of the periodic table combine with one another using its chemical valence. These are simple numbers that indicate the ratio of atoms in a molecule. For example, the valence of Mg is 2+ (we write this as Mg^{2+}), whereas the valence of Cl is 1− (we write this as Cl^-). When Mg^{2+} and Cl^- combine, they exchange their valences forming Mg_1Cl_2, that should be written as $MgCl_2$. Notice that the ones are always assumed and not displayed. The signs indicates which elements come first in the compound, and positive valences are arranged first whereas negative valences are arranged secondly. Sometimes, and because the subscripts are a ratio of atoms, one needs to simplify the valences and, for example, the result of combining Mg^{2+} and O^{2-} is Mg_2O_2, that is written as MgO. Some elements, typically the transition metals, can have more than one valence. In that case, the valence has to be specified in the name using roman numbers. Oftentimes, you will encounter polyatomic ions, such as ${SO_4}^{2-}$ or ${NO_3}^-$.

Example

Write down the formula resulting of combining the following pairs of ions: Ca^{2+} *with* F^-, *and* Ca^{2+} *with* ${SO_4}^{2-}$.

Answer: The resulting formulas would be: CaF_2 and $CaSO_4$.

Ionic compound naming

Ionic compounds are named by a combinations of two words, one for each ion in the formula. The first word refers to the name of the first ion, the one on the left, whereas the second word comes from the name of the second ion, the one on the right, with an *-ide* ending. For example, the name of NaCl is *sodium chloride*. Sodium is the name of the ion on the left, whereas chloride comes from the name chlorine with an *-ide* ending.

Covalent compound naming

Covalent compounds are named by a combinations of two words making use of a prefix to indicate the number of atoms in the molecule. The first word refers to the name of the first element on the left in the compound formula, whereas the second word comes from the name of the second element, the element on the right, with an *-ide* ending. You need to use prefixes (mono, di, tri, tetra,...) to indicate the number of atoms of each element. For example, the name of CO_2 is *carbon dioxide*. Carbon is the name of the element C, whereas dioxide comes from the name oxygen with the *-ide* ending. The prefix *di-* in dioxide indicates that there are two oxygen atoms in the molecule.

Acids and bases

Acids in most of the cases contain a hydrogen atom, *H*. Examples of that are HCl and H_2SO_4, both of which are acids. The later is also called an oxoacid because it contains oxygen. Oxoacid formulas are written starting with H and ending with O, and contain another element in between. Hydracids, on the other hand, do not contain oxygen. An example would be HCl which is named hydrochloric acid. The names of oxoacids and hydracids can be found in tables.
The most common bases contain *OH* in their formula. These bases are hydroxides, such as NaOH, $Ca(OH)_2$. Acids names are nonstandard and one needs to rely of naming tables. Bases on the other hand are named with the word *hydroxide*. For example, NaOH is named *sodium hydroxide*.

Oxo salts

Oxosalts are the results of the reaction of an oxoacid and a base. They normally contain three elements listed as a metal first, then a nonmetal such as N, S, or C, and ended with oxygen. $NaNO_3$ is an example of an oxosalt. HNO_3 is an oxoacid named as *nitric acid*. The ion resulting from the oxoacid, ${NO_3}^-$ is called nitrate and hence $NaNO_3$ is named as *sodium nitrate*.

Example

Name the following compounds: $MgCl_2$, P_2O_3, *HCl, KOH,* H_2CO_3, *and* Na_2CO_3.

Answers: $MgCl_2$ is named magnesium chloride, whereas P_2O_3 is diphosphorus trioxide. HCl is hydrochloric acid, whereas KOH is potassium hydroxide. H_2CO_3 is an oxoacid named carbonic acid. Its corresponding ion would be ${CO_3}^{2-}$, named carbonate. Hence Na_2CO_3 is named sodium carbonate.

Oxoacids	*Acid name*	*Anion*	*Anion name*
$HMnO_4$	Permanganic acid	${MnO_4}^-$	Permanganate
H_2CO_3	Carbonic acid	${CO_3}^{2-}$	Carbonate
HNO_3	Nitric acid	${NO_3}^-$	Nitrate
HNO_2	Nitrous acid	${NO_2}^-$	Nitrite
H_3PO_4	Phosphoric acid	${PO_4}^{3-}$	Phosphate
H_2SO_4	Sulfuric acid	${SO_4}^{2-}$	Sulfate
H_2SO_3	Sulfurous acid	${SO_3}^{2-}$	Sulfite
$H_2S_2O_2$	Thiosulfurous acid	${S_2O_2}^{2-}$	Thiosulfite
$HClO_4$	Perchloric acid	${ClO_4}^-$	Perchlorate
HIO_4	Periodic acid	${IO_4}^-$	Periodate

Hydracids	*Acid name*
HCl	Hydrochloric acid
HBr	Hydrobromic acid
HF	Hydrofluoric acid
HI	Hydroiodic acid
HCN	Hydrocyanic acid

References

(1) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

(2) Beran, J. A., *Laboratory manual for principles of general chemistry*; John Wiley and Sons: 2010.

Procedure

Indicate your hypothesis

☐ *Step 1:* – Obtain the unknown compound, the blanks and the reagents.

☐ *Step 2:* – Compare the unknown with the blanks and guess which blank looks alike the unknown. That is your hypothesis. Write down the name of your guess in the results page.

☐ *Step 3:* – Wash a 100 mL graduated cylinder and 400 mL beaker. Place a pea-size portion of the unknown in the beaker, and add 150 mL of water using the cylinder to measure it. Swirl until all the solid has totally or partially dissolved.

☐ *Step 4:* – Set aside the remaining portion of your unknown.

> **Good Lab Practice**
>
> Obtain directions from your lab instructor for discharging the solutions that you will use in this experiment. ✌

Testing for gas

☐ *Step 5:* – In this first experiment you will use the solid version of the blanks, not the droppers. You can find solids at the entrance of the lab. Obtain six test tubes. In each test tube you will add a blank reagent in solid state, that is, one blank per test tube. To each test tube you will add a pea-size sample of a different blank in solid state. You will add the unknown in the sixth test tubes. This way, place a pea-size solid sample of NaCl in the first test tube. Add a pea-size sample of NaI in the second test tube, and so on with the rest of the blanks, Na_2CO_3, Na_2HPO_4, Na_2SO_4 and your solid unknown in the following test tubes. Mark your test tubes for recognition.

☐ *Step 6:* – Add five drops of HNO_3 ⚠ to each test tube and record your observations in the results table: *GE, for gas evolution /NGE for no gas evolution.*

☐ *Step 7:* – Discharge the solutions from the test tubes into the right waste container. Wash the test tubes, and rinse them with distilled water.

Testing for a barium precipitate

☐ *Step 8:* – For the following test will use only liquid samples in the dropper, no more solids or powders. Use the solutions of NaCl, NaI, Na_2CO_3, Na_2HPO_4 and Na_2SO_4 that you will find in the laboratory, and the solution of your unknown prepared in Step 3 for the remaining tests in this practice.

☐ *Step 9:* – Using six clean test tubes, add 20 drops of each of the solutions as well as the unknown to separate test tube. In short, five of the test tubes should have a different blank and the sixth test tube should have the unknown solution. Add 3 drops of a solution of ammonia (NH_3) in each test tube ⚠ and swirl until solved.

☐ *Step 10:* – Add five drops of the $Ba(NO_3)_2$ ⚠ solution to each of the test tubes. Shake gently to obtain homogeneity. Examine each test tube carefully and look at the results. Record your observations: *P for precipitate /NP for no precipitate.*

> ⚠CAUTION!
>
> ⚠ Nitric acid (HNO_3) and ammonia (NH_3) ⚠ can cause chemical burns.
>
> ⚠ Solutions containing barium are toxic. !

Testing for silver precipitate

☐ *Step 11:* – Using six clean test tubes, add 20 drops of each of the solutions as well as the unknown to separate test tube. In short, five of the test tubes should have a different blank and the sixth test tube should have the unknown solution.

☐ *Step 12:* – Add five drops of the nitric acid $AgNO_3$ solution to each test tube. Record your observations: *P for precipitate /NP for no precipitate.*

Testing with thymol blue

☐ *Step 13:* – Using the cleaned 6 test tubes, add 20 drops of a different solution and the unknown to each test tube.

☐ *Step 14:* – Add five drops of the thymol blue solution to each test tube. Shake each test tube gently before recording your observations: *write down Blue, Yellow or Red.*

Identifying the unknown compound

☐ *Step 15:* – The unknown is one of these compounds: $NaCl$, NaI, Na_2CO_3, Na_2HPO_4, Na_2SO_4. You should be able to identify your unknown compound by matching its reactions with those of one of the known samples.

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

The scientific method

---

1. Classify the following compounds as *ionic, covalent, acids, bases* or *oxo salts*:

NH_3	______________________	SO_3	______________________
HNO_3	______________________	NaF	______________________

2. Give the names or formulas for the following compounds:

NaCl	______________________	HNO_3	______________________
Sodium iodide	______________________	Sodium carbonate	______________________

3. Name or give the formula of the following compounds:

SO_2	______________________	NO_2	______________________
Lithium fluoride	______________________	Sodium oxide	______________________

4. What does a pea size of solid mean?

5. Where in the lab you will be disposing the disposals from this experiment?

6. The following table resulted from a experiment, very similar to the one you are going to carry. The goal was to identify an unknown compound between KCl or KI. Look at the reactivity patters and identify the unknown.

		Reagents			
		HNO_3	$Ba(NO_3)_2$	$AgNO_3$	Thymol Blue
Blanks	KCl	NP	NP	P	Brown
	KI	NP	NP	NP	Brown
	Unknown	NP	NP	NP	Brown

My unknown is = ______________________

STUDENT INFO

Name: ____________________ Date: __________

Results
EXPERIMENT
The scientific method

Unknown Label # = ______________________________Your guess = ______________________________

		Reagents			
		HNO_3	$Ba(NO_3)_2$	$AgNO_3$	Thymol Blue
Blanks	NaCl				
	NaI				
	Na_2CO_3				
	Na_2HPO_4				
	Na_2SO_4				
	Unknown				

(write GE for gas evolution, NGE for no gas evolution, P for precipitate, NP for no precipitate and Blue, Yellow or Red for the color for the Thymol blue experiment)

My unknown is = ______________________________

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

The scientific method

---

1. Classify the following compounds as *ionic, covalent, acids, bases* or *oxo salts*:

$Ba(OH)_2$ ____________________

HCl ____________________

$LiCl$ ____________________

Na_2SO_4 ____________________

CO_2 ____________________

H_2SO_4 ____________________

$AuCl$ ____________________

$NaOH$ ____________________

2. Give the names or formulas for the following compounds:

HNO_3 ____________________

Barium nitrate ____________________

$AgNO_3$ ____________________

Sodium sulfate ____________________

Na_2CO_3 ____________________

Nitric acid ____________________

3. Name or give the formula of the following compounds:

$CuCl_2$ ____________________

Nickel (II) chloride ____________________

CoO ____________________

Sodium iodide ____________________

NH_3 ____________________

Hydrogen chloride ____________________

BaO ____________________

MgO ____________________

Potassium iodide ____________________

LiH ____________________

EXPERIMENT

Volumetric measurements

Goal

The goal of this experiment is to determine the *precision* of your measurements and to learn how to express correctly a *measured* result. You will learn also how to use a pipet and how to determine the precision of your measured volume. You will do this by repeatedly adding water from a filled pipet into a flask, and then measure the mass of the liquid added to the flask with a scale. Knowing the density of water, you will be able to calculate the volume delivered into the flask. Next, you will repeat the process using a graduated cylinder and compare the precision of the two methods. You will also carry simple statistics on the measured results.

Materials

- ☐ 250 mL beaker
- ☐ 10 mL volumetric pipet, 10ml graduated pipet and 10 mL cylinder
- ☐ Thermometer
- ☐ 50 mL Erlenmeyer flask with a rubber stopper
- ☐ Analytical balance
- ☐ Tab water

Background

Most chemistry experiments require the measurement of a property (mass, volume, temperature, color...), and the validity of those experiments will depend on the reliability of each measurement. The reliability of a measurement is usually considered in terms of its *accuracy* and its *precision*. The *accuracy* of a measurement is the degree of closeness of measurements of a quantity to that quantity's true, exact, value. On the other hand, the *precision* of a measurement refers to reproducibility of the experiment, that is the degree to which repeated measurements show the same results. This is a very simple and insightful experiment, implemented in many different forms in the chemistry lab literature along the years [1–4].

Significant Figures

The digits of a quantity that has been experimentally measured are not exact, and every measurement has a certain degree of uncertainty, which depends on the device used for the measurement. We should take the uncertainty in our measurements into account in order to know the error associated to them. One way to do this is to report the result of a calculation with the correct number of *significant figures* (SF). The digits different than zero are always significant. Differently, the zeros can or can not be significant. To identify the correct number of significant zero of a quantity, one needs to follow these four rules:

(a) *Rule one:* zeros at the beginning of a decimal number (these are called leading zeros) are not significant: 0.123 has 3 SF.

(b) *Rule two:* terminal zeros at the right of the decimal point (these are called trailing zeros) are significant: 0.120 has 3 SF

(c) *Rule three:* terminal zeros in a whole number (a number without an explicit decimal point) are not significant: 1230 has 3 SF.

(d) *Rule four:* for a number in scientific notation, all numbers before the power of ten are significant: 1.50×10^{-5} has 3 SF.

The number of significant figures in a measured property is determined by the tool employed on the measurement. When the data is used in subsequent calculations the uncertainty must be carried on with the results. The following two rules hold when determining the correct number of significant figures to report for the answer to a calculation:

(a) *Rule one:* when *multiplying or dividing* measured quantities, give as many significant figures in the answer as there are in the measurement with the *least number of significant figures.*

(b) *Rule two:* when *adding or subtracting* measured quantities, give the same number of decimal places in the answer as there are in the measurement with the *least number of decimal places.*

Example

Indicate the number of significant figures (SF) for the following values; (a) and (b), and perform the following calculations; (c) and (d), rounding the answers to the correct number of significant figures:

(a) 0.1250

(b) 56,700

(c) $\frac{1.235 \times 3.2}{2.34}$

(d) $1.235 + 3.2$

Answers: (a) 0.1250 has 4 SF. (b) 56,700 has 3 SF. (c) The number with the fewest significant figures is 3.2; 2 SF. Hence, the result of the calculation, 1.68889, should be rounded to two significant figures; 1.7 (d) The lowest place value in common for the 2 amounts is the tenths, hence the result of the addition, 4.435, should be rounded to one decimal place; 4.4

Average value and Standard deviation

If we measure a property repeatedly, some measurements will be closer to the true, exact, value than others. This is because there can be variations between the amount of the property being measured and in the actual measurement itself. This variability contributes to the dispersion in the measured values. The general approach is to perform a number, n, of replicated measurements under similar conditions. Obtaining several measurements allows us to calculate the sample *mean value*, $\bar{x}$, and the *standard deviation*, σ.

The sample mean $\bar{x}$ is the average value for a finite set of replicate measurements on a sample, and is calculated using the formula:

$$\bar{x} = \frac{1}{n}\sum_{i}^{n} x_i$$

where $\sum_i^n x_i$ represents the sum of all measurements, x. The sample standard deviation σ provides estimates of the population values and it is calculated using the formula:

$$\sigma = \sqrt{\frac{1}{n-1}\sum_{i}^{n}(x_i - \bar{x})^2}$$

Reporting measured results

The final value for a measurement should be written as:

$$x = \bar{x} \pm \sigma$$

notice that the standard deviation should have the same number of decimals as the average value. Read the following example to see how the calculation is done with real numbers.

Example

We measure one sample mass several times and obtain the values: 108.6 g, 104.2 g, 96.1 g, 99.6 g, and 102.2 g.

(a) Compute the average, $\bar{x}$.

(b) Compute the standard deviation of the measurement, σ.

(c) Report the measured mass in the form $\bar{x} \pm \sigma$.

Answer: (a) The average mass is given by:

$$\bar{m} = \frac{1}{5}(108.6\ g + 104.2\ g + 96.1\ g + 99.6\ g + 102.2\ g) = 102.1\ g.$$

(b) The standard deviation is given by:

$$\sigma = \sqrt{\frac{1}{4}\left((108.6\,g - 102.1\,g)^2 + (104.2\,g - 102.1\,g)^2 + (96.1\,g - 102.1\,g)^2 + (99.6\,g - 102.1\,g)^2 + (102.2\,g - 102.1\,g)^2\right)} = 4.7\,g.$$

(c) The measured mass is given by: $102.1 \pm 4.7\,g$

Use of analytical balance and volumetric material

Learn more about the correct use of an analytical balance and a pipet.

Visual

An *analytical balance* is a class of balance designed to measure small masses. The measuring pan of an analytical balance is inside a transparent container with doors in order to avoid dust and air currents. On the other hand, the volumetric measurement of liquids is a very common task in the laboratory, and volumetric instruments, such as volumetric flasks, bulb pipettes, graduated pipettes, graduated cylinders and burettes are standard laboratory equipment. On the contrary, graduated beakers, beakers, Erlenmeyer flasks, dropping funnels and the like containers *are not volumetric instruments*. They are not precisely calibrated, and the volume marks serve only as approximate guides. Transfer pipet are sensitive volumetric material calibrated to deliver a specific volume of a liquid.

Correct meniscus setting is a must in order to accurately measure a volume. In order to read the correct volume, the volumetric instrument should be held upright and the observer's eye must be at the same height as the meniscus. In this position the ring mark on the measuring tool will flatten into a straight, horizontal line. For a better reading, the meniscus will turn up darker and more easily readable in front of a light background if a piece of dark paper is held behind the instrument immediately beneath the ring mark or graduation mark.

Density

The density of an object (d) is its mass (m) per unit volume (V). You can express this as:

$$d = \frac{m}{V}$$

Density is an important physical property of a material. The value of density changes with the temperature. Water, for example, has a density of $1.000\,g/mL$ at $4.0°C$ and a density of $0.998\,g/mL$ at $20.0°C$.

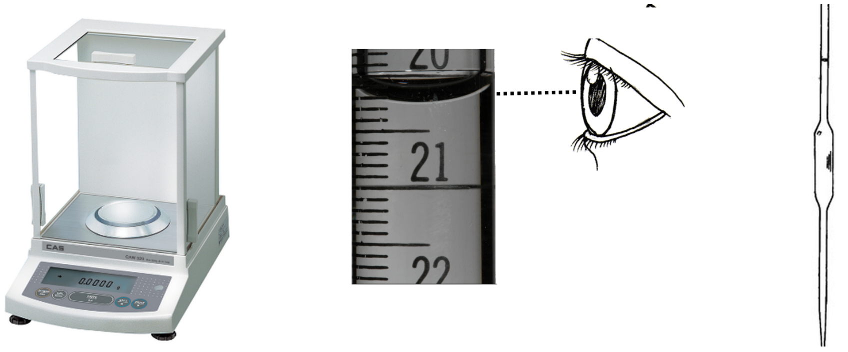


Figure 2: (Left panel) A picture of an analytical balance. (Right panel) A close up picture of a meniscus, with an indication of the graduating mark.

References

(1) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

(2) Brescia, F., *Fundamentals of Chemistry Laboratory Studies*; Elsevier: 1966.

(3) Sawyer, C. P.E.H. D., *Quantitative Analysis*; Wiley: 1948.

(4) NIST., `https://www.nist.gov/sites/default/files/documents/calibrations/74-461.pdf`, Accessed: 2017-04-22.

Procedure

Getting started

☐ *Step 1:* – Obtain 100 mL of water in a beaker, a 10 mL pipet, a thermometer and a 50 mL Erlenmeyer flask with a rubber stopper.

☐ *Step 2:* – Record the temperature of the water you have set aside with the thermometer.

☐ *Step 3:* – Find out the water density for that temperature by visiting the following website: 🌍 `http://antoine.frostburg.edu/chem/senese/javascript/water-density.html`. Copy the result with four significant figures.

☐ *Step 4:* – Locate an analytical scale in the lab. Check the bubble of the leveling eye to make the scale pan is flat. If its not, adjust the corners of the scale until the bubble enters the inner circle.

Good Lab Practice

Do not handle objects to be weighed with bare hands.
Moisture, grease and dirt on you fingers
will affect the weight of the objects.

Volume using a pipet

☐ *Step 1:* – Practice pipetting liquid several times, until you are comfortable with the technique.

☐ *Step 2:* – Bring the balance to zero and measure the mass of the dry Erlenmeyer with the stopper on. Write down the mass in the Results table.

☐ *Step 3:* – Remove the stopper and pipet 10.00 mL of the room-temperature water into the flask.

☐ *Step 4:* – Bring the balance to zero and measure and record the combined mass of the water and the stoppered flask.

☐ *Step 5:* – Remove the flask from the balance and remove the stopper. *Do not empty the flask.* Add another 10.00 mL sample into the flask, leaving the first water sample in the flask. The volume of water in the flask should now be approximately 20 mL. Put back the stopper and record the mass of the water and the stoppered flask copying all decimals provided by the scale.

☐ *Step 6:* – Keep pipetting 10.00 mL of water and recording the mass until four samples of water have been delivered to the flask and the final volume is 40.00 mL.

☐ *Step 7:* – The mass of water that was delivered each time from your pipet, is the difference between masses after each addition. Calculate the mass of water added by subtraction. These masses should be quite similar.

☐ *Step 8:* – Using the water density from your records, calculate the volume of each water sample.

Volume using a cylinder

☐ *Step 1:* – Practice reading the meniscus of a 10mL cylinder filled with water until you are comfortable with the technique.

☐ *Step 2:* – With a thermometer, record *again* the temperature of the water you have set aside, and write it down in the Results sheet.

☐ *Step 3:* – Find out the water density for that temperature by visiting the following website: 🌍 `http://antoine.frostburg.edu/chem/senese/javascript/water-density.html`

☐ *Step 4:* – Bring the balance to zero and measure and record the mass of *the same, empty, dry* Erlenmeyer with the stopper on.

☐ *Step 5:* – Using tissue paper and without touching the flask, remove the stopper and add to the flask *exactly* 10.00 mL of the room-temperature water using the cylinder.

☐ *Step 6:* – Measure and record the combined mass of the water and the stoppered flask.

☐ *Step 7:* – Remove the flask from the balance and remove the stopper. Add then another 10.00 mL of water into the flask, leaving the first water sample in the flask. The volume of water in the flask should now be approximately 20 mL. Replace the stopper and record the mass of the water and the stoppered flask.

☐ *Step 8:* – Keep adding 10.00 mL of water and recording the mass until four samples of water have been delivered to the flask and the final volume is 40.00 mL.

☐ *Step 9:* – The mass of water that was delivered each time from your cylinder, is the difference between masses after each addition. Calculate the mass of water added by subtraction. These masses should be quite similar.

☐ *Step 10:* – Using the water density from your records, calculate the volume of each water sample.

Volume using a graduated pipet

☐ *Step 1:* – With a thermometer, record *again* the temperature of the water you have set aside, and write it down in the Results sheet. Find out the water density for that temperature by visiting the following website:
🌍 `http://antoine.frostburg.edu/chem/senese/javascript/water-density.html`

☐ *Step 2:* – Bring the balance to zero and measure and record the mass of *the same, empty, dry* Erlenmeyer with the stopper on.

☐ *Step 3:* – Using tissue paper and without touching the flask, remove the stopper and add to the flask *exactly* 10.00 mL of the room-temperature water using a graduated pipet. To add 10mL you need to fill the pipet all the way to the zeroth mark and let the liquid go until the liquid level reaches the 10mL mark.

☐ *Step 4:* – Measure and record the combined mass of the water and the stoppered flask.

☐ *Step 5:* – Remove the flask from the balance and remove the stopper. Add then another 10.00 mL of water into the flask, leaving the first water sample in the flask. The volume of water in the flask should now be approximately 20 mL. Replace the stopper and record the mass of the water and the stoppered flask.

☐ *Step 6:* – Keep adding 10.00 mL of water and recording the mass until four samples of water have been delivered to the flask and the final volume is 40.00 mL.

☐ *Step 7:* – The mass of water that was delivered each time from your pipet, is the difference between masses after each addition. Calculate the mass of water added by subtraction. These masses should be quite similar.

☐ *Step 8:* – Using the water density from your records, calculate the volume of each water sample.

Calculations

⓪ Obtain the density of water at the measured temperatures using the link below:

🌍 `http://antoine.frostburg.edu/chem/senese/javascript/water-density.html`

① The mass added of water is calculated by means of:

(mass of water after the addition) – (mass of water before the addition)

② The volume of water in each addition is:

$$v_i = \frac{m_i}{d}$$

where d is the density of water at the lab temperature and m_i is the mass of water added to the beaker.

③ The average volume of the four measurements can be computed by means of the formula:

$$\bar{v} = \tfrac{1}{4}(v_1 + v_2 + v_3 + v_4)$$

④ The standard deviation in the volume measurements can be computed by means of the formula:

$$\sigma = \sqrt{\frac{1}{3}\left[(v_1 - \bar{v})^2 + (v_2 - \bar{v})^2 + (v_3 - \bar{v})^2 + (v_4 - \bar{v})^2\right]}$$

where $\bar{v}$ is the average volume calculated in the previous calculation step. You can also use the link below to calculate the standard deviation. Make sure you select the "sample" option:

🌍 `https://www.calculator.net/standard-deviation-calculator.html`

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Volumetric measurements

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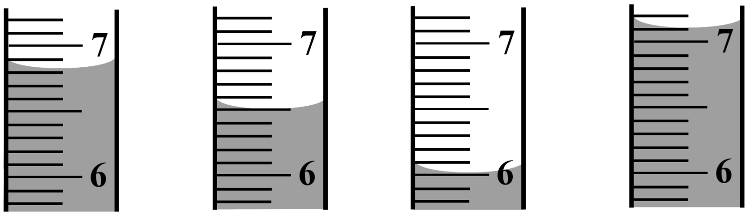
1. Throughout this experiment you will use tab water. How different would the results be if you have used distilled water instead?

2. Indicate the measurement of the following meniscus:



3. Express the answer to the following calculations to the correct number of significant figures:

(a)
$$\frac{60.875 \times 2.700}{60.4 + 107.1} =$$

(b)
$$\frac{(604.01 + 0.73) \times 321.81 \times 0.00480}{3.630 \times 10^{-3}} =$$

4. We measure the mass of an object four times and the results obtained in grams are: 2.092, 2.100, 2.053, 1.998g.
   (a) Calculate the average mass $\bar{m}$ with the right number or significant figures.

   (b) Calculate the standard deviation σ of the mass measurements.

   (c) Express the measured value of the mass in the form: $\bar{m} \pm \sigma$, and indicate how many significant figures are justified in the measured value.

5. Calculate the water density at 35.0°C by visiting the following website: `http://antoine.frostburg.edu/chem/senese/javascript/water-density.html`

6. Determine what mass of water is delivered from a 10.00 mL pipet at 39.0°C if the density of water at that temperature is 0.9918 g/mL.

STUDENT INFO

Name: ____________________ Date: __________

Results
EXPERIMENT

Volumetric measurements

Temperature (°C)=______________________________ Density of water (g/mL) ⓪ =______________________________

Using the Pipet

Addition Number	1	2	3	4
Mass before addition (g)				
Mass after addition (g)				
① Mass of added water (g)				
② Volume of water delivered (mL)				
③ Average Volume delivered, $\bar{v}$ (mL)				
④ Standard Deviation, σ (mL)				
$\bar{v} \pm \sigma$ (mL)				

(The boxes connected with an arrow should contain the same information)

Temperature (°C)=____________________ Density of water (g/mL) ⓪ =____________________

Using the Cylinder

Addition Number	1	2	3	4
Mass before addition (g)				
Mass after addition (g)				
① Mass of added water (g)				
② Volume of water delivered (mL)				
③ Average Volume delivered, $\bar{v}$ (mL)				
④ Standard Deviation, σ (mL)				
$\bar{v} \pm \sigma$ (mL)				

The boxes connected with an arrow should contain the same information)

Temperature (°C)=____________________ Density of water (g/mL) ⓪ =____________________

Using the graduated pipet

Addition Number	1	2	3	4
Mass before addition (g)				
Mass after addition (g)				
① Mass of added water (g)				
② Volume of water delivered (mL)				
③ Average Volume delivered, $\bar{v}$ (mL)				
④ Standard Deviation, σ (mL)				
$\bar{v} \pm \sigma$ (mL)				

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Volumetric measurements

---

1. The results of measuring the mass of an *stopped Erlenmeyer flask* mass are: 50.123g, 50.013g, 49.933g and 51.002g,
   (a) Calculate the standard deviation in your measured mass for the empty stopped flask.

   (b) How does the result compare with the claimed precision of your balance? The balance precision is ± 0.0001 g.

   (c) Express the measured value of the mass in the form: $\bar{m} \pm \sigma$.

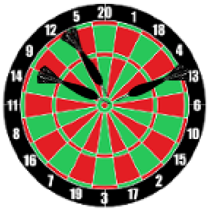
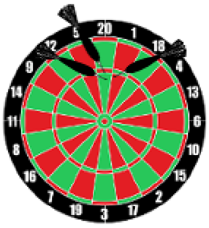
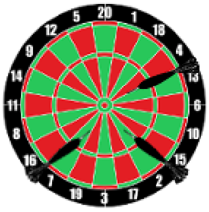
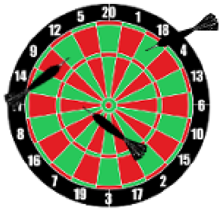
2. For the measurement of the *volume of water in a pipet:* compare the precision of the pipet with your result. The pipet precision is ± 0.01 mL.

3. Compare the accuracy of the volume measurements when using a pipet with the one when using a cylinder.

4. Compare the precision of the volume measurements when using a pipet with the one when using a cylinder.

5. Using the data from:
   `http://antoine.frostburg.edu/chem/senese/javascript/water-density.html`
   plot the values of density vs. temperature for water: (use a range between 0.0°C and 90.0°C)

6. The darts in these targets represent measurements that are precise and accurate (PaA), precise but inaccurate (PbI), imprecise but accurate on average (IbA), and both imprecise and inaccurate (IaI). Correlate the targets with the most suitable description.

EXPERIMENT

Mass spectrometry

Goal

The goal of this experiment is to determine the *isotopic composition* and *atomic masses* of several elements. You will achieve this goal by studying mass spectra. By measuring the height of the mass spectrum peaks–as the height of a signal is proportional to the fractional abundance–you will be able to calculate the fractional abundances of the different isotopes.

Materials

This is a modeling-based experiment, and you will not need a real mass spectrometer to carry out the experiment. You will only need a ruler and the mass spectra provided here, and extracted from *www.webelements.com* and the *National Institute of Standards and technology*, a US measurement standards laboratory.

Background

A mass spectrometer is an instrument used to determine atomic and molecular masses. This instrument can also be used to determine the masses and isotopic composition of an element. In this theory-based experiment, adapted from the literature [1], you will first analyze a set of mass spectra. Then, from their peaks, you will calculate the isotopic composition of a series of elements.

The nuclear model

According to the current nuclear model of the *atom*, atoms contain a small positively charged nucleus, which is responsible for almost all atomic mass, and a negatively charged electronic cloud. When the positive charge of the nucleus exactly cancels the negative charge of the surrounding electrons, we say that atoms are electrically neutral. The positively charged particles inside the nucleus are called *protons* (denoted p), whereas *electrons* (denoted e^-) have a single negative charge. The number of protons of an element is called the *atomic number*, Z, of that element. For neutral atoms, the number of protons equals the number of electrons. As an example, the nucleus of a hydrogen atom (H) has one proton, so its atomic number is Z=1, whereas the nucleus of a helium atom (He) has two protons, so its atomic number is Z=2. The atomic nucleus also contains electrically neutral particles called *neutrons* (denoted n). Neutrons have approximately the same mass as protons, and so they add substantially to the mass of the nucleus. Therefore, differences in the number of neutrons in a nucleus give rise to different atomic masses, even though those atoms are of the same element. Apart from their charge, neutrons and protons are jointly known as *nucleons*.

Isotopes and atomic mass

Learn more about *isotopes*.

Visual

The total number of protons and neutrons in a nucleus is called the *mass number*, A, of the atom. All the atoms of a particular element have the same atomic number. However, different atoms of the same element can have different mass numbers. *Isotopes* are atoms of the same element that have the same atomic number but a different mass number. Different isotopes have different masses and different nuclear properties. Isotopes from the same element are differentiated by writing their mass number after the name of the element, as in neon-20, neon-21, and neon-22. Its symbol such as $^{A}_{Z}\mathrm{X}$

is obtained by writing the mass number A as a superscript to the left of the chemical symbol of the element and the atomic number Z included as a subscript on the lower left, as in ${}^{20}_{10}\text{Ne}$. Some isotopes of an atom are more common than others and this is quantified using the isotope *percent abundance*, F. For example, hydrogen has three isotopes. The most common is ${}^{1}_{1}\text{H}$ called protium with a 99.985% abundance. The second isotope, deuterium ${}^{2}_{1}\text{H}$, is less common, with 0.015% abundance. The third one called tritium is a very rare isotope. The abundance can also be expressed as a *fractional abundance* f and for example the fractional abundance of ${}^{1}_{1}\text{H}$ is 0.99985.

The *atomic mass* of an element, expressed in amu (atomic mass units), is the weighted average of the masses of the individual isotopes of the element. For an element with n isotopes with different masses (A_1, A_2, ..., A_n) and different fractional abundances for each isotope (f_1, f_2, ..,f_n), the atomic mass is given by

$$\text{Atomic mass} = \sum_{i=1}^{n} A_i \cdot f_i$$

Example

Naturally occurring copper (Cu) consists of 69.17% ^{63}*Cu and* 30.83% ^{65}*Cu. The mass of* ^{63}*Cu is 62.939598 amu, and the mass of* ^{65}*Cu is 64.927793 amu. What is the atomic mass of copper?*

Answer: The weighted average is the sum of the mass of each isotope times its fractional abundance:

$$62.939598\, amu \times \frac{69.17}{100.} + 64.927793\, amu \times \frac{30.83}{100.} = 63.55\, amu$$

Mass spectrometry

Mass spectrometry is a very powerful analytical technique that provides both quantitative and qualitative information about the chemical being analyzed. In a mass spectrometer, the sample is first vaporized to generate atoms, and then introduced into the sample inlet chamber that leads to an ionization chamber. In this chamber, electrons collide with the vaporized species producing ions and new electrons. The resulting ions are accelerated by an electric field into a beam that is deflected toward a detector by a magnetic field. The heavier the particle, the lower the bean deflection. Hence, the degree of deflection is related to the relative masses of the deflected particles.

Learn more about how a 🌍 *mass spectrometer* works.

Visual

Mass spectrum

The positions of the peaks in the mass spectrum give the mass-to-charge ratios of the ions. If the sample consists of atoms of one element, the mass spectrum gives the isotopic distribution of the sample. Mass spectra of molecules are more complex as bond breaking can occur. Thus the mass spectrum of a molecule contains signals that are due to fragments of molecules in the mass spectrum, as well as those that are due to the unfragmented molecule. The mass of this molecular ion, the parent ion, is called the parent mass. The parent ion has essentially the same molar mass as the compound itself.

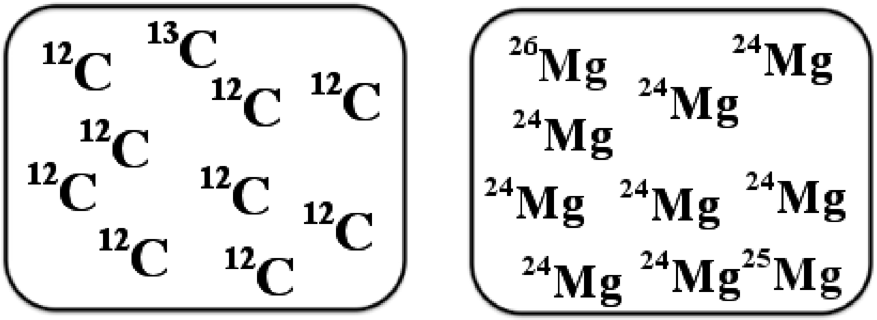
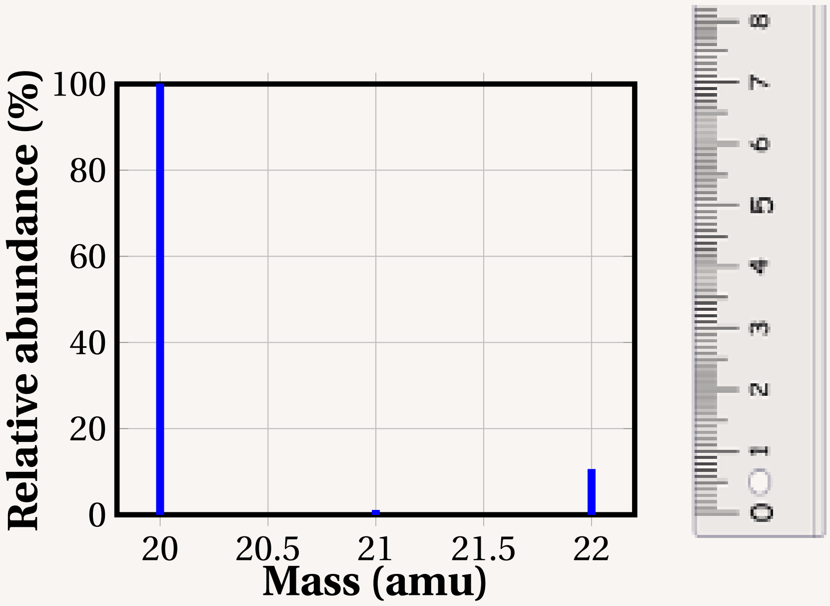


Figure 5: Image representing the atomic composition of two different elements

Example

Calculate the % abundance of the naturally occurring isotopes of Ne from the lines of the spectrum below.



Answer: the mass spectrum of He consists of an intense signal at mass number 20, a much smaller signal at a mass of 22, and a very weak signal at 21 amu. The horizontal coordinates of each peak represent the mass number of each isotope (A_i), as listed in the table below. The heights of these peaks are proportional to the number of counts at each mass number and, in return, reflect the fractional abundances of the isotopes. By using a ruler we can measure the length of the peaks, as we list them in the table below. By adding the length of each peak we obtain I_{total}, also listed on the table. We obtain the fractional abundance of each isotope f_i by dividing the intensity of each peak by the total intensity, also listed in the table below. The fractional abundance is related to the isotope abundance. For example, the fractional abundance of $^{22}_{9}$Ne is 0.903, hence, the abundance of this isotope is 90.3%.

Isotope ($^{A_i}_{Z}$X)	Mass Number (A_i)	Intensity (I_i)	Fractional Abundance (f_i)	Isotope contribution ($f_i \cdot A_i$)
^{20}Ne	20	7.00 cm	$\frac{7.00}{7.75} = 0.903$	$0.903 \cdot 20 = 18$ amu
^{21}Ne	21	0.05 cm	$\frac{0.05}{7.75} = 0.006$	$0.006 \cdot 21 = 0.1$ amu
^{22}Ne	22	0.70 cm	$\frac{0.70}{7.75} = 0.090$	$0.090 \cdot 22 = 2.0$ amu
	$I_{total} = 7.00 + 0.70 + 0.05 = 7.75$ cm		AW$_{experimental}$ =	$18 + 0.1 + 2.0 = 20.1$ amu
			AW$_{Theoretical}$ =	20.2 amu
			% Error AW =	4.9%

Finally, we obtain the isotope contribution to the atomic mass of the element by multiplying the fractional abundance by the mass number of each isotope. By adding all isotope contributions we obtain the average atomic mass, which can be compared from the value found on the periodic table. The error between the calculated and experimental values is less than 5%.

References

(1) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

Calculations

The mass spectrum of a series of atoms is provided next page. Select each of the spectra, while following the next steps:

⓪ Pick up one of the mass spectra. Write down the name of the element.

① Write down the name of the isotope. Z is the atomic mass from the periodic table. A is the mass number of the pick. X is the name of the element.

② Write down the mass number of each pick (horizontal reading of the mass spectra).

③ With a ruler measure the length of each pick. Consider the precision of your ruler and use the last decimal possible (estimated value).

④ Add the intensity of each pick.

⑤ Calculate the fractional abundance by doing: ③/④

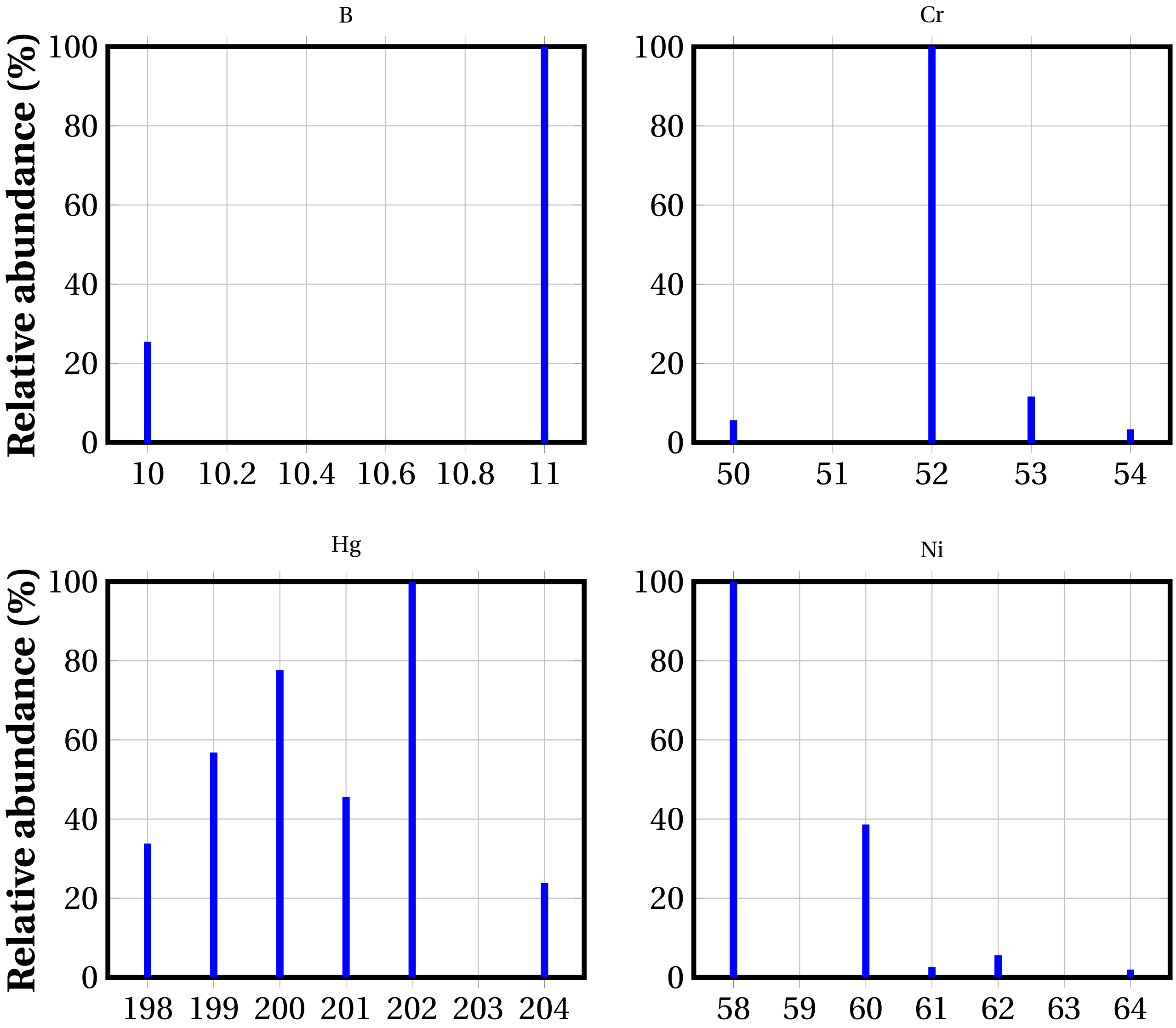
⑥ Calculate the isotope abundance: ⑤ × ②

⑦ Calculate the average atomic mass by adding all isotope contributions.

⑧ Obtain the experimental atomic mass of the element from the periodic table

⑨ Calculate the % Error AW as indicate below. The value should be positive. If the error is lower than 10% you need to review the accuracy of your measurements and redo the numbers:

% Error AW=| ⑦ − ⑧ | / ⑧ × 100%


B
Cr
Hg
Ni
Relative abundance (%)
Relative abundance (%)
0
20
40
60
80
100
10
10.2
10.4
10.6
10.8
11
50
51
52
53
54
198
199
200
201
202
203
204
58
59
60
61
62
63
64

STUDENT INFO

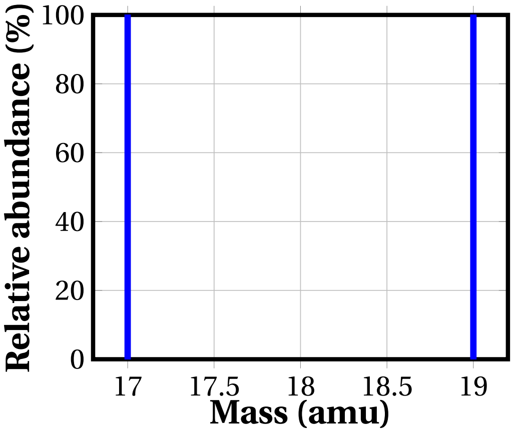
Name: ____________________ Date: __________

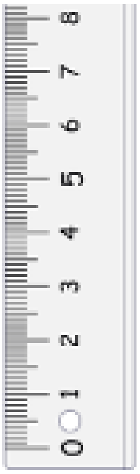
Pre-lab Questions

Mass spectrometry

---

1. Consider the mass spectrum below of an element X: how many isotopes are present and what is the fractional abundance of the isotopes? Given that the average atomic mass of the element from the periodic table is 17.898 amu, use the table below to indicate the experimental atomic weight and the percent error.





Isotope ($^{A_i}_{Z}X$)	Mass Number (A_i)	Intensity (I_i)	Fractional Abundance (f_i)	Isotope contribution ($f_i \cdot A_i$)
	I_{total} =		$AW_{experimental}$	
			$AW_{Theoretical}$ =	
			% Error AW =	

STUDENT INFO

Name: ____________________ Date: __________

Results

EXPERIMENT

Mass spectrometry

<table>
<tr><td colspan="5">⓪ Element=</td></tr>
<tr><td>Isotope $({}^{A_i}_{Z}X)$ ①</td><td>Mass Number (A_i) ②</td><td>Intensity (I_i) ③</td><td>Fractional Abundance (f_i) ⑤</td><td>Isotope contribution $(f_i \cdot A_i)$ ⑥</td></tr>
<tr><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td></td><td></td><td></td><td></td><td></td></tr>
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<tr><td></td><td></td><td></td><td></td><td></td></tr>
<tr><td></td><td>I_{total} = ④</td><td></td><td>$AW_{experimental}$ ⑦</td><td></td></tr>
<tr><td></td><td></td><td></td><td>$AW_{Theoretical}$ = ⑧</td><td></td></tr>
<tr><td></td><td></td><td></td><td>% Error AW = ⑨</td><td></td></tr>
</table>

⓪ Element=				
Isotope ($^{A_i}_{Z}X$) ①	Mass Number (A_i) ②	Intensity (I_i) ③	Fractional Abundance (f_i) ⑤	Isotope contribution ($f_i \cdot A_i$) ⑥
	I_{total} = ④		$AW_{experimental}$ ⑦	
			$AW_{Theoretical}$ = ⑧	
			% Error AW = ⑨	

⓪ Element=				
Isotope ($^{A_i}_{Z}X$) ①	Mass Number (A_i) ②	Intensity (I_i) ③	Fractional Abundance (f_i) ⑤	Isotope contribution ($f_i \cdot A_i$) ⑥
	I_{total} = ④		$AW_{experimental}$ ⑦	
			$AW_{Theoretical}$ = ⑧	
			% Error AW = ⑨	

⓪ Element=				
Isotope $({}^{A_i}_{Z}X)$ ①	Mass Number (A_i) ②	Intensity (I_i) ③	Fractional Abundance (f_i) ⑤	Isotope contribution $(f_i \cdot A_i)$ ⑥
	I_{total} = ④		$AW_{experimental}$ ⑦	
			$AW_{Theoretical}$ = ⑧	
			% Error AW = ⑨	

STUDENT INFO

Name: ______________ Date: __________

Post-lab Questions

Mass spectrometry

---

1. The atomic mass of Ga is 69.72 amu. There are only two naturally occurring isotopes of gallium: 69Ga, with a mass of 69.0 amu, and 71Ga, with a mass of 71.0 amu. Calculate the natural abundance of the 69Ga isotope.

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2. Magnesium contains three different isotopes: magnesium-24 with an abundance of 79% and a mass of 23.9850423 amu, magnesium-25 with an abundance of 10% and a mass of 24.9858374 amu, and magnesium-26 with a mass of 25.9825937 amu. Calculate the average atomic mass of a sample of magnesium.

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EXPERIMENT

Empirical Formula of an Oxide

Goal

The goal of this experiment is to determine the *empirical formula* and *molecular formula* of magnesium oxide, produced by burning a known amount of Mg with oxygen from the air. Unfortunately, the oxidation of Mg to give MgO in air also produce a small amount of magnesium nitride (Mg_2N_3), think that there is N_2 in the air. Adding water will get rid of the nitride by liberating ammonia (NH_3). You will determine the mass of oxygen that is present in the oxide by subtracting the mass of the product from the original mass of magnesium employed.

Materials

☐ 0.2 g of Mg ribbon
☐ Metallic wool
☐ Crucible with lid, ring stand with clay triangle
☐ Bunsen burner
☐ Crucible tongs and a wire gauze

Background

Scientists often gain insight into the nature of chemicals by studying the properties of the compounds found in nature. Once they have extracted a compound from a natural resource, they try to identify its formula. In this experiment, we will identify the molecular formula of a chemical compound resulting of the reaction between Mg and O_2. This experiment is a well-known classic in college chemistry, implemented previously in numerous laboratory manuals [1–4].

Chemical formulas

The *empirical formula* of a compound shows the relative numbers of atoms of each element present in that compound. For example, the empirical formula of glucose, which is CH_2O, tells us that carbon, hydrogen, and oxygen atoms are present in the ratio 1:2:1. Differently, the *molecular formula* shows the actual numbers of atoms of each element in a molecule. The molecular formula for glucose, which is $C_6H_{12}O_6$, tells us that each glucose molecule consists of 6 carbon atoms, 12 hydrogen atoms, and 6 oxygen atoms.
To obtain the empirical formula, first you need to convert the mass percentage of each type of atom in the compound into the relative number of atoms of that element. The simplest procedure is to imagine that we have a sample with a mass of exactly 100. g. In this case the mass percentage composition gives the mass in grams of each element. Then we use the molar mass of each element to convert these masses into amounts in moles and go on to find the relative numbers of moles of each type of atom.

Example

The mass percentage composition of a compound is: 18.59% O*,* 37.25% S*, and* 44.16% F*. Calculate its empirical formula.*

Answer: the mass m of each atom in 100. g of the compound is:

$$m(O) = 18.59\,g; \qquad m(S) = 37.25\,g; \qquad m(F) = 44.16\,g.$$

We convert the mass of each element into moles n:

$$n(O) = 1.16\,mol; \qquad n(S) = 1.16\,mol; \qquad n(F) = 2.32\,mol.$$

We now divide each amount by the smallest amount (1.16):

$$O = 1; \qquad S = 1; \qquad F = 2.$$

The empirical formula is: OSF_2

Determining molecular formulas

The empirical formula of glucose is CH_2O. However, the empirical formula tells us only that the C, H, and O atoms are present in the sample in the ratio 1:2:1, not the number of each type of atom in a molecule. The molecular formula could be $C_2H_4O_2$ or $C_4H_8O_4$, or any other whole-number multiple of the empirical formula. To find the molecular formula of a compound, we need its molar mass.

Example

The molar mass of a compound is $104\,g\cdot mol^{-1}$*, and its empirical formula is CH. Deduce the molecular formula of the compound.*

Answer: the molecular formula of the molecule is $n\cdot$CH, where n is an integer. Knowing that the molecular mass if $104\,g\cdot mol^{-1}$ and the molecular mass of CH is $13\,g\cdot mol^{-1}$, we can solve for n:

$$n\cdot 13 = 104$$

and $n = 8$. Hence the molecular formula is: C_8H_8.

The oxidation of Magnesium

Air is a mixture of different gasses such as oxygen, nitrogen, water or carbon dioxide. Oxygen (O_2) in the air, which is a diatomic molecule, is a very reactive substance and many elements will react with it forming what we call oxides. When Mg reacts with oxygen, magnesium oxide is formed:

$$2\,Mg \quad + \quad O_2 \quad \longrightarrow \quad 2\,MgO$$

Nitrogen in the air N_2, also a diatomic molecule is normally inert, which means that it will hardly react. Still at very high temperature metals can react with nitrogen producing what we called nitrides:

$$3\,Mg \quad + \quad N_2 \quad \longrightarrow \quad Mg_3N_2$$

By means of water and working still at high temperatures, we can eliminate Mg_3N_2 producing ammonia:

$$Mg_3N_2(s) \quad + \quad 6\,H_2O(l) \quad \longrightarrow \quad 3\,Mg(OH)_2(aq) \quad + \quad 2\,NH_3(g).$$

Figure 7: (Left panel) A crucible. (Central panel) Use of the crucible tongs. (Right panel) Correct arrangement of the ring stand.

References

(1) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

(2) Russo, T., *Merrill Chemistry Lab Manual*; McGraw-Hill: 1998.

(3) Fetzer-Gislason, P. R.D.W. S., *Lab Experiments in Introductory Chemistry*; Freeman: 2003.

(4) Bull, W. B. E., *Laboratory Manual for College Chemistry*; SHarper and Row: 1980.

Procedure

Getting started

☐ *Step 1:* – Obtain a crucible with a lid, a clay triangle and an iron ring attached to a ring stand.

☐ *Step 2:* – Place the covered crucible in the clay triangle on an iron ring attached to a ring stand. Adjust the height of the ring so that the bottom of the crucible will be in the hottest part of the flame. The correct arrangement of the equipment, crucible, and burner is shown in the figure (Right panel).

☐ *Step 3:* – Learn how to use the Bunsen burner ✍. Heat the covered crucible in the hottest part of the flame for about 5 min while keeping the lid ajar, making sure that the bottom of the crucible attains a red glow.

☐ *Step 4:* – Stop the burner and allow the crucible to cool down completely ⚠.

☐ *Step 5:* – Weight the covered crucible and record the mass of the covered crucible. Now you know the mass of the empty– and clean–crucible with lid.

Good Lab Practice

✍ Discharge the magnesium oxide in the disposal container.

✍ In the following video you can learn the correct use of a 🌍 *Bunsen burner*.

Burning Magnesium

☐ *Step 6:* – Obtain 0.2 g of magnesium ribbon. Clean the surface of the metal with metallic wool until it shines.

☐ *Step 7:* – Cut the magnesium ribbon into tiny bits, and place them inside the crucible.

☐ *Step 8:* – Cover the crucible, obtain and record the mass again. Now you know the mass of the crucible+lid+Mg.

☐ *Step 9:* – Set the crucible on the clay triangle with the lid on and heat the crucible ✍ in the hottest part of the flame another 5 min ⚠. Keep the lid close.

☐ *Step 10:* – Using the crucible tongs, lift the lid carefully by a slight amount. The metal should glow brightly without flames. Continue until all Mg is burned and the product does not glow.

☐ *Step 11:* – Patiently cool down the crucible with lid. The content should be white or slightly gray. At this point, add a few drops of water using a plastic pipet on the crucible content. You might notice a smell of ammonia at this point.

☐ *Step 12:* – Place the lid back onto the crucible, slightly ajar, and heat the crucible ⚠in the hottest part of the flame for 15 more minutes.

☐ *Step 13:* – After that time, allow the covered crucible and its content to cool down. Obtain the mass of the covered crucible.

⚠CAUTION!

⚠ Do not burn yourself with the crucible: it is hot.

⚠ Make sure the gas outlet is closed before leaving the lab.

Calculations

① This is the mass of the empty and clean crucible with lid.

② This is the mass of the clean crucible with lid and the Mg.

③ This is the mass of Mg added to the crucible: ② – ①

④ This is the moles of Mg (Atomic weight 24.305 $g \cdot mol^{-1}$):

$$n_{Mg} = \frac{\textcircled{3}\, g}{24.305\, g \cdot mol^{-1}}$$

⑤ This is the mass of the clean crucible with lid and the product.

⑥ This is the mass of product: ⑤ – ①

⑦ This is the mass of O in the product: ⑥ – ③

⑧ This is the moles of O (Atomic weight 15.999 $g \cdot mol^{-1}$) in the product: $n_O = \frac{\textcircled{7}\, g}{15.999\, g \cdot mol^{-1}}$

STUDENT INFO

Name: ____________________ Date: ________

Pre-lab Questions

Empirical Formula of an Oxide

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1. Caffeine has a molar mass of 194.19 $g \cdot mol^{-1}$ and a mass percentage composition of 49.48% C, 5.19% H, 28.85% N, and 16.48% O. What is the molecular formula of caffeine?.

2. The percent composition of chloroform is 10.06% carbon, 0.85% hydrogen, and 89.09% chlorine. Determine the empirical formula.

   of salicylic acid knowing that the mass percentage composition is: 60.87% C, 4.38% H, and 34.75% O.

3. In an experiment, 4.14 g of Ni are burned to produce 4.88 g of an oxide. Calculate the empirical formula of the compound? Mind that the indexes of a formula are always integer numbers.

STUDENT INFO

Name: ______________________ Date: ________

Results

EXPERIMENT

Empirical Formula of an Oxide

---

①	Mass of empty crucible and lid (g)	
②	Mass of crucible and lid with Mg (g)	
③	Mass of Mg (g)	
④	Moles of Mg (mole)	
⑤	Mass of crucible and lid with MgO (g)	
⑥	Mass of MgO (g)	
⑦	Mass of O (g)	
⑧	Moles of O (moles)	

	Mg	O
Moles of (moles)		
Moles/smallest amount		
Empirical Formula		

STUDENT INFO

Name: ____________ Date: ________

Post-lab Questions

Empirical Formula of an Oxide

---

1. Write correctly balanced chemical equations for the following reactions:
   (a) Magnesium with molecular oxygen to produce magnesium oxide

   (b) Magnesium with molecular nitrogen to produce magnesium nitride

   (c) Magnesium nitride with water to produce magnesium hydroxide and ammonia

2. A 1.27-g sample of metallic copper was placed in a 26.32-g crucible and heated until all the copper had reacted with the oxygen in air to form an oxide. The crucible and product together were found to weigh 27.75 g. Calculate the empirical formula of the oxide and give the name.

3. Explain why did you have to shine the magnesium ribbon before the reaction.

10 Spectrochemistry

Fpost

EXPERIMENT

Spectrochemistry

Goal

The goal of this experiment is to calculate the *concentration* of an unknown Co^{2+} solution by means of spectrochemistry and to get experience on the measurement of absorbance. To get started, you will need to find the *correct wavelength* of light to use in the measurements. This is the wavelength at which the maximum absorbance occurs. You will also practice dilutions, and you will *dilute* a concentrated cobalt solution to then measure the absorbance of the diluted solution. Finally, you will calculate backwards the concentration in the original solution. The absorbance of a known cobalt solution will be measured before and after the dilutions.

Materials

- ☐ a $CoCl_2$ solution of known concentration
- ☐ 8 test tubes
- ☐ a spectrophotometer
- ☐ an unknown
- ☐ a 50 mL beaker
- ☐ a Mohr pipet

Background

When atoms are heated they emit light with very characteristic frequencies. At the same time, matter can also absorb discrete frequencies of light. The process of light absorption is exactly the opposite of light emission. During the absorption of light, an electron undergoes a transition from a lower-energy level to a higher-energy level, and the difference in energy corresponds to the energy of the light absorbed. As a result of the transition, a component of light–remember white light have components of different frequencies–with frequency ν is absorbed and the other frequencies are transmitted across the material. This experiment is truly a classic in college chemistry, historically implemented in numerous lab manuals [1–4].

Learn more about light and chemistry in this Video Link

Visual

Transmittance and Absorbance

When light of a specific wavelength illuminates a solution of an absorbent substance, the intensity of the light diminishes as it passes through the solution. This is due to light absorption. If I_o is the intensity of the light that enters and I the intensity of the light that emerges from the solution, the transmittance is defined as:

$$T = \frac{I}{I_o}$$

A related quantity called absorbance, A, is defined as the negative logarithm of the transmittance

$$A = -log(T)$$

Spectrophotometers

Learn more about how spectrophotometers work
🌍 Video Link

Visual

Absorbance is measured with an instrument called a spectrophotometer. This instrument separates light into its component wavelengths and selectively measures the intensity of light of a given wavelength after it goes through a solution. All spectrophotometers have certain parts in common such as the source of light, a prism to select a particular wavelength, a device for holding the sample and a detector that measures the intensity of the light.

Beer's Law

Beer's law states that the absorption A is directly related to the concentration c of a substance that absorbs light:

$$A = k_\lambda \cdot c$$

where k_λ is a constant. Because A is a dimensionless number and the unit of measurement for c is mol/L or M, it follows that the unit of measurement for k_λ is L/mol or M^{-1}. k is only a constant for a given substance at a particular wavelength, hence the subscript λ.

The determination of k_λ

The equation for Beer's law, $A = k \cdot c$, has the same form as the equation for a straight line, $y = mx + b$, where A is y, c is x and k_λ is the slope, m. Consequently, you should obtain a straight line when you plot the absorbances obtained at various concentrations against those concentrations. Moreover, the slope of that line will be given by k_λ and the line must pass through the origin because at zero concentration, zero absorbance.

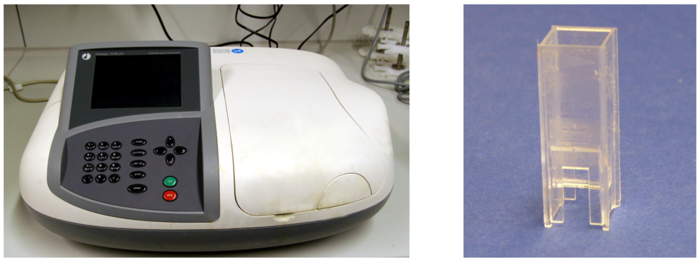

Figure 8: (Left panel) A spectrophotometer. (Right panel) A cuvette.

Dilutions

Sometimes, in chemistry labs one needs to obtain from a concentrated solution a more diluted one. To dilute a stock–more concentrated–solution to a desired–more diluted–concentration, a pipet is used in order to transfer the appropriate volume of stock solution to a volumetric flask with a given volume. Then solvent is added to increase the volume of the solution to the volume of the flask. When diluting a certain volume V_1 of stock solution with concentration c_1 into a larger volume V_2, the formula that relates the stock solution concentration with the diluted solution c_2 is:

$$c_1 \cdot V_1 = c_2 \cdot V_2$$

Example

You need to prepare 250. mL of a diluted 1.25×10^{-3} *M NaOH(aq) solution using a concentrated 0.0380 M NaOH(aq) stock solution. How much stock solution do you need?*

Answer: The original, stock solution has a concentration of $c_1 = 0.0380\,M$, whereas for the diluted solution should have a concentration $c_2 = 1.25 \times 10^{-3} M$. We need a volume $V_2 = 250.\,mL$. We solve for V_1 to know how much of the stock solution we need to take and later dilute with the solvent.

$$c_1 \cdot V_1 = c_2 \cdot V_2 \qquad V_1 = \frac{c_2 \cdot V_2}{c_1}$$

$$V_1 = \frac{1.25 \times 10^{-3} M \cdot 250.\,mL}{0.0380 M}$$

and this gives $V_1 = 8.22\,mL$.

References

(1) Murov, S., *Experiments in General Chemistry*; Cengage: 2013.

(2) Sawyer, R. B.J.B. D., *Laboratory Inquiry in Chemistry*; Cengage: 2009.

(3) Gunnerson, C. C. K., *Exploring General Chemistry in the Laboratory*; Morton: 2017.

(4) Rossi, E. S.W.W. R., *Chemical Principles in the Laboratory*; Cengage: 2014.

Procedure

Calculating the maximum wavelength λ_{max}

☐ *Step 1:* – Make sure that there is no cuvette in the spectrophotometer. Close the lid and start the device.

☐ *Step 2:* – Obtain 7 test tubes.

☐ *Step 3:* – Mark each test tube with one of a series of identification numbers running from #1 to #7.

☐ *Step 4:* – Get 30 mL of the stock Co^{2+} solution in a 50 mL beaker in order to prepare a series of dilutions from this solution. Record the concentration (Molarity) written on the bottle in the Results sheet.

☐ *Step 5:* – Use Mohr pipets to make additions of Co^{2+} and distilled water as shown in the table below.

Test tube #	$V_{Co^{2+}}$ (mL)	$V_{Water(mL)}$
#1	5.0	0.0
#2	4.0	1.0
#3	3.5	1.5
#4	3.0	2.0
#5	2.5	2.5
#6	2.0	3.0
#7	1.0	4.0

☐ *Step 6:* – Gently swirl the solutions from side to side to make sure they are mixed.

☐ *Step 7:* – Set the wavelength at 400 nm.

☐ *Step 8:* – Fill a cuvette with water, make sure the walls of the cuvette are clean and dry, and make the absorbance of this sample zero–this is the blank.

☐ *Step 9:* – Fill another cuvette with the solution in the first test tube. Ideally you will rinse the cuvette first with 1 mL of the solution to be used. Place the cuvette in the holder and read the absorbance for 400 nm wavelength. You do not need to touch anything to measure. Just record your result. Mind not to dispose of this solution after the measurement.

☐ *Step 10:* – Repeat the previous 3 steps, but this time for 425 nm, then for 450 nm, and so on. Until reaching 600 nm. You need to redo the blank every time you change the wavelength. In another words, you need to change wavelength, place the water cuvette, press zero, add the sample and record.

☐ *Step 11:* – Graph Absorbance vs. wavelength.

☐ *Step 12:* – From the plot, select the wavelength at which the absorbance is largest.

☐ *Step 13:* – **Use this wavelength for all subsequent measurements.**

Good Lab Practice

✍ Mind not to use your finger as tube stopper.

Constructing the calibration curve

☐ *Step 14:* – Select the maximum wavelength obtained in the previous step in the spectrophotometer, and do the blank for this wavelength. After this point, there is no need to do the blank anymore.

☐ *Step 15:* – Measure again the absorbance of the first test tube and record the results.

☐ *Step 16:* – Measure the absorbance of the remaining tubes and record your results.

☐ *Step 17:* – Plot Absorbance vs. concentration with Absorbance on the Y and concentration on the X. To do that, follow the steps to fill the values on the Results' table.

☐ *Step 18:* – Calculate the slope of the graph by using the formula:

$$k = \frac{\sum c \cdot A}{\sum c^2}$$

☐ *Step 19:* – Show the results to your instructor to ensure your plot is correct.

Measuring the unknown

☐ *Step 20:* – Using a new test tube, add 2 mL of the unknown and 2 mL of water (we will call this V_{water}) and measure the absorbance. If the absorbance does not lie within the range of your previous results, change relative quantities of water and unknown.

☐ *Step 21:* – Knowing k_λ from the previous part of the experiment, solve for the concentration of the diluted solution $c_{Diluted}$:

$$c_{Diluted} = \frac{A}{k_\lambda}$$

☐ *Step 22:* – Now go back from the concentration of the diluted solution $c_{Diluted}$ to calculate the concentration of the unknown using the formula:

$$c_{Unknown} \cdot V_{unknown} = c_{Diluted} \cdot V_{Diluted}$$

$$c_{Unknown} = \frac{c_{Diluted} \cdot V_{Diluted})}{V_{unknown}}$$

☐ *Step 23:* – Make sure you stop the spectrophotometer and remove the cuvette.

Calculations

① This is the volume of cobalt solution added to the tube.

② This is the volume of water added to the tube.

③ This is the total volume: ① + ②

④ This is the concentration of cobalt after dilutions in each tube:

$$c_{diluted} = \frac{c_{initial} \cdot V_{Co^{2+}}}{V_{Total}} \qquad \textcircled{4} = \frac{c_{Co^{2+}} \cdot \textcircled{1}}{\textcircled{3}}$$

⑤ The measured absorbance

⑥ The product of the concentration and the absorbance for each measurement ⑤ × ④

⑦ The square of each diluted concentration $\textcircled{4}^2$

⑧ The sum (∑) of all the values in column ⑥ and the sum for all the values for column ⑦ respectively.

⑨ Here you calculate the slope of the A vs. c plot by doing:

$$k = \frac{\sum c \cdot A}{\sum c^2}$$

⑩ The volume of unknown you used.

⑪ The volume of water you added.

⑫ The absorbance you measured.

⑬ The concentration of cobalt in the sample you measured that is diluted is:

$$c_{Co,diluted} = \frac{\textcircled{12}}{\textcircled{9}}$$

⑭ The concentration of cobalt in the real original sample is:

$$c_{Unknown} = \frac{c_{Diluted} \cdot V_{Diluted}}{V_{unknown}} \qquad c_{Unknown} = \frac{\textcircled{13} \cdot \left(\textcircled{10} + \textcircled{11}\right)}{\textcircled{10}}$$

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Spectrochemistry

---

1. The absorbance of a colored substance in a colorless liquid is measured at each of a series of wavelengths, and the data is given below:

λ(nm)	325	350	375	400	425	450	475	500	525	550
A	0.016	0.144	0.341	0.578	0.681	0.558	0.281	0.092	0.031	0.004

   (a) Plot A vs. λ in the graph below.

   (b) Calculate the λ value that gives a maximum A.

$\lambda(A_{max})$ = ______________________________

2. A set of absorbances, A, for different concentrations, c, are given below.
   (a) Plot A vs. c in the graph below.
   (b) Compute the slope of the graph by using the formula:

$$k = \frac{\sum c \cdot A}{\sum c^2}$$

c (M)	A	$c \cdot A$ (M)	c^2 (M^2)
0.0120	0.681		
0.00960	0.540		
0.00720	0.389		
0.00480	0.270		
0.00240	0.133		
Sum			

k (M^{-1})=________________________________

3. After calculating k, now write down the formula for absorbance in the form: $A = k \cdot c$

STUDENT INFO

Name: ________________ Date: __________

Results
EXPERIMENT

Spectrochemistry

Calculating the wavelength

λ (nm)	400	425	475	500	525	550	575	600
A								

Max Abs. λ (nm)=________________ c_{CoCl_2} (M)=________________

Constructing the calibration curve

Test tube #	$V_{Co^{2+}}$ (mL) ①	V_{water} (mL) ②	V_{Total} (mL) ③	$c_{Co^{2+}}$ (M) ④	Absorbance A ⑤	$A \cdot c_{Co^{2+}}$ ⑥	$c^2_{Co^{2+}}$ ⑦
#1	5.0	0.0					
#2	4.0	1.0					
#3	3.5	1.5					
#4	3.0	2.0					
#5	2.5	2.5					
#6	2.0	3.0					
#7	1.0	4.0					
					⑧ Sum ($\sum$)=		
						⑨ k_λ=	

Measuring the unknown

Unknown Number =______________________________

Measurement of the unknown				
$V_{unknown}$ (mL) ⑩	V_{water} (mL) ⑪	Absorbance A ⑫	$c_{diluted}$ (M) ⑬	$c_{unknown}$ (M) ⑭

Figure 9: λ (X axis) vs. A (Y axis)

Figure 10: *c* (X axis) vs. A (Y axis)

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Spectrochemistry

---

1. Write a brief procedure outlining how would you prepare 250. mL of a 0.150 M solution of $CoCl_2$ from solid $CoCl_2$ and distilled water.

2. Using the solution from the previous question, how would you prepare 100. mL of a 0.060M diluted $CoCl_2$ solution.

EXPERIMENT

Thermal decomposition of a hydrate

Goal

The goal of this experiment is to study the thermal decomposition of a hydrate–barium chloride dihydrate–, and to compare this reaction with the thermal decomposition of a few other hydrates. The thermal decomposition of barium chloride dihydrate ($BaCl_2 \cdot 2\,H_2O$) produces water, whereas other oxides often times generate hydrochloric acid.

Materials

- ☐ Crucible and lid.
- ☐ Crucible tongs and wire gauze.
- ☐ about 1.5 g of $BaCl_2 \cdot 2\,H_2O\,(s)$.
- ☐ small test tubes.
- ☐ litmus paper.
- ☐ $CrCl_3 \cdot 6\,H_2O\,(s)$ and $CoCl_2 \cdot 6\,H_2O\,(s)$.

Background

The amount of product generated in a chemical decomposition can be predicted by means of stoichiometric calculations. Knowing an initial mass of the reactant the amount (mass) of products and more importantly, the mass increase or decrease in the reaction can be estimated . The thermal decomposition of $BaCl_2 \cdot 2\,H_2O$–an extensively studied reaction in the college laboratory literature [1–4]–generates $BaCl_2(s)$ and $H_2O\,(g)$. Clearly, if you start with certain mass of $BaCl_2 \cdot 2\,H_2O$ the weight of the final product should be lower, as the hydrate looses water molecules during the reaction.

Figure 11: Several chemical hydrates used in this experiment: (Left panel) $CaCl_2 \cdot 6\,H_2O$; (Central left panel) $CrCl_3 \cdot 6\,H_2O$; (Central right panel) anhydrous $CrCl_3$; (Right Panel) The set up for the experiment.

Basic reaction Stoichiometry

When studying chemical reactions, often times one needs to predict how much product can be produced from a given amount of reactants. To obtain this type of information we use *reaction stoichiometry*, and the key to employ reaction stoichiometry is a balanced chemical equation. The coefficients from the balanced chemical equation, known as stoichiometric coefficients, are needed to relate one substance with another.

Example

What is the number of moles of H_2O *that can be produced from 2.0 mol* $BaCl_2 \cdot 2H_2O$ *in the reaction:*

$$BaCl_2 \cdot 2H_2O\,(s) \longrightarrow BaCl_2\,(s) + H_2O\,(g)$$

Answer: the first step would be to balance the reaction:

$$\mathrm{BaCl_2 \cdot 2\,H_2O\,(s)} \longrightarrow \mathrm{BaCl_2\,(s)} + \mathbf{2}\mathrm{H_2O\,(g)}$$

The reaction produces two water molecules. We will relate the number of H_2O moles with the given number of $BaCl_2 \cdot 2H_2O$ moles using a conversion factor from the mole–to–mole ratio

$$n(\mathrm{H_2O}) = 2.0\,mol\,\mathrm{BaCl_2 \cdot 2H_2O} \times \frac{2\,mol\,\mathrm{H_2O}}{3\,mol\,\mathrm{BaCl_2 \cdot 2H_2O}} = 1.3\,mol\,\mathrm{H_2O}$$

Mass-to-mass predictions

In order to calculate how much product can be produced from a given mass of a reactant, we need first to convert the grams of reactant into moles, using its molar mass, to then use the mole ratio from the balanced equation. Finally, we need to convert the moles of product formed, back into mass units by means of the molar mass.

Example

Sulfur trioxide $BaCl_2 \cdot 2H_2O$ *(MW* $244.266\,g \cdot mol^{-1}$*) decomposes to produce* $BaCl_2$ *(MW* $208.236\,g \cdot mol^{-1}$*) and* H_2O *(MW* $18.015\,g \cdot mol^{-1}$*):*

$$BaCl_2 \cdot 2H_2O\,(s) \longrightarrow BaCl_2\,(s) + 2H_2O\,(g)$$

What mass of $BaCl_2$ *will be produced from 3.00 g of* $BaCl_2 \cdot 2H_2O$*?*

Answer: knowing that the equation is already balanced, we will first convert the grams of $BaCl_2 \cdot 2H_2O$ into moles to then relate the $BaCl_2 \cdot 2H_2O$ with the $BaCl_2$ moles produced. Finally, we will convert the $BaCl_2$ moles back into grams:

$$m(\mathrm{BaCl_2}) = 3.00\,g\,\mathrm{BaCl_2 \cdot 2H_2O} \times \frac{1\,mol\,\mathrm{BaCl_2 \cdot 2H_2O}}{244.266\,g\,\mathrm{BaCl_2 \cdot 2H_2O}} \times \frac{1\,mol\,\mathrm{BaCl_2}}{1\,mol\,\mathrm{BaCl_2 \cdot 2H_2O}} \times \frac{208.236\,g\,\mathrm{BaCl_2}}{1\,mol\,\mathrm{BaCl_2}} = 2.56\,g\,\mathrm{BaCl_2}$$

Hydrates

Learn about the crystallization of a *hydrate*.

Visual

Some compounds in contact with the moisture of air form crystals that incorporate water molecules. These compounds are called *hydrates*. Hydrates are named by first giving the name of the compound, followed by the word *hydrate* with a Greek prefix (mono, di, tri, etc.) indicating how many molecules of water are found in each formula unit. For example, the name of $CuSO_4 \cdot 5H_2O$ is copper(II) sulfate pentahydrate. The quantity of water will not change as long as the temperature (and pressure) is not changed. A substantial increase in temperature by heating the chemical however, will cause the loss of water. When we want to emphasize that a compound has lost its water of hydration, we call it *anhydrous*.

Example

Name $BaCl_2 \cdot 2H_2O$ *and give the formula for Sodium carbonate pentahydrate.*

Answer: $BaCl_2 \cdot 2H_2O$ is named **barium chloride dihydrate**, whereas the formula for sodium carbonate pentahydrate is $\mathbf{Na_2CO_3 \cdot 5H_2O}$.

Water composition of a hydrate

Different hydrates have different water compositions. For example $BaCl_2 \cdot 2H_2O$ contains two water molecules, whereas $CuSO_4 \cdot 5H_2O$ contains five water molecules. At the same time, each hydrate will have different mass % of water, depending on its water composition. $BaCl_2 \cdot 2H_2O$ is 15% water whereas $CuSO_4 \cdot 5H_2O$ is 36% water.

Example

Calculate the mass % of water in $BaCl_2 \cdot 2H_2O$ *(MW 244.266* $g \cdot mol^{-1}$*) and in* $CuSO_4 \cdot 5H_2O$ *(MW 250* $g \cdot mol^{-1}$*).*

Answer: $BaCl_2 \cdot 2H_2O$ has two water molecules and a molecular weight of 244.266 $g \cdot mol^{-1}$, and hence the mass % of water will be $\frac{2 \times 18.015\, g\, H_2O}{244.266\, g BaCl_2 \cdot 2H_2O} \times 100 = 14.750$ % of water. $CuSO_4 \cdot 5H_2O$ has five water molecules and a molecular weight of 249.681 $g \cdot mol^{-1}$, and hence the mass % of water will be $\frac{5 \times 18.015\, g\, H_2O}{249.681\, g\, CuSO_4 \cdot 5H_2O} \times 100 = 36.076\%$ water.

Litmus paper

Some hydrates produce acid when decomposing. Litmus paper is a quick test to identify whether a hydrate generates an acid. Litmus paper is a standard acid-base indicator. There are two variants for litmus paper. You can find red litmus paper and blue litmus paper. Blue litmus paper turns pink under acidic conditions and red litmus paper turns blue under basic or alkaline conditions.

Example

A solution is tested with red litmus and the paper turned blue. Indicate whether the solution is acidic or basic.

Answer: Acids turn blue litmus into pink and bases turn red litmus into blue. In this case as the red indicator turned blue, that means that the solution was a base.

	Red litmus paper	Blue litmus paper
Acidic solution	Red	Red
Basic solution	Blue	Blue

Heating hydrates

As hydrates contain water, if we warn them they will release this water and will reduce its mass. A way to calculate the amount of water in a hydrate is to warm up a sample and measure the reduction in mass. This mass reduction corresponds to the mass of water in the compound.

Example

A mass of 1g of $BaCl_2 \cdot xH_2O(s)$ *becomes 0.9 after heating up the hydrate for long enough time. Calculate the percentage of mass in the hydrate.*

Answer: The mass of water if 0.1g, as 1g of hydrate weights 0.9g after the thermal decomposition. Therefore, the chemical has 10% of mass.

References

(1) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.
(2) Bishop, K. W.C.B. M., *Experiments in General Chemistry*; Saunders College Publishing: 1992.
(3) Corwin, C., *Laboratory Experiments*; Prentice-Hall: 1985.
(4) Beran, J. A., *Laboratory manual for principles of general chemistry*; John Wiley and Sons: 2010.

Procedure

Getting started

☐ *Step 1:* – Place a clean, covered crucible in a clay triangle on an iron ring attached to a ring stand. Adjust the height of the ring so that the bottom of the crucible will be in the hottest part of the flame. The correct arrangement of the equipment, crucible, and burner is shown in the figure (Right panel).

☐ *Step 2:* – Learn how to use the Bunsen burner ✍. Heat the covered crucible in the hottest part of the flame for about 5 min, making sure that the bottom of the crucible attains a red glow.

☐ *Step 3:* – Stop the burner and allow the crucible to cool down completely ⚠.

☐ *Step 4:* – Weight the covered crucible and record the mass of the covered crucible.

Good Lab Practice

✍ Learn the correct use of a 🌍 *Bunsen burner*.

✍ Learn how to 🌍 *warm a test tube* on the flame.

✍ Make sure the gas outlet is closed before leaving the lab.

Decomposition of $BaCl_2 \cdot 2H_2O$

☐ *Step 5:* – Weight about 1.5 g of $BaCl_2 \cdot 2H_2O$.

☐ *Step 6:* – Add the $BaCl_2 \cdot 2H_2O$ sample onto the crucible and cover the crucible again. Weight the covered crucible with the $BaCl_2 \cdot 2H_2O$ and record the exact mass in the results sheet.

☐ *Step 7:* – Heat up the crucible in the hottest part of the flame for about 15 min. The bottom of the crucible should be red hot during this time.

☐ *Step 8:* – When the crucible is cool, weight and record the mass of the product.

Hydrates producing acid

☐ *Step 9:* – Place a pea-sized portion of $CrCl_3 \cdot 6H_2O$ in a test tube.

☐ *Step 10:* – Using the Bunsen Burner ⚠ and with a test tube clamp, hold the test tube at an angle of about 45° while swinging it above the hottest part of the flame.

☐ *Step 11:* – When steam forms test with blue litmus paper the liquid drops that condensate near the opening of the test tube.

☐ *Step 12:* – Repeat the previous 3 steps using $BaCl_2 \cdot 2H_2O$ and $CoCl_2 \cdot 6H_2O$ instead.

⚠ CAUTION!

⚠ Do not touch the crucible or the iron ring during this experiment as it should be hot.

⚠ Be aware that hot glassware look the same as when cold.

Calculations

① Record the mass of the empty crucible with the lid. Remember to weight the crucible in the balance only when completely cool.

② Record the mass of the empty crucible with the lid with $BaCl_2 \cdot 2H_2O$.

③ The mass of $BaCl_2 \cdot 2H_2O$ added to the crucible should be:

$$\text{Mass of } BaCl_2 \cdot 2H_2O = \textcircled{2} - \textcircled{1}$$

④ After you heat the crucible with $BaCl_2 \cdot 2H_2O$ a product will form. Weight the crucible and lid with the final product inside.

⑤ You should calculate the mass of product by doing:

$$\text{Mass Product} = \textcircled{4} - \textcircled{1}$$

⑥ Calculate the mass % of water in the hydrate:

$$\frac{(\text{Mass } BaCl_2 \cdot 2H_2O\,) - (\text{Mass Product})}{(\text{Mass of } BaCl_2 \cdot 2H_2O)} \times 100 = \frac{\textcircled{5} - \textcircled{3}}{\textcircled{3}} \times 100$$

⑦ Calculate the theoretical mass % of water in the hydrate

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Thermal decomposition of a hydrate

---

1. Name or give the formula of the following compounds:

Magnesium sulfate	____________________
$MgSO_4 \cdot H_2O$	____________________
Barium chloride	____________________
$BaCl_2 \cdot 2\,H_2O$	____________________
Chromium(III) chloride	____________________
$CoCl_2 \cdot 6\,H_2O$	____________________
Nickel(II) sulfate heptahydrate	____________________

2. The thermal decomposition of 2.00 g of $MgSO_4 \cdot H_2O$ (MW 138.383 $g \cdot mol^{-1}$) produces 1.7 g of a$MgSO_4$ (MW 120.366 $g \cdot mol^{-1}$).

$$MgSO_4 \cdot H_2O\,(s) \xrightarrow{\Delta} MgSO_4(s) + H_2O\,(g)$$

   (a) Calculate the mass of water in the compound.

   (b) Calculate the water % mass in the compound.

3. Calculate the theoretycal mass % of water of the hydrate: $Li_2SO_4 \cdot H_2O\,(s)$.

4. A solution is tested with blue litmus and the paper turned red. Indicate whether the solution is acidic or basic.

STUDENT INFO

Name: ______________________ Date: __________

Results

EXPERIMENT

Thermal decomposition of a hydrate

Decomposition of $BaCl_2 \cdot 2H_2O$

① Mass of empty crucible and lid (g)	
② Mass of crucible and lid with $BaCl_2 \cdot 2H_2O$ (g)	
③ Mass of $BaCl_2 \cdot 2H_2O$ (g)	
④ Mass of crucible, lid and product (g)	
⑤ Mass of dry product (g)	
⑥ Mass % of water	
⑦ Theoretical mass % of water in $BaCl_2 \cdot 2H_2O$	

Hydrates producing acid

	Litmus Color Before	Litmus Color After	Acid produced?
$BaCl_2 \cdot 2H_2O$			
$CrCl_3 \cdot 6H_2O$			
$CoCl_2 \cdot 6H_2O$			

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Thermal decomposition of a hydrate

---

1. Name or give the formula of the following compounds:

Zinc sulfate	____________
$ZnSO_4 \cdot 7\,H_2O$	____________
Iron(II) sulfate	____________
$FeSO_4 \cdot 7\,H_2O$	____________
Sodium carbonate	____________
$Na_2CO_3 \cdot 10\,H_2O$	____________

2. 3.0 grams of an unknown hydrate is thermally decomposed giving 2.0 grams of product.

   (a) Calculate the mass of water in the hydrate

   (b) Calculate the mass % of water in the hydrate.

EXPERIMENT

Precipitation and acid-base reactions

Goal

The goal of this experiment is to carry a set of simple acid-base and precipitation reactions in aqueous solutions. On one hand you will carry precipitation reactions and you will put in practice the *solubility rules*. You will have to observe the final product of the reaction and identify it as *soluble* or *insoluble*. On the other hand, you will examine some simple *acid-base reactions*, identify the acidic/basic character of each chemical and figure out whether there is heat involved in the acid-base reaction.

Materials

- ☐ Red and blue litmus paper.
- ☐ 4 small test tubes.
- ☐ solutions of: NH_4NO_3 (aq), HCl (aq), CH_3COOH (aq),
- ☐ solutions of: NaOH (aq), Na_2CO_3 (aq) and NH_3 (aq)
- ☐ solutions of: NaCl (aq), NaBr (aq), Na_2SO_4 (aq), and KCl (aq)
- ☐ solutions of: $Ba(NO_3)_2$ (aq), $AgNO_3$ (aq), $Pb(NO_3)_2$ (aq), and $Ni(NO_3)_2$ (aq)

Background

Learn about 🌍 *precipitation* reactions.

Visual

Many chemical reactions take place in aqueous solution, and most of these reactions involve ions. Let's consider as an example what happens when mixing a colorless solution of silver nitrate ($AgNO_3$) with a colorless solution of sodium chloride (NaCl). The solution of silver nitrate contains Ag^+ cations and $NO_3{}^-$ anions, whereas the solution of sodium chloride contains Na^+ cations and Cl^- anions. When we mix these two aqueous solutions, a white precipitate (AgCl) forms immediately due to the ion exchange process.

$$AgNO_3\,(aq) \quad + \quad NaCl\,(aq) \quad \longrightarrow \quad AgCl\,(s) \quad + \quad NaNO_3\,(q)$$

On the other hand, acid-base reactions are an important type of aqueous reactions. Acids–substances with a sharp or sour taste–and bases–chemicals with a soapy feel– react by means of a neutralization reaction. An example would be:

$$HCl\,(aq) \quad + \quad NaOH\,(aq) \quad \longrightarrow \quad NaCl\,(aq) \quad + \quad H_2O\,(l)$$

This experiment addresses these two important types of chemical reactions, studying several classic precipitation and acid-base reactions. Similar experiments have been performed in the literature [1, 2], and some of the reagents of this experiment are adapted from [3].

Weak and strong electrolytes

Electrolytes are chemicals that when placed in water break or dissociate into ions with positive and negative charges. An example of an electrolyte would be NaCl, which is a solid. When dissolved in water it produces Na^+ and Cl^-. On the other hand, non-electrolytes are chemicals that when dissolved in water they stay on its molecular form. An example of this is ethanol, CH_3CH_2OH, that you can find in wine. Once dissolved in water it stays as a molecule and does not produce positive and negative ions. Electrolytes can be classified as strong or weak electrolytes, depending on their degree of dissociation in water. Strong electrolytes dissociate completely producing large amounts of ions, whereas weak electrolytes dissociate only partially to produce just a few ions.

Precipitation reactions

Some ionic compounds are soluble in water whereas others are not. In a precipitation reaction, two electrolyte solutions are mixed to produce an insoluble solid called *precipitate*. When an insoluble substance is formed in water, it immediately precipitates, separates from the solution in solid form.

Ions that form *soluble* compounds...	... **except** when combined with
Group I ions (Na^+, Li^+, K^+, etc)	no exceptions
Ammonium (NH_4^+)	no exceptions
Nitrate (NO_3^-)	no exceptions
Acetate (CH_3COO^-)	no exceptions
Hydrogen carbonate (HCO_3^-)	no exceptions
Chlorate (ClO_3^-)	no exceptions
Halides (F^-, Cl^-, Br^-)	Pb^{2+}, Ag^+ and Hg_2^{2+}
Sulfate (SO_4^{2-})	Ag^+,Ca^{2+}, Sr^{2+},Ba^{2+}, Hg_2^{2+} and Pb^{2+}
Ions that form *insoluble* compounds...	... **except** when combined with
Carbonate (CO_3^{2-})	group I ions (Na^+, Li^+, K^+, etc) or ammonium (NH_4^+)
Chromate (CrO_4^{2-})	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+} or ammonium (NH_4^+)
Phosphate (PO_4^{3-})	group I ions (Na^+, Li^+, K^+, etc) or ammonium (NH_4^+)
Sulfide (S^{2-})	group I ions (Na^+, Li^+, K^+, etc) or ammonium (NH_4^+)
Hydroxide (OH^-)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}, Sr^{2+} or ammonium (NH_4^+)

Ionic and Net ionic equations

The ionic equation for a precipitation reaction shows all the species as they actually exist in solution. Because dissolved ionic compounds exist as separate aqueous ions, in an ionic equation the ions should be shown separately. Some of these ions appear as both reactants and products. This means that they play no role in the reaction: they are spectator ions. In the ionic equation, you can simplify the chemical equation by canceling the spectators out on each side of the arrow.

Example

Write the net ionic equation for the reaction between aqueous solutions of silver nitrate ($AgNO_3$) and potassium chromate (K_2CrO_4) to give a precipitate of silver chromate (Ag_2CrO_4) and a solution of potassium nitrate (KNO_3):

$$AgNO_3(aq) \quad + \quad K_2CrO_4(aq) \quad \longrightarrow \quad Ag_2CrO_4(s) \quad + \quad KNO_3(aq)$$

Answer: the complete ionic equation, which shows all the dissolved ions, both those that react and the spectator ions, is:

$$\mathrm{Ag^+(aq)} \ + \ \mathrm{NO_3{}^-(aq)} \ + \ \mathrm{2\,K^+(aq)} \ + \ \mathrm{CrO_4{}^{2-}(aq)} \longrightarrow \mathrm{Ag_2CrO_4(s)} \ + \ \mathrm{K^+(aq)} \ + \ \mathrm{NO_3{}^-(aq)}$$

Now cancel the spectator ions K^+ and $NO_3{}^-$

$$\mathrm{2\,Ag^+(aq)} \ + \ \cancel{\mathrm{2\,NO_3{}^-(aq)}} \ + \ \cancel{\mathrm{2\,K^+(aq)}} \ + \ \mathrm{CrO_4{}^{2-}(aq)} \longrightarrow \mathrm{Ag_2CrO_4(s)} \ + \ \cancel{\mathrm{2\,K^+(aq)}} \ + \ \cancel{\mathrm{2\,NO_3{}^-(aq)}}$$

and obtain the net ionic equation:

$$\mathrm{2\,Ag^+(aq)} \quad + \quad \mathrm{CrO_4{}^{2-}(aq)} \quad \longrightarrow \quad \mathrm{Ag_2CrO_4(s)}$$

Acids and bases

Acids are chemical that produce hydrogen ions (H^+, also known as protons) in water. Bases, on the other hand, accept hydrogen ions, producing hydroxide ions (OH^-) in water. Acids and bases change the color of certain chemicals called indicators, and litmus is a well-known acid-base indicator. Aqueous solutions of acids turn litmus red, whereas aqueous solutions of bases turn it blue. Acids and bases can be classified as strong or weak according to the extent to which they dissociate in solution. This way, a strong acid is completely dissociated in solution, whereas a weak acid is only slightly dissociated. The same can be applied to bases.

Strongs acids	*Strongs* bases
Hydrobromic acid (HBr)	Group I hydroxides (NaOH, KOH, etc.)
Hydrochloric acid, (HCl)	Alkaline earth metal hydroxides ($Ba(OH)_2$, $Mg(OH)_2$, etc.)
Hydroiodic acid, (HI)	Group I and Group II oxides (Na_2O, BaO, etc.)
Nitric acid (HNO_3)	
Perchloric acid ($HClO_4$)	
Chloric acid ($HClO_3$)	
Sulfuric acid (H_2SO_4)	

Weak acids	*Weak* bases
HF (aq), CH_3COOH (aq) , H_2SO_3 (aq)	NH_3 (aq)

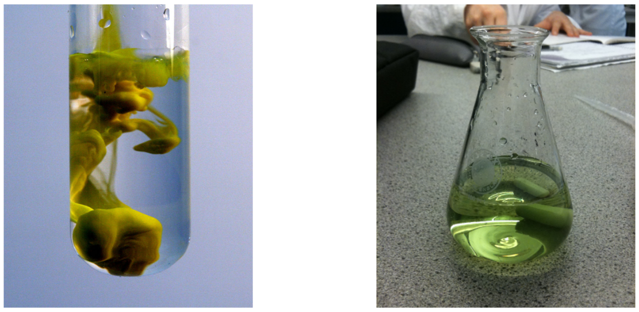

Figure 12: (Left panel) A precipitate of PbI_2; (Right panel) The result of an acid-base neutralization.

Neutralization

A neutralization reaction is a reaction between a strong acid and a strong base to give a salt and water:

$$\text{Acid (aq)} \quad + \quad \text{Base (aq)} \quad \longrightarrow \quad \text{Salt (aq)} \quad + \quad H_2O\,(l)$$

When we write the net ionic equation for the neutralization of a weak acid or a weak base, we use the molecular form of the weak acid or base and not its ionic form, because molecules are the dominant species in solution.

Litmus paper

Litmus paper is a quick test to identify whether a solution is acidic or basic. There are two variants for litmus paper. You can find red litmus paper and blue litmus paper. Blue litmus paper turns pink under acidic conditions whereas red litmus paper turns blue under basic conditions.

	Red litmus paper	Blue litmus paper
Acidic solution	Red	Red
Basic solution	Blue	Blue

References

(1) Fetzer-Gislason, P. R.D.W. S., *Lab Experiments in Introductory Chemistry*; Freeman: 2003.
(2) Murov, S., *Experiments in General Chemistry*; Cengage: 2013.
(3) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

Procedure

Getting started

☐ *Step 1:* – Obtain the material you need (litmus paper, test tubes, reactants).

☐ *Step 2:* – Obtain instructions of how to use the centrifuges from the lab instructor in order to separate some of the insoluble compounds. You can learn more in the following video link.

☐ *Step 3:* – Prepare solutions of NaCl and KCl by weighting 0.06g of KCl and 0.03g of NaCl. Dissolve each salt in 100mL of water. Label the solutions.

Good Lab Practice

✍ Learn the correct use of a 🌍 *centrifuge*.

✍ Discharge the contents of each test tube as directed by your lab instructor.

Precipitation reactions

☐ *Step 4:* – Arrange in the following order the set of reactants Type A: $NaCl\,(aq)$, $NaBr\,(aq)$, $Na_2SO_4\,(aq)$ and $KCl\,(aq)$

☐ *Step 5:* – Arrange in the following order the set of reactants Type B: $Ba(NO_3)_2(aq)$, $AgNO_3(aq)$, $Pb(NO_3)_2(aq)$ and $Ni(NO_3)_2(aq)$

☐ *Step 6:* – Pick up one of the reactants Type A such as $NaCl\,(aq)$ and add 20 drops to a set of four test tube. Each tube should have now 20 drops of a reactant Type A.

☐ *Step 7:* – Pick up one of the reactants Type B such as $Ba(NO_3)_2(aq)$. Add 20 drops to the first test tubes containing the reactant A. Write down your observations on the Results table. Indicate if soluble as (S) or if insoluble as (I).

☐ *Step 8:* – Now pick another Type B reactant and add 20 drops to the second test tubes containing the reactant A. Write down the result on the Results table as soluble (S) or insoluble (I). Do this until you run all the reactants Type B.

☐ *Step 9:* – Repeat this process for another Type A reactant, such as NaBr (aq). Add 20 drops to a set of four test tubes and try adding different reactants Type B to each of these test tubes. Write down the result on the Results table as soluble (S) or insoluble (I). Once done, continue until you run all the reactants Type A.

☐ *Step 10:* – In a new, clean test tube mix 20 drops of $Ni(NO_3)_2$ and 20 drops of NaOH. You have not done this mixture in the experiments above. Centrifuge the test tube for 3 seconds. Then, add HCl to the precipitate and described the observations along the whole experiment.

⚠CAUTION!

⚠ When using the centrifuge, do not attempt to stop the rotor with your fingers or anything else.

⚠ Sodium hydroxide, hydrochloric acid, acetic acid and ammonia can cause chemical burns.

⚠ Barium is poisonous. Wash your hands thoroughly after using it.

Acidic and Basic character

☐ *Step 11:* – Have handy each of the following reactants: HCl, NH_3, NH_4NO_3 (aq), CH_3COOH (aq) and Na_2CO_3 (aq).

☐ *Step 12:* – Add a drop of each reactant to a small piece of blue litmus. Write down the resulting color.

☐ *Step 13:* – Add a drop of each reactant to a small piece of red litmus. Write down the resulting color.

Acids-Base reactions

☐ *Step 14:* – Have handy the set of reactants Type C: NaOH (aq), Na_2CO_3 (aq) and NH_3 (aq)

☐ *Step 15:* – Have handy the set of reactants Type D: NH_4NO_3 (aq), HCl (aq), and CH_3COOH (aq)

☐ *Step 16:* – Pick up one of the reactants Type C such as NaOH (aq) and add 20 drops to a set of four test tube. Each tube will have now 20 drops of a reactant Type C.

☐ *Step 17:* – Pick up one of the reactants Type D such as NH_4NO_3 (aq). Add 20 drops to the first test tube containing the reactants C. Using litmus paper, test the acidity of the resulting solution. Write down the result on the Results table as acidic (A) or basic (B).

☐ *Step 18:* – Now pick another Type D reactant and add 20 drops to the second test tube containing the reactant C. Using litmus paper, test the acidity of the resulting solution. Write down the result on the Results table as acidic (A) or basic (B).

☐ *Step 19:* – Now, pick up another Type C reactant and follow the cycle above. Add 20 drops to a set of four test tubes. And try adding reactants Type D to each test tube. Using litmus paper, test the acidity of the resulting solution. Write down the acid/base character on the Results table as acidic (A) or basic (B). Do this until you run all the reactants Type D.

STUDENT INFO

Name: ______________________ Date: ________

Pre-lab Questions

Precipitation and acid-base reactions

---

1. When mixing sodium phosphate (Na_3PO_4) and lead(II) nitrate ($Pb(NO_3)_2$):
    (a) Predict the precipitate likely to be formed, if any.
    (b) Write the net ionic equation for the reaction.

2. Predict the solubility (*soluble* or *insoluble*) of the following substances.

Compound	Solubility
$NaNO_3$	
NH_4Br	
$BaBr_2$	
$Ni(OH)_2$	
$BaCO_3$	

3. Write the balanced overall, complete ionic, and net ionic equations corresponding to each of the following reactions:
    (a) $BaCl_2(aq) + K_3PO_4(aq) \longrightarrow Ba_3(PO_4)_2(s) + KCl(aq)$
    (b) $2\,HNO_3(aq) + Ba(OH)_2(aq) \longrightarrow Ba(NO_3)_2(aq) + 2\,H_2O(l)$

STUDENT INFO

Name: ______________________ Date: __________

Results
EXPERIMENT

Precipitation and acid-base reactions

Precipitation reactions				
	NaCl	KCl	Na_2SO_4	NaBr
$Ni(NO_3)_2$				
$Pb(NO_3)_2$				
$AgNO_3$				
$Ba(NO_3)_2$				

(write *S* for soluble product and *I* for insoluble product. If insoluble, indicate the color of the precipitate)

Acids and bases			
	Red Litmus	Blue Litmus	Acid/Base
HCl			
NH_3			
NH_4NO_3 (aq)			
CH_3COOH (aq)			
Na_2CO_3 (aq)			

(write the color of the Litmus paper (red/blue).)

Acid-base reactions			
	NaOH (aq)	Na_2CO_3 (aq)	NH_3 (aq)
NH_4NO_3 (aq)			
HCl (aq)			
CH_3COOH (aq)			

(write the color of the Litmus paper (red/blue).)

STUDENT INFO

Name: ______________________ Date: ________

Post-lab Questions

Precipitation and acid-base reactions

---

1. When mixing barium nitrate ($Ba(NO_3)_2$) and sodium hydroxide (NaOH):
   (a) Predict the precipitate likely to be formed, if any.

   (b) Write the net ionic equation for the reaction.

2. Complete the overall equation, and write the complete ionic equation and the net ionic equation for the following acid-base reaction. If the substance is a weak acid or base, leave it in its molecular form in the equations.

$$HF\,(aq) \quad + \quad NaOH\,(aq) \quad \longrightarrow$$

3. Identify the following compounds as strong acid, weak acid, strong base or weak base:

HF	______________________	HF	______________________
HCl	______________________	NH_3	______________________
CH_3COOH	______________________	NaOH	______________________

EXPERIMENT

Titration of a weak acid

Goal

Weak acids are acids that do not dissociate completely, releasing only some of its hydrogen atoms into the solution. Acetic acid (CH_3COOH) is a very important weak acid, produced from the fermentation ethanol from the wine. Commercial acetic acid–what you know as vinegar–is just an aqueous solution of acetic acid. The goal of this experiment is to calculate the *molar concentration* of a sample of acetic acid by means of a standard chemical procedure known as *titration*. In order to do that you will react the weak acid with a basic solution of sodium hydroxide (NaOH), which has a known concentration. You will also use phenolphthalein as *indicator*.

Materials

☐ a 10 mL transfer pipet and a 50 mL buret

☐ solutions of acetic acid and NaOH

☐ phenolphthaleine

Background

A titration is a technique where a solution of known concentration–often times a base–is used to determine the unknown concentration of another solution–often times an acid [1]. Both substances react with each other in an acid-base reaction. The solution of known concentration is delivered carefully from a buret until an indicator–a third substance added to indicate the end of the titration–changes color. This experiment is a classic practice, previously implemented in numerous college chemistry laboratory manuals [2–5].

Concentration

We can express the composition of a solution as the *mass percentage* (%mass) of each component: solute and solvent.

$$\%mass = \frac{\text{mass of solute}}{\text{mass of solute + mass of solvent}} \times 100.\%$$

For example, if we dissolve 15 g of NaCl in 60. g of water, the total mass of the solution –solute plus solvent– is 75 g and the mass percentage of NaCl in the solution is $(15\,g/75\,g)\cdot 100\% = 20.\%$ NaCl. In chemistry, the molar concentration, c, of a solute in a solution, the *molarity* of the solute, is the moles of solute n present in a given volume, V, of the solution in liters. The units of molarity are moles per liter ($mol \cdot L^{-1}$), and it is denoted as M.

$$M = \frac{\text{moles of solute}}{\text{L of solution}} = \frac{n}{V}$$

Density, although not a measurement of concentration, is a property of liquid that we can use to convert mass into volume or volume into mass. The formula for density d is:

$$d = \frac{\text{grams of solution}}{\text{mL of solution}}$$

Density can be used in combination with the molar mass, M_w, which converts mass into moles, to link the two expression of concentration learned above, mass percent and Molarity.

$$\frac{c \times M_W}{d} \times 10^{-1} = \%\ \text{mass}$$

use dimensional analysis to verify the expression above.

$$\frac{\frac{mol}{L} \times \frac{g}{mol}}{\frac{g}{mL}} \times \frac{1\,L}{1000\,mL} \times 100\% = 10^{-1}\%$$

Example

What is the molarity of a sodium hydroxide solution (NaOH, M_W=39.997 $g \cdot mol^{-1}$) prepared by dissolving 15.00 g of the solute in enough water to make 350.0 mL of solution?

Answer: the molecular mass of NaOH is 39.997 $g \cdot mol^{-1}$ and the number of NaOH moles are:

$$n_{solute} = 15.00\,g \times \frac{1\,mol}{39.997\,g} = 0.3750\,mol$$

Do not forget to convert the volume from mL to L.

$$350.0\,mL \times \frac{1\,L}{1000\,mL} = 0.3500\,L$$

The molarity will be:

$$M = \frac{0.3750\,mol}{0.3500\,L} = 1.071\,M$$

Volumetric analysis

Learn how to do a 🌍 *acid-base titration.*

Visual

The determination of concentration by measuring volumes is called *volumetric analysis*. Titrations are volumetric analyses where a buret is used to add and measure the volume of one of the reactants. Acid-base titrations are extensively used chemical techniques employed to determine solute concentration in a solution. In a *acid-base titration*, an acid reacts with a base by gradually adding one solution to the other. The volume of the second solution is known, and the volume of the first solution required for the complete reaction is measured. The formula to use in a titration is:

$$n_a \cdot c_a \cdot V_a = n_b \cdot c_b \cdot V_b,$$

where n_a and n_b is the number of protons and hydroxides, respectively, of the acid and the base involved in the titration. c_a and V_a are the concentration of the acid and the volume of acid employed, and c_b and V_b are the concentration of the base and the volume of base employed.

Example

A 25 mL solution of sulfuric acid, H_2SO_4 –which has two acidic protons– is titrated with NaOH 0.10 M. The end point for the reaction is reached after 40. mL of the NaOH solution are added. Find the molarity of the sulfuric acid solution.

Answer: the balanced equation for the acid-base reaction is:

$$H_2SO_4(aq) \quad + \quad 2\,NaOH(aq) \quad \longrightarrow \quad 2\,H_2O \quad + \quad Na_2SO_4$$

The acid (H_2SO_4) has two protons, hence $n_a = 2$. Its molarity, c_a, is unknown whereas the volume used, $V_a = 25\,mL$, is given.
The base (NaOH) has one hydroxide, hence $n_b = 1$. The base concentration, $c_b = 0.10\,M$, and volume, $V_b = 40.\,mL$, are given:

$$c_a = \frac{n_b \cdot c_b \cdot V_b}{n_a \cdot V_a}$$

$$c_a = \frac{1 \cdot 0.10\, M \cdot 40.\, mL}{2 \cdot 25\, mL} = 0.080\, M$$

An *indicator* is used to indicate the exact end of the reaction. The indicator chosen will have one color before the reaction is complete and a different color when the acid-base reaction finishes. For example in the reaction between acetic acid (CH_3COOH) and sodium hydroxide (NaOH):

$$CH_3COOH \quad + \quad NaOH \longrightarrow CH_3COONa \quad + \quad H_2O$$

using phenolphtalein as the indicator, the solution will be colorless before the competition of this reaction but pink after completion. At a specific point during the titration, a single drop of the NaOH from the buret will cause the solution being titrated to turn from colorless to a barely discernible pink color. This point is called the *endpoint*.

Figure 13: (Left panel) A buret employed in a titration. (Center panel) Illustration showing the correct way to control the buret valve. (Right panel) The titration set up

References

(1) Crouch, D. S.D.W.F.L.H. S., *Fundamentals of analytical chemistry*; Nelson Education: 2013.

(2) Beran, J. A., *Laboratory manual for principles of general chemistry*; John Wiley and Sons: 2010.

(3) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

(4) Russo, T., *Merrill Chemistry Lab Manual*; McGraw-Hill: 1998.

(5) Holman, G. H. J., *Chemistry in Context*; McGraw-Hill Education: 2014.

Procedure

Getting started

☐ *Step 1:* – Obtain a 10 mL transfer pipet and a 50 mL buret with a stand and buret clamp.

☐ *Step 2:* – Obtain about 30 mL of acetic acid solution in a 50 mL beaker and about 80 mL of the NaOH solution in a clean, dry Erlenmeyer flask. Keep the NaOH solution containing Erlenmeyer closed with a rubber stopper.

☐ *Step 3:* – Clean your buret and fill it with the NaOH solution using a plastic funnel.

☐ *Step 4:* – Record the initial volume in the buret as zero. Mind you read the buret from the top and as you filled it up with NaOH the initial reading should be zero.

Good Lab Practice

✍ Learn how to execute an 🌍 *acid-base titration.*

Doing the titrations

☐ *Step 1:* – Pipet 5.00 mL of acetic acid into a clean 125 mL Erlenmeyer flask that has 20 mL of distilled water and 2 drops of phenolphthalein.

☐ *Step 2:* – Record the molarity of the NaOH solution indicated in the stock solution bottle.

☐ *Step 3:* – Place the flask under the buret. Use a piece of white paper under the flask to distinguish better the color change.

☐ *Step 4:* – Add the NaOH solution from the buret in 1 mL portions, while swirling the solution in the flask.

☐ *Step 5:* – The titration is completed when an addition of 1 mL causes the color to change from colorless to any shade of pink.

☐ *Step 6:* – Record the final buret volume.

☐ *Step 7:* – Repeat the steps above four times and average the resulting acetic acid concentration.

⚠CAUTION!

⚠ All acids and bases in this experiment could cause chemical burns.

Calculations

① Record the initial volume of the buret. This value is not necessarily 0.00 mL.

② Record the final volume of the buret, after you reached the end point.

③ The volumen of NaOH used should be:②–①

④ You can calculate the molarity of the acetic acid solution by means of:

$$c_a = \frac{\textcircled{3} \cdot c_b}{5\text{ mL}}$$

If the professor ask you to use a different volume of acetic acid, the value '5 mL' will need to be adjusted. where c_b is the given molarity of the NaOH solution.

⑤ Is the average of the 4 concentrations calculated.

$$\frac{\sum \textcircled{4}}{4}$$

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Titration of a weak acid

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1. A 10.00 mL sample of aqueous HCl requires 31.00 mL of 0.0900 M NaOH to reach the endpoint. What is the molar concentration of HCl. The equation for the reaction is:

$$HCl + NaOH \longrightarrow NaCl + H_2O$$

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2. The molarity of a vinegar solution is 0.90 M. Calculate the number of acetic acid moles in 10. mL of this solution.

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3. Phosphoric acid (H_3PO_4) is an acid with three protons. Suppose you titrate 5.00 mL of of this acid with NaOH 0.10 M. Knowing that the end point is reached after 25.00 mL of the base is added, find the molarity of the acid solution.

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STUDENT INFO

Name: ______________________ Date: __________

Results
EXPERIMENT

Titration of a weak acid

Titration

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Molarity of NaOH, c_b = ______________________

		1	2	3	4
①	Initial Buret Volume (mL)				
②	Final Buret Volume (mL)				
③	NaOH Volume used (mL)				
④	CH_3COOH Concentration (M)				

⑤ Mean CH_3COOH Concentration (M) = ______________________

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Titration of a weak acid

---

1. You need to prepare a sample containing 0.20 g of $CuSO_4$ from a solution that is 10.% $CuSO_4$ by mass. What mass of solution do you need?

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2. A 10.00 mL sample of aqueous H_2SO_4 requires 20.00 mL of 0.201 M NaOH to reach the endpoint. Calculate the molarity of H_2SO_4.

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3. You titrate a vinegar sample–an acetic acid solution in water–with 0.30 M NaOH. Using 10. mL of vinegar, you reach the endpoint after 10. mL of the bases are added. Indicate the molarity of the acetic acid solution.

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EXPERIMENT

Gravimetry

Goal

The goal of this experiment is to study the reaction between barium sulfate ($Ba(NO_3)_2$) and sulfamic acid (NH_2SO_3H) in hot water to produce barium sulfate ($BaSO_4$). This reaction is called hydrolysis. Using hot gravity filtration and by means of a gravimetric analysis you will calculate the reaction yield of the reaction.

Materials

- ☐ $Ba(NO_3)_2$ and NH_2SO_3H
- ☐ A 250 mL beaker and a graduated cylinder
- ☐ Stirring rod and rubber policeman (see the Figure)
- ☐ Filter paper and funnel
- ☐ Stand and iron ring
- ☐ Tongs

Background

Sulfamic acid (NH_2SO_3H), also known as amidosulfonic acid, is a very common chemical used to remove grout and mortar haze, as well as rust and mineral deposits [1]. It can indeed be found in many home improvement retail store [2]. The hydrolysis of this acid (the reaction with water at high temperature) produces ammonium hydrogen sulfate: $NH_2SO_3H + H_2O \longrightarrow NH_4HSO_4$ The hydrolysis of sulfamic acid to give sulfates can be used to precipitate barium in the form of barium sulfate [3].

$$Ba(NO_3)_2 \quad + \quad NH_2SO_3H \quad + \quad H_2O \quad \longrightarrow \quad BaSO_4 \quad + \quad NH_4NO_3 \quad + \quad HNO_3$$

By weighting the amount of precipitate, you can trace the amount of barium in the solution. In this experiment, you will react sulfamic acid with sodium nitrate in hot water to produce a precipitate of barium sulfate, focusing on the yield of the reaction. The advantage of using sulfamic acid instead of other type of sulfur acids, such as for example sulfuric acid, is that sulfamic acid produces coarse crystalline precipitate with fewer impurities [4].

Reaction yield

Stoichiometric calculations can predict the amount of product that will be formed in a chemical reaction. However, chemical reactions are not perfect, and often times reactions do not proceed to fully completed, and the amount of product predicted is not the one obtained in the experiment. The *yield* of a chemical reaction refers to the amount (mass or moles) of product obtained in the experiment (the actual) with respect to the theoretical quantity obtained according to stoichiometric calculations (the theoretical).

$$\%\,\text{Yield} = \frac{\text{Actual amount}}{\text{Theoretical amount}} \times 100\%$$

Example

We mix 2.0 moles of $Ba(NO_3)_2$ *with an excess of* H_2SO_4 *to produce a* $BaSO_4$ *precipitate. We obtain 1.0 moles of* $BaSO_4$ *:*

$$Ba(NO_3)_2(aq) \quad + \quad H_2SO_4(aq) \quad \longrightarrow \quad BaSO_4(s) \quad + \quad 2HNO_3(aq)$$

Calculate the reaction yield.

Answer: according to stoichiometric calculations, (coefficients in the balance chemical equation) 2 moles of $\mathrm{Ba(NO_3)_2}$ produce 2 moles of $\mathrm{BaSO_4}$). Since we start with 2.0 moles of $\mathrm{Ba(NO_3)_2}$, one would expect to also produce 2 moles of $\mathrm{BaSO_4}$. But the actual numbers of moles of $\mathrm{BaSO_4}$ is 1.0 mole. Hence the yield is

$$\%\,\text{Yield} = \frac{\text{Actual amount}}{\text{Theoretical amount}} \times 100\% = \frac{1.0\text{ moles}}{2.0\text{ moles}} \times 100\% = 50\%.$$

Limiting Reactant

In a chemical reaction involving two reactants normally one of the reactants limits the reactions, that means, limits the amount of product generated. This means that the chemical reaction will proceed as long as there is limiting reactant. After all limiting reactant has reacted the process will stop and an excess of the other reactant will remain unused. Most of the times is not so obvious to identify the limiting reactant and one needs to do some basic calculations to identify the limiting reactant as well as the excess reactant.

Example

We mix 2 moles of NaOH with an 1.5 moles of $MgCl_2$:

$$2NaOH(aq) \quad + \quad MgCl_2(aq) \quad \longrightarrow \quad Mg(OH)_2(s) \quad + \quad 2NaCl(aq)$$

Identify the limiting and the excess reactant.

Answer: we first choose one of the reactants, for example NaOH, to then convert the number of moles of that reactant into how many moles will be needed from the other reactant, according to the stoichiometry of the balanced chemical equation:

$$n_{\mathrm{MgCl_2}} = 2\text{ moles of NaOH} \times \frac{1\text{ mole MgCl}_2}{2\text{ moles NaOH}} = 1\text{ mole of MgCl}_2$$

This means that in order to react with 2 moles of NaOH we would need 1 mole of $\mathrm{MgCl_2}$. Because we have indeed 1.5 moles of $\mathrm{MgCl_2}$ that means we do have enough $\mathrm{MgCl_2}$ and hence NaOH should be the limiting reactant. $\mathrm{MgCl_2}$ would be the excess reactant.

Filtration

Learn more about **filtration** in this video link 🌍 Video Link

Visual

Filtration is a technique employed in chemistry to separate an insoluble solid compound (the precipitate) from the liquid. The mixture containing the solid is poured through a fine mesh (the filter) and gravity makes the liquid go through the filter, while the solid remains on it. When the mixture containing the solid is very hot, we call the filtration procedure: "hot gravity filtration". Different types of filter paper are available with a variety of porosities in order to separate different types of solids. Filter paper must be folded before proceeding with the filtration, forming a cone. In order to do this, one needs to first fold the filter in half, and then in half again, as shown in the Figure. Do not press the tip of the cone while folding, because it will weaken the paper.

Gravimetric Analysis

Learn more about **gravimetric analysis** in this video link 🌍 Video Link

Visual

Gravimetric analysis is a technique used to find out how much of a metal ion, such as barium, is in a solution. Think about a water sample containing $\mathrm{Ba^{2+}}$ dissolved. To determine how much barium is in the solution it would be necessary to separate it from the solvent. One way to separate the metal would be to precipitate the $\mathrm{Ba^{2+}}$ ions by means of a chemical reaction. Later, by filtering off and weighting the precipitate (containing all the barium) you would be able to determine the amount of $\mathrm{Ba^{2+}}$ in the original solution. A successful precipitation reaction would need to proceed to completion in order to precipitate the complete amount of $\mathrm{Ba^{2+}}$ ions in the solution.

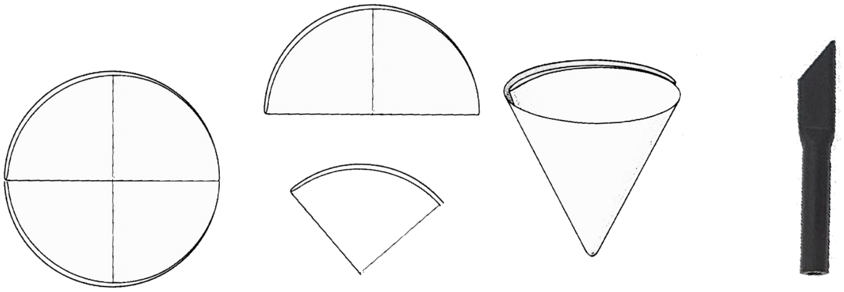

Figure 14: (Left panel) How to properly fold the filter. (Right panel) Picture of a rubber policeman, which is a hand-held flexible natural-rubber scraper attached to a glass rod used in chemical laboratories to transfer residues of precipitate or solid on glass surfaces when performing gravimetric analysis.

References

(1) pubchem., `https://pubchem.ncbi.nlm.nih.gov`, Accessed: 2017-04-22.

(2) www.homedepot.com/., `http://www.homedepot.com/`, Accessed: 2017-04-22.

(3) Wagner, W. F.; Wuellner, J. A. *Analytical Chemistry* **1952**, *6*, 1031–1032.

(4) Notley, J. M. *Journal of Applied Chemistry and Biotechnology* **1973**, *10*, 717–723.

Procedure

Initiating the Reaction

☐ *Step 1:* – You will need 1.2 g to 1.5 g of $Ba(NO_3)_2$ and 2.3 g to 2.5 g NH_2SO_3H. Use separate weighting boats to get approximate amounts of the 2 reactants.

☐ *Step 2:* – Weight a clean, dry, 250 mL beaker and record the exact mass in the Results section.

☐ *Step 3:* – Transfer the $Ba(NO_3)_2$ samples to a 250 mL beaker and record the mass of the beaker with the compound in the Results section.

☐ *Step 4:* – Transfer the NH_2SO_3H samples to the beaker and record the mass of the beaker with the two compounds in the Results section.

☐ *Step 5:* – Add 150 mL of distilled water and mix with a glass stirring rod.

☐ *Step 6:* – Prepare a stand with two metal rings and a wire gauze. The second metal ring will serve as a fence to avoid the beaker falling. Get a bunsen Burner ready. Set the height of the metal rings according to the size of the flame. It is important that the size of the flame can be reduced using the burner valve.

☐ *Step 7:* – If your instructor want you to use a hot place you will only need a single metal ring. Mind the beaker might jump if you use too much heat. If this happens reduce the heat immediately.

☐ *Step 8:* – On a different spot of your lab table, prepare an smaller metal ring with a funnel for the filtration.

☐ *Step 9:* – Place the solution on the wire gauze. Leave the stirring rod inside. Bring the water to boil and reduce the size of the flame when it starts boiling to avoid spills. Allow the solution to boil for about 30 min. Make sure the amount of liquid ($\approx$ 150 mL) does not change, by adding extra water with the wash bottle. Stir the solution occasionally.

⚠CAUTION!

⚠ Barium solutions are toxic. Make sure you to wash your hands thoroughly after working with Barium compounds.

⚠ If your beaker starts jumping when using a hot place, immediately reduce the heat.

Filtering and weighting of the dry precipitate

☐ *Step 10:* – After 30 minutes boiling, turn off the Bunsen Burner and cool the beaker to room temperature. Do not move the beaker until you can handle it with your hand without burning.

☐ *Step 11:* – Obtain a piece of filter paper and record its mass.

☐ *Step 12:* – When the reaction mixture containing the precipitate is cooled, proceed to filter the mixture. Make sure no trace of precipitate remains in the beaker. Use a wash bottle to rinse any product left.

☐ *Step 13:* – Transfer the filter paper and its contents to a watch glass labeled with your name on it. Make sure the precipitate do not touch the glass.

☐ *Step 14:* – Place the beaker in an oven at 90°C for at least 1 hour, until the paper is fully dry. Make sure you use gloved to protect yourself from the heat.

☐ *Step 15:* – Remove the dry filter paper with dry precipitate from the oven after 1 hour, always using gloves. Make sure the filter paper is completely dry. Leave the watch glass in the oven, so that you can place the paper with the precipitate directly in the tare balance using a cool watch glass. Record the mass of the filter paper and its content.

☐ *Step 16:* – Calculate the % Yield.

Calculations

① Measured using the balance.

② Measured using the balance.

③ Measured using the balance.

④ Calculate mass by subtraction ②–①

⑤ Calculate mass by subtraction ③–②

⑥ Measured using the balance.

⑦ Measured using the balance.

⑧ Calculate mass by subtraction ⑦–⑥

⑨ Convert ④ using the reciprocal of the molar mass of the corresponding chemical compound.

⑩ Convert ⑤ using the reciprocal of the molar mass of the corresponding chemical compound.

⑪ Using the balanced chemical equation and the number of moles calculated in ⑨ and ⑩, calculate the limiting reactant as in the example and in the pre-lab.

⑫ Convert moles to moles, and to mass: ⑨ × 261.337

⑬ % Yield = ⑧ / ⑫ × 100%

STUDENT INFO

Name: ________________ Date: __________

Pre-lab Questions

Gravimetry

---

1. Explain what to do if you warm up a beaker filled with liquid using a hot plate and it starts to jump.

2. For the hydrolysis of sulfamic acid NH_2SO_3H (MW 97.1 $g \cdot mol^{-1}$) reacting with $Ba(NO_3)$ (MW 261.34 $g \cdot mol^{-1}$) to produce $BaSO_4$ (MW 233.38 $g \cdot mol^{-1}$), we use a solution containing 1.50 g $Ba(NO_3)_2$ and 2.50 g of NH_2SO_3H :

$$Ba(NO_3)_2(aq) \;+\; NH_2SO_3H(aq) \;+\; H_2O(l) \longrightarrow BaSO_4(s) \;+\; NH_4NO_3(aq) \;+\; HNO_3(aq)$$

   (a) Use the expression below and calculate the grams of sulfamic acid needed to react with 1.50 g of $Ba(NO_3)_2$:

$$\text{grams } NH_2SO_3H = 1.50\text{grams of } BaNO_3 \times \frac{\underline{\qquad}\,\text{mol } BaNO_3}{\underline{\qquad}\,\text{g } BaNO_3} \times \frac{\underline{\qquad}\,\text{mol } NH_2SO_3H}{\underline{\qquad}\,\text{mol } BaNO_3}$$
$$\times \frac{\underline{\qquad}\,\text{g } NH_2SO_3H}{\underline{\qquad}\,\text{mol } NH_2SO_3H} = \underline{\qquad}\,\text{g } NH_2SO_3H \text{ needed to react}$$

   (b) Identify how many grams of sulfamic acid you actually have and see if you have more or less sulfamic acid than the amount you need. If you have more than you need, then $BaNO_3$ is the limiting reactant. If you do not have enough, then sulfamic acid should be the limiting reactant.

   (c) Assuming $BaNO_3$ is the limiting reactant, calculate the grams of $BaSO_4$ produced using the following expression:

$$1.5 \text{ grams of } BaNO_3 \times \frac{\underline{\qquad}\,\text{mol } BaNO_3}{\underline{\qquad}\,\text{g } BaNO_3} \times \frac{\underline{\qquad}\,\text{mol } BaSO_4}{\underline{\qquad}\,\text{mol } BaNO_3} \times \frac{\underline{\qquad}\,\text{g } BaSO_4}{\underline{\qquad}\,\text{mol } BaSO_4} =$$
$$= \underline{\qquad}\,\text{g } BaSO_4 \text{ produced}$$

   (d) Knowing that the chemical reaction produces 1.0 g of $BaSO_4$. Calculate the % yield of the reaction.

STUDENT INFO

Name: ____________________ Date: __________

Results

EXPERIMENT

Gravimetry

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①	Mass of empty, clean, dry 250 mL beaker (g)	
②	Mass of beaker + $Ba(NO_3)_2$ (g)	
③	Mass of beaker + $Ba(NO_3)_2$ + NH_2SO_3H (g)	
④	Mass of $Ba(NO_3)_2$ (g)	
⑤	Mass of NH_2SO_3H (g)	
⑥	Mass of dry, unused filter paper (g)	
⑦	Mass of filter paper + precipitate after oven (g)	
⑧	Mass of dry precipitate (g)	
⑨	Moles of $Ba(NO_3)_2$ (g)	
⑩	Moles of NH_2SO_3H (g)	
⑪	Limiting reactant	
⑫	Theoretical mass of precipitate	
⑬	$\%\ \text{Yield}=\frac{\text{Actual mass of precipitate}}{\text{Theoretical mass of precipitate}} \times 100\%$	

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Gravimetry

---

1. For the reaction of NaOH (M_W 39.997 $g \cdot mol^{-1}$) with $MgCl_2$ (M_W 95.221 $g \cdot mol^{-1}$) to produce $Mg(OH)_2$ (M_W 53.320 $g \cdot mol^{-1}$), we use a solution containing 1.00 g NaOH and 2.00 g of $MgCl_2$:

$$2\,\text{NaOH}\,(aq) \quad + \quad \text{MgCl}_2\,(aq) \quad \longrightarrow \quad \text{Mg(OH)}_2\,(s) \quad + \quad 2\,\text{NaCl}\,(aq)$$

   (a) Use the expression below and calculate the grams of $MgCl_2$ needed to react with 1.00 g of NaOH :

$$\text{grams MgCl}_2 = 1.00 \text{ grams of NaOH} \times \frac{\underline{\hspace{2cm}}\,\text{mol NaOH}}{\underline{\hspace{2cm}}\,\text{g NaOH}} \times \frac{\underline{\hspace{2cm}}\,\text{mol MgCl}_2}{\underline{\hspace{2cm}}\,\text{mol NaOH}}$$
$$\times \frac{\underline{\hspace{2cm}}\,\text{g MgCl}_2}{\underline{\hspace{2cm}}\,\text{mol MgCl}_2} = \underline{\hspace{2cm}}\ \text{g MgCl}_2 \text{ needed to react}$$

   (b) Identify how many grams of $MgCl_2$ you actually have and see if you have more or less $MgCl_2$ than the amount you need. If you have enough, then NaOH is the limiting reactant. If you do not have enough, then $MgCl_2$ should be the limiting reactant.

   (c) Assuming NaOH is the limiting reactant, calculate the grams of $Mg(OH)_2$ produced using the following expression:

$$1 \text{ grams of NaOH} \times \frac{\underline{\hspace{2cm}}\,\text{mol NaOH}}{\underline{\hspace{2cm}}\,\text{g NaOH}} \times \frac{\underline{\hspace{2cm}}\,\text{mol Mg(OH)}_2}{\underline{\hspace{2cm}}\,\text{mol NaOH}} \times \frac{\underline{\hspace{2cm}}\,\text{g Mg(OH)}_2}{\underline{\hspace{2cm}}\,\text{mol Mg(OH)}_2} =$$
$$= \underline{\hspace{2cm}}\ \text{g Mg(OH)}_2 \text{ produced}$$

   (d) Knowing that the chemical reaction produces 1.00 g of $Mg(OH)_2$. Calculate the % yield of the reaction.

EXPERIMENT

Heat of neutralization

Goal

The goal of this experiment is to measure the heat involved in two acid-base chemical reactions: the neutralization of NaOH and NH_3 with HCl. You will determine the enthalpy change for the two acid-base neutralization reactions using a coffee-cup calorimeter and a thermometer. By measuring the heat released or consumed you will be able to calculate the enthalpy change ΔH for both reactions.

Materials

☐ A coffee-cup calorimeter

☐ A thermometer and a lit for the calorimeter

☐ HCl (aq), NH_3 (aq) and NaOH (aq) solutions

Background

Reactants and products in a chemical reaction normally have different energy, and this difference in energy is reflected by a release or absorption of heat. By means of a coffee-cup calorimeter–a constant pressure calorimeter–and a thermometer, we can measure temperature changes produced by a chemical reaction. When a reaction produces heat, the temperature inside the calorimeter will rise, whereas the temperature will decrease if the reaction consumes heat. These energy changes are related to the enthalpy change–the heat exchanges at constant pressure–of the chemical reaction. This is a classical college chemistry experiment implemented in many lab manuals with different choices of acids and bases [1–4].

Measurement of heat

Heat–thermal energy q–is the energy transferred as a result of a difference in temperature. q flows from regions of high-temperature to low-temperature regions. Heat is connected to change on temperature ΔT by means of the specific heat C:

$$q = m \cdot C \cdot \Delta T$$

where m is the mass of the object. Large heat capacity means that a small supply of heat produces a big temperature increase. Metals for example have small specific heats. For aqueous solutions C is the specific heat of water, $4.184\, J \cdot ^{\circ} C^{-1} \cdot g^{-1}$, and assuming the density of the solution is $1.00\, g \cdot mL^{-1}$ the energy change due to a change in temperature $\Delta T = T_f - T_i$ is given by:

$$q = \left(d \cdot V\right) \cdot C \cdot \Delta T$$

where V is the volume of solution in mL and the temperatures are in $^{\circ}C$.

Enthalpy change of a chemical reaction

Learn more about heat and reactions in this 🌍 Video Link

Visual

The energy change in a chemical reaction measured at constant pressure–in an open container–is called enthalpy change, ΔH. All chemical reactions are accompanied by a change in energy. For example, the reaction between an acid and a base is called neutralization, and for the case of HCl and NaOH is the following reaction:

$$\mathrm{HCl} + \mathrm{NaOH} \longrightarrow \mathrm{NaCl} + \mathrm{H_2O} \qquad \Delta H_1 = -52\ KJ$$

The heat involved in this process is -52 kJ. The negative sign means that the heat is being released. Reactions that release heat are called exothermic ($\Delta H < 0$), whereas reaction that absorb heat are called endothermic ($\Delta H > 0$). The heat released in a chemical reaction depends on the external conditions. The standard reaction enthalpy is the reaction enthalpy when reactants in their standard states change into products in their standard states, that is its pure form at exactly 1 bar. The enthalpy of an element on its standard state is always zero. For example for H_2(g) $\Delta H = 0$.

Using a coffee-cup calorimeter

Two nested polystyrene cups–coffee-cups–make a good constant–pressure calorimeter, as polystyrene is a good heat insulator. With a coffee-cup calorimeter and a thermometer, we can measure the enthalpy change ΔH of a chemical reactions happening inside the calorimeter: the heat q consumed or released inside a calorimeter is related to the enthalpy change of the reaction by:

$$\Delta H_r + q = 0$$

Often times heat flows through the walls of calorimeters, as these devices are never perfect. To take into account this loose of heat, we can use the following formula:

$$\Delta H_r + q + C_{Cal}\Delta T = 0$$

where C_{Cal} is the calorimeter constant and it is calculated to be $C_{Cal} = 10.0\,J\cdot^{\circ}C^{-1}$ for a standard calorimeter. It accounts for the energy absorbed by the calorimeter. Because of this effect, we will not be able to observe the highest (or lowest) temperature that could have been achieved in a perfectly insulated calorimeter, as when the temperature in a calorimeter starts rising due to a reaction, heat starts leaving the cup. We can compensate this effect and estimate the temperature when the heat leakage is minimal, by plotting temperature as a functions of time and then extrapolating the values to the beginning of the reaction.

Example

When 10.00 ml of nitric acid 1.0 M, HNO_3, reacts with excess of NaOH:
$HNO_3(aq) + NaOH(aq) \longrightarrow NaNO_3(aq) + H_2O(aq)$
in a coffee-cup calorimeter with a heat capacity of 10. $J\cdot^{\circ}C^{-1}$, the temperature of the calorimeter rises by 10.0° C. Calculate the reaction enthalpy for this process assuming the density of the mixture is $1.00\frac{g}{mL}$.

Answer: the heat absorbed by the calorimeter is the heat exchanged during the chemical reaction:

$$\Delta H_r + \left(d\cdot V\right)\cdot C\cdot\Delta T + C_{Cal}\cdot\Delta T = 0$$

where V is the volume used of nitric acid. Plugging all the values:

$$\Delta H_r + \left(1.00\,g\cdot mL^{-1}\cdot 10.00\,mL\right)\cdot 4.184\,J\cdot^{\circ}C^{-1}\cdot g^{-1}\cdot 10.0^{\circ}C^{-1} + 10.J\cdot^{\circ}C^{-1}\cdot 10.0^{\circ}C = 0$$

Solving for ΔH_r we have $-520\,J$. The number of moles of HNO_3 is:

$$n = 10.00\,mL \times \frac{1.0\,mol}{L} = 0.010\,mol$$

hence $\Delta H_r = -5,200\,J/mol = -52\,kJ/mol$

Hess's Law

Hess's law states that the total enthalpy change during a chemical reaction is the same whether the reaction results of one or several steps. For example:

$$
\begin{array}{rcll}
\mathrm{HCl} + \mathrm{NaOH} & \longrightarrow & \mathrm{NaCl} + \mathrm{H_2O} & \Delta H_1 \\
\mathrm{NH_3} + \mathrm{HCl} & \longrightarrow & \mathrm{NH_4Cl} & \Delta H_2 \\
\hline
\mathrm{NaOH} + \mathrm{NH_4Cl} & \longrightarrow & \mathrm{NaCl} + \mathrm{NH_3} + \mathrm{H_2O} & \Delta H_3
\end{array}
$$

You might have noted there is a relationship among these equations: if you reverse the second equation and add it to the first equation, the third equation is generated. This provided the basis for Hess's law and you will find you can predict the enthalpy change for the third reaction by combining the enthalpy change for the first and second reactions:

$$\Delta H_3 = \Delta H_1 - \Delta H_2$$

Example

Calculate the standard enthalpy of the reaction $3\,C(gr) + 4\,H_2(g) \longrightarrow C_3H_8(g)$, ΔH°_R, *using the following data:*

$$
\begin{array}{rcll}
C_3H_8(g) + 5\,O_2(g) & \longrightarrow & 3\,CO_2(g) + 4\,H_2O(l) & \Delta H^\circ_1 = -2220.\,kJ \\
C(g) + O_2(g) & \longrightarrow & CO_2(g) & \Delta H^\circ_2 = -394\,kJ \\
H_2(g) + \frac{1}{2}O_2(g) & \longrightarrow & H_2O(l) & \Delta H^\circ_3 = -286\,kJ
\end{array}
$$

Answer: Using Hess's law, we can build up the reaction by reversing the first reaction, multiplying the second reaction times 3, multiplying reaction 3 times 4 and adding them. If we do the same to the enthalpies we obtain:

$$\Delta H^\circ_R = 3 \cdot \Delta H^\circ_2 + 4 \cdot \Delta H^\circ_3 - \Delta H^\circ_1 = -106\,kJ$$

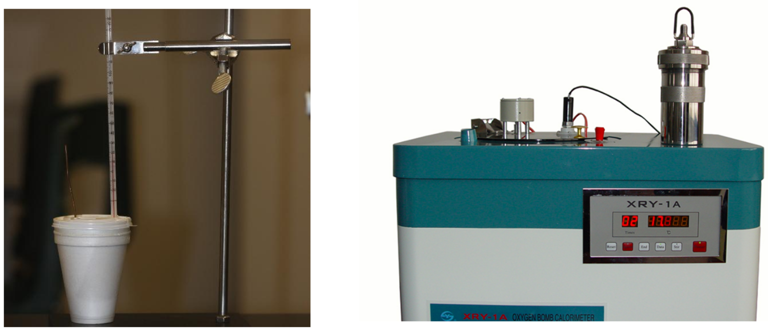


Figure 15: (Left panel) A coffee-cup calorimeter. (Right panel) A constant-volume calorimeter.

How to find extrapolated temperature

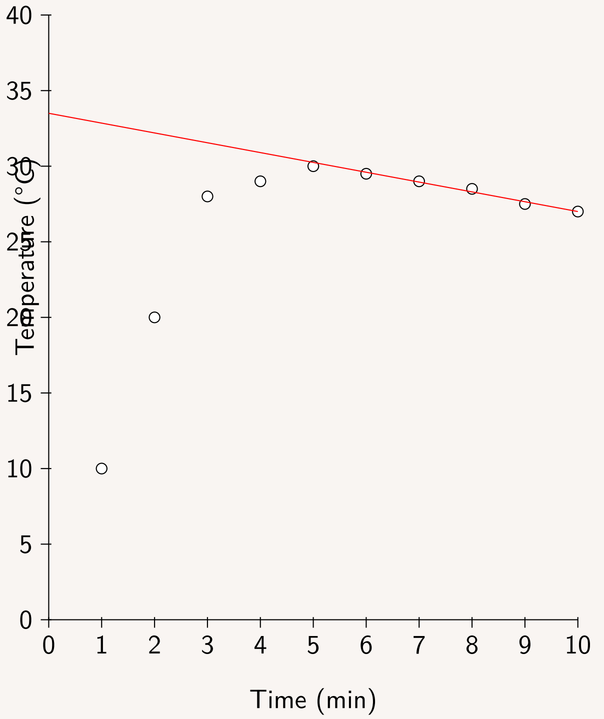
A calorimeter is ideally an isolated system. That means, the heat produced does not leave the cup. If a reaction is exothermic–if it produces heat–the temperature inside the calorimeter will increase. Differently, for endothermic reaction–a reaction that consumes heat–the temperature inside the calorimeter will decrease. When the reaction stops, as there is no more heat being produced or consumer, the temperature inside the calorimeter will change. In order to estimate the temperature at that point, we need to extrapolate the temperature inside the calorimeter. If you plot temperature versus time you will see that for early times, the temperature inside a calorimeter in which a reaction happens, changes abruptly. After that, the temperature slowly decreases with time. In order to calculate the temperature when the reaction stops you need to connect the dots of the later stages, when the reaction already has stopped. We call this interpolation. All these points will be part of a line. If you extend the line, it will cross the Y axis, and that crossing point is the temperature at what the reaction has stopped. We call this the extrapolation method.

Example

The following data reports the temperature inside a cup calorimeter. Indicate whether the reaction is exothermic or endothermic and calculate the temperature when the reaction stops by means of the extrapolation method.

t (min)	1	2	3	4	5	6	7	8	9	10
T (°C)	10	20	28	29	30	29.5	29	28.5	27.5	27

Answer: if we plot temperature versus time we will see the two different regions of the plot. For early times, temperature increase and hence the reaction in the calorimeter is exothermic. For later times, the reaction stops and temperature decreases due to the lack of heat release. We will use the last points and connect them with a line. The crossing point with the Y axis is 33 °C and that is the final temperature of the reaction.



Procedure

Measuring ΔH for the HCl–NaOH pair

☐ *Step 1:* – Obtain a coffee-cup calorimeter with a thermometer and a plastic lid.

☐ *Step 2:* – Place 50.0 mL of HCl and 50.0 ml of NaOH in separate graduated cylinders. Make sure the cylinders are dry and clean.

☐ *Step 3:* – Write down the concentration of the acid and the base on the results section.

☐ *Step 4:* – Measure the temperature of each of these solution. If you use the same thermometer for both solutions, make sure your clean the thermometer in between measurements as the chemicals will reacts. Record the mean of the two temperatures as your initial temperature t_i.

☐ *Step 5:* – Get the timer ready. Time should be taken as soon as the second reactant is added.

☐ *Step 6:* – Add the acid to the calorimeter. If you are using a magnetic stirrer set the speed to a medium speed. Then add the base. Start the timer immediately and place the lid on the calorimeter.

☐ *Step 7:* – Record the temperature for 4 min. Do not stop the timer during the experiment. Record the results in the results table.

☐ *Step 8:* – Plot the temperature of the calorimeter against time and use a straight line to extrapolate your results against the starting time (time = 0 s). Record the extrapolated temperature as t_f.

☐ *Step 9:* – Calculate the change of enthalpy, ΔH.

Good Lab Practice

✍ Make sure you always use the same thermometer all through the experiment so that the results are consistent. ✍

Measuring ΔH for the HCl – NH_3 pair

☐ *Step 10:* – Obtain a coffee-cup calorimeter with a thermometer and a plastic lid.

☐ *Step 11:* – Place 50.0 mL of HCl and 50.0 ml of NH_3 in separate graduated cylinders. Make sure the cylinders are dry and clean.

☐ *Step 12:* – Measure the temperature of each of these solution. If you use the same thermometer for both solutions, make sure your clean the thermometer in between measurements as the chemicals will reacts. Record the mean of the two temperatures as your initial temperature t_i.

☐ *Step 13:* – Get the timer ready. Time should be taken as soon as the second reactant is added.

☐ *Step 14:* – Add the acid, HCl to the calorimeter. If you are using a magnetic stirrer set the speed to a medium speed. Then add the base, NH_3. Start the timer immediately and place the lid on the calorimeter.

☐ *Step 15:* – Record the temperature for 4 min. Do not stop the timer during the experiment. Record the results in the results table.

☐ *Step 16:* – Plot the temperature of the calorimeter against time and use a straight line to extrapolate your results against the starting time (time = 0 s). Record the extrapolated temperature as t_f.

☐ *Step 17:* – Calculate the change of enthalpy, ΔH.

⚠CAUTION!

⚠ All acids and bases in this experiment, HCl, NaOH and NH_3, can cause chemical burns. !

Calculations

① This is the initial temperature before you mix the reagents.

② This is the final temperature after the reaction stops. You need to plot the data and extrapolate to get this value.

③ This is the number of moles reacting: $c_{acid} * 50/1000$

④ You can calculate the enthalpy change (ΔH) by means of the following formula:

$$\Delta H = -\frac{1}{\textcircled{3}} \cdot \left(4.184 \cdot 1.0 \cdot 100.0 \cdot (\textcircled{2} - \textcircled{1}) + 10 \cdot (\textcircled{2} - \textcircled{1})\right)$$

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Heat of neutralization

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1. An acid and a base are mixed in a coffee-cup calorimeter. We recorded the temperatures with time and the data is given below. The initial temperature, t_i was 25.0 °C.

t (s)	30	60	90	120	150	180	210	240
T (°C)	40.2	40.6	40.5	40.4	40.3	40.2	40.1	40.0

   (a) Plot T vs. t. Obtain t_f, the final temperature after the reaction starts, by extrapolating to the time of mixing (time = 0 s) with a straight line.

   t_f = ____________________

   (b) Indicate whether this is an exothermic or an endothermic reaction.

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(c) Explain the rise and fall of temperature inside the calorimeter.

STUDENT INFO

Name: ____________ Date: ________

Heat of neutralization

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c_{acid} = ____________

c_{base} = ____________

HCl–NaOH Trial 1														
t (s)	0	30	60	90	120	150	180	210	240		T_i (°C) ①	T_f (°C) ②	n (moles) ③	$\Delta H (kJ/mol)$ ④
T (°C)														
HCl–NaOH Trial 2														
t (s)	0	30	60	90	120	150	180	210	240		T_i (°C)	T_f (°C)	n (moles)	$\Delta H (kJ/mol)$
T (°C)														
											Average ΔH (kJ/mol) =			

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c_{acid} = ______________ c_{base} = ______________

HCl−NH_3 Trial 1														
t (s)	0	30	60	90	120	150	180	210	240		T_i (°C) ①	T_f (°C) ②	n (moles) ③	$\Delta H (kJ/mol)$ ④
T (°C)														
HCl−NH_3 Trial 2														
t (s)	0	30	60	90	120	150	180	210	240		T_i (°C)	T_f (°C)	n (moles)	$\Delta H (kJ/mol)$
T (°C)														
											Average ΔH (kJ/mol) =			

Figure 16: HCl – NaOH

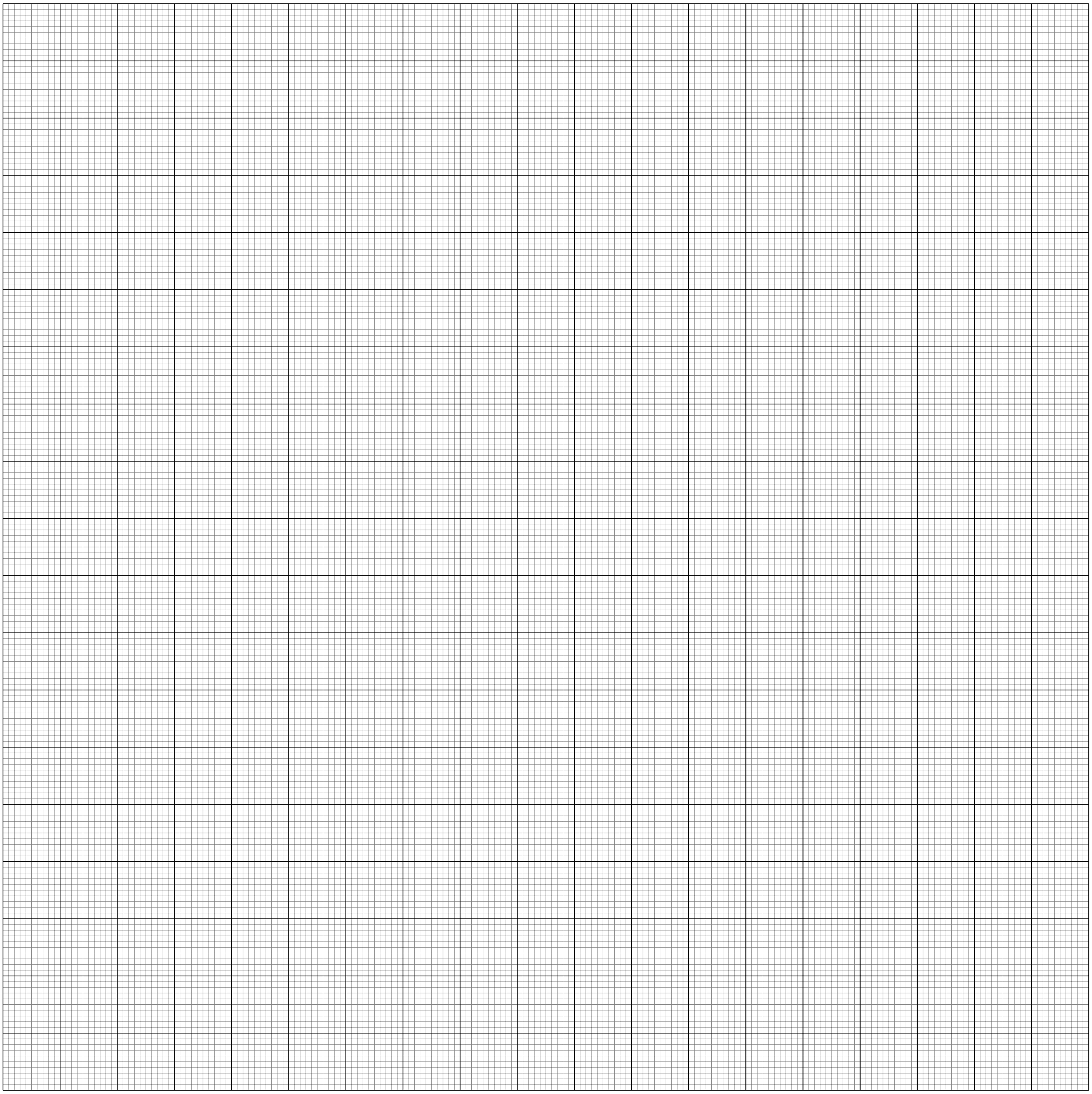

Figure 17: HCl–NH_3

STUDENT INFO

Name: | Date:

Post-lab Questions

Heat of neutralization

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1. 5.0 ml of a 1 M acid reacts with 5.0 ml of 1 M base in a coffee-cup calorimeter with a heat capacity of 10. $J \cdot (°C)^{-1}$. Calculate the rise in temperature of the calorimeter assuming the density of the mixture is 1 $\frac{g}{ml}$ and given that $\Delta H_r = 50.\, kJ/mol$.

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2. When 5.0 mL of 1.00 M HNO_3 solution at 25.0 °C reacts with 5.0 mL of 1.00 M NaOH solution at the same temperature in a calorimeter, the temperature of the resulting mixture increases to 31.8 °C. Assuming that the calorimeter absorbs only a negligible quantity of heat, that the specific heat capacity of the solution is 4.18 $J \cdot (°C)^{-1} \cdot g^{-1}$, and that the density of the final solution is 1.0 $g \cdot mL^{-1}$. Calculate the enthalpy of neutralization.

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3. Calculate the energy to increase the temperature of a 5 g piece of iron from 25.0 °C to 75.0 °C, knowing the specific heat capacity of iron is 0.45 $J \cdot (°C)^{-1} \cdot g^{-1}$.

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EXPERIMENT

Molecular Mass of a Volatile Liquid

Goal

The goal of this experiment is to calculate the molar mass (molecular weight, MW) of a volatile gas using the Dumas method, by means of the formula of the ideal gases and by measuring the mass of gas.

Materials

☐ An unknown liquid
☐ A 600 mL beaker
☐ A 125 mL Erlenmeyer
☐ A 10 and 100 mL graduated cylinders
☐ Hot plate
☐ Thermometer
☐ Stand and iron ring
☐ Rubber band and aluminum foil

Background

Gases have a very different behavior than liquids or solids. In a gas, particles (atoms, or molecules) are far apart from each other, while in solids or liquid particles are closer to each other and interact by means of stronger forces. As a consequence, gases are less dense–have less particles per volume unit–than liquids or solids and are easy to compress as well. Think for example of how easy is to compress a ballon full if air, even with your own hands. In this process the gas particles, which are far away from each other, get closer together. Another characteristic exclusive of gases is that they will fill the volume of its container. All these properties of gases are explained by the kinetic molecular theory of gases.

Gases or vapors can be easily produced by boiling a liquid. But even below their boiling temperature liquids can generate gas; a property called vapor pressure. The molecules in the exterior layer of the liquid are able to escape into the air creating a vapor near the surface of the liquid, and this vapor exert pressure. The warmer a liquid is, the more molecules are able to jump into the gas and hence the vapor pressure increases. Liquids that evaporate at low temperatures are called *volatile*, and have a high vapor pressure. In this experiment we will create a gas by boiling a small amount of a volatile liquid.

Ideal gas law

In an ideal gas particles do not interact at all. In real gases particles do actually interact by means of attractive forces. A lot of gases act as an ideal gas at high temperatures and low pressures. On the other hand, the state of a gas is determined by a set of properties: pressure, volume, number of moles and temperature. The volume (V) of a gas is the volume of its container, as gases occupy the whole container. Pressure (P), refers to the force than the gas particles exert on the walls of the container. Gas particles move fast and randomly, changing its direction after a collision with another particle or with the walls of the container. The stronger the collisions and the higher the frequency of collision, the higher the pressure. The number of moles of a gas (n), is the amount of particles. Finally, temperature (T) is related to the kinetic energy of the gas molecules. The higher the temperature the faster the average speed of the gas particles. All the variables that define the state of an gas are related by means of the ideal gas law:

$$\boxed{PV = nRT} \quad \text{Ideal Gas Law}$$

Pressure units

Gas pressure can be expressed in several units. Units of pressure are: atmospheres (atm), torr, pascals (Pa), millimeters of mercury (mmHg) and in inHg. In order to covert pressure units, you can use the following conversion factors:

$$\boxed{\frac{1\text{ atm}}{760\text{ mmHg}}}\quad\boxed{\frac{1\text{ torr}}{1\text{ mmHg}}}\quad\boxed{\frac{1\text{ atm}}{101325\text{ Pa}}}\quad\boxed{\frac{1\text{ atm}}{29.92\text{ inHg}}}$$

The atmospheric pressure is normally close to one atm. However, its exact value depends on factors like altitude of even the weather. The higher you are with respect to the sea level the lower pressure. Similarly, the atmospheric pressure on a hot day is higher than on a cold day.

Example

The atmospheric pressure on a hot summer day in NYC is 29.84 inHg. Calculate this value in atm and mmHg.

Answer: In order to convert from inHg to atm we will use the conversion factor: 1 atm = 29.92 inHg and start by using the given pressure and the conversion factor with the inHg part on the bottom

$$29.84\,\cancel{\text{inHg}} \times \frac{1\text{ atm}}{29.92\,\cancel{\text{inHg}}} = 0.9973\text{atm}$$

Now, we will convert this value into mmHg using 1 atm = 760 mmHg

$$0.9973\,\cancel{\text{atm}} \times \frac{760\text{ mmHg}}{1\,\cancel{\text{atm}}} = 757.9\text{mmHg}$$

Calculating the molar weight of a volatile gas: the Dumas method

The procedure for calculating the molar weight (or molar mass) of a volatile gas–also called the Dumas method–consists of boiling completely a small amount of the liquid, so that the gas formed will occupy the volume on the container. In this conditions, you will know the volume of the gas (same as the container) and the temperature, which can be measured with a thermometer. The pressure of the gas will be the atmospheric pressure, as when a liquid boils its vapor pressure equal the external pressure. Using all these variables (P, V, T) we will be able to calculate the number of moles of the gas. At a given point we will stop heating the container and the vapor contained in the flask will condensate back into a liquid. Knowing the mass of the empty container and the mass of the container with condensate (the liquid) we will be able to measure the mass of the gas. Finally, the molar mass of a gas is the result of dividing its mass and its number of moles.

Example

We set up the experiment to calculate the molar mass of a volatile liquid. The mass of the dry flask is 90.00 g and after the experiment is done, the wet flask with the condensate liquid weights 92.00 g. The volume of the flask is 150.0 mL. The temperature of the hot bath is 98.0° C and the atmospheric pressure is 0.98 atm. Calculate the molar mass of the substance.

Answer: In order to calculate the molar mass of the liquid we need the mass of the gas in the experiment and the number of moles. The mass of gas can be obtained from the difference in mass between the wet and dry beaker, as the condensate liquid comes from the gas in the flask, and hence the mass of this liquid is indeed the mass of gas in the experiment. The value is 2.00 grams. We will obtain the number of moles of gas using the ideal gas law. The pressure of the gas equals to the atmospheric pressure. This is because when a liquid boils the vapor pressure equals to the atmospheric pressure: $P = 0.98\,atm$. On the other hand the temperature of the gas is the temperature of the hot bath: $T = 98.0 + 273 = 371\,K$. The volume of the gas is the volume of the flask: $V = 0.150\,L$. Remember in order to use the ideal gas law we need to use atm, L and Kelvins. Solve the ideal gas law for the n.

$$PV = nRT \qquad n = \frac{PV}{RT}$$

Plugging the values in the law:

$$n = \frac{0.98\cancel{atm}\cdot 0.150\cancel{L}}{0.082\frac{\cancel{atm\cdot L}}{mol\cdot\cancel{K}}\cdot 371\cancel{K}} = 4.83\cdot 10^{-3}\,moles$$

Solving for the number of moles: n=$4.83\cdot 10^{-3}$ moles, and using the mass we can calculate the molar mass of the liquid:

$$\text{MW} = \frac{2.00\text{ grams}}{4.83\cdot 10^{-3}\text{ moles}} = 414\,g\cdot mol^{-1}$$

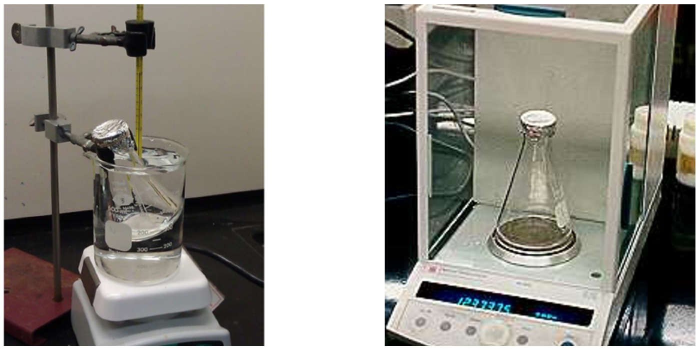

Figure 18: (Left panel) Set up for the experiment. (Right panel) A scale with the empty flask covered with aluminum foil and a rubber band.

Procedure

Generating the gas

☐ *Step 1:* – First, Prepare a hot water bath, using a 600 mL beaker, large enough to comfortably fit a 125 mL flask, and a hot plate. Add 200 mL of water to the beaker, place it on the hotplate and start heating.

☐ *Step 2:* – Weight and record the mass of a dry 125 mL Erlenmeyer, covered with a small square of aluminum foil and fastened using a rubber band.

☐ *Step 3:* – The instructor will give you an unknown liquid. While you wait for the water to boil add approximately 2 mL of the unknown liquid into the Erlenmeyer flask.

☐ *Step 4:* – Place the aluminum foil in the mouth of the flask and fasten it with the rubber band. After that, make one very tinny hole on the aluminum foil cover, enough for excess vapor to scape while the liquid boils.

☐ *Step 5:* – When the water in the hot bath is boiling, place the Erlenmeyer flask inside the boiling water, close to the bottom of the beaker. Clamp the flask so that it does not move and twist slightly the flask so that you can clearly see the level of the liquid in the flask. Wait until the unknown liquid has completely disappeared from the flask, becoming gas, and continue boiling the system for 10 minutes. At this point you already have a gas inside the flask, at a certain temperature (that of the water) and pressure (the atmospheric pressure).

Measuring T and P

☐ *Step 6:* – While the gas is in the Erlenmeyer, measure its temperature by measuring the temperature of the boiling water. Record this measurement in the results and make sure you write down this value in Kelvins.

☐ *Step 7:* – Measure the pressure of the gas–remember this is equal to the atmospheric pressure–with a barometer in the lab. Record this measurement in the results, and make sure you write down this results in atmospheres.

☐ *Step 8:* – After 10 minutes, stop the hot plate and remove the beaker from the bath. Wait patiently until the flask is completely cool. At this point measure the mass of the flask, with the rubber band and aluminum foil and the condensed liquid. At this point you will have the mass of the gas, which is the same as the weight of the condensed liquid.

Calculating the volume of an Erlenmeyer

☐ *Step 9:* – Calculate the volume of the Erlenmeyer by filling the flask completely with water, all the way to the brim. Use a graduated cylinder larger enough to accurately measure the volume of liquid in the flask. Help yourself with a plastic pipet to transfer all the water without spilling one drop. The volume should be more than 125 mL.

☐ *Step 10:* – Write down the volume of the Erlenmeyer in the results section, and make sure you write the result in liters. At this point you will have the volume of gas.

☐ *Step 11:* – Use the ideal gas law to calculate the number of moles of gas. Use the mass measurement to calculate the molar mass of the unknown gas.

⚠CAUTION!

⚠ Be aware that hot glass looks like cold glass.

⚠ All chemicals used must be disposed properly.

Calculations

① Measured using the balance.

② The measured temperature of the water boiling, measures using a thermometer.

③ Measured using the barometer.

④ Measured using the balance.

⑤ Calculate mass by subtraction ④–①

⑥ Measured with a graduated cylinder.

⑦ Calculate using the Ideal Gas Law.

$$n = \frac{\textcircled{3} \cdot \textcircled{6}}{0.082\,\frac{atm \cdot L}{mol \cdot K} \cdot \textcircled{2}}$$

⑧ $MW = \frac{\textcircled{5}}{\textcircled{7}}$

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Molecular Mass of a Volatile Liquid

---

1. The atmospheric pressure of a hot summer day in NYC is 29.94 inHg. Calculate this value in mmHg and atm.

2. The atmospheric pressure of a cold winter day in NYC is 25.94 inHg. Calculate this value in Pa.

3. We set up an experiment utilizing the Dumas method to calculate the molar mass of an unknown chemical. A 4.00 mL pure liquid sample of this chemical is vaporized in an Erlenmeyer flask with a volume of 155 mL, when the barometric pressure is 1.01 atm. The empty flask weighs 25.2436 g. After the excess gas escapes, the temperature is measured as 97.0°C. The flask and contents are subsequently cooled and the vapor condenses to a liquid. The weight of the flask and contents is found to be 25.7843 g. Calculate the molar mass of the unknown liquid.

STUDENT INFO

Name: ____________ Date: ________

Results

EXPERIMENT

Molecular Mass of a Volatile Liquid

Unknown # = ____________

①	Mass of dry flask with foil and rubber (g)	
②	Temperature (K)	
③	Pressure (atm)	
④	Mass of wet flask with foil and rubber (g)	
⑤	Grams of gas (g)	
⑥	Volume (L)	
⑦	Moles of gas (mol)	
⑧	Molar Mass of the gas ($g \cdot mol^{-1}$)	

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Molecular Mass of a Volatile Liquid

---

1. Dumas studied mercury with the aim to estimate its molar mass. He found that at 446 °C and 765 torr, 0.812 g of mercury vapor filled a vessel of volume 0.235186 L. From this data, compute the molar mass of mercury.

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2. Calculate the mass of 45.0 L of NH_3 at 27 °C and 890 mmHg.

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3. A volume of 30.0 mL of a gas was collected in a tube at a temperature of 15°C and 800. mm Hg. The next day the volume of the same gas was 26.0 mL with the barometer still reading the same pressure. Calculate the temperature the second day.

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EXPERIMENT

Empirical rules for polarity and Paper Chromatography

Goal

This experiment will introduce you to paper chromatography, which is a standard chemical technique used to separate chemicals in a mixture. The goal of the experiment is to identify an unknown chemical in a mixture of several chemicals. You will do this by means of paper chromatography.

Materials

☐ 4in tall × 9in long piece of filter paper

☐ 8 capillary tubes

☐ A 600 mL beaker

☐ A large watch glass to be used as a lid

☐ An unknown substance, some blanks solutions and a mixture solution.

Background

Often, the chemicals in nature exist as mixtures of several compounds and one needs to figure out the components. There are many analytical techniques used to separate and identify the components of a mixture. These techniques exploit the differences in the chemical or physical properties of the components of the mixture in order to separate the different elements. Chromatography is a very sensitive as well as affordable chemical technique, extensively employed to separate mixtures of chemicals and identify the nature of the elements of the mixture. This is a classic experiment, previously implemented in numerous college chemistry laboratory manuals with different reagents [1–4].

Mixtures

A lot of the materials that you use in your everyday life are indeed mixtures of chemicals. For example, air is a mixtures of gases. Sometimes one needs to specify the composition of mixtures quantitatively, identifying each of the elements. In this experiment we will learn to separate and identify the different compounds in a mixture.

Paper Chromatography

In the following video you can learn more about 🌍 *paper chromatography*.

Visual

Paper chromatography relies on capillarity, which is the tendency of liquid substances to rise on the surface of a material. In this technique, a drop of a liquid solution containing a different substances (the sample) is deposited on a rectangular piece of filter paper, close to the bottom edge. This paper is called the stationary phase. The bottom end of the paper is immersed in a liquid called the mobile phase, to a point that is just below the spot where the sample was placed. Due to capillarity, the mobile phase will move up along the stationary phase. When the mobile phase reaches the sample, the different components of the mixture will begin migrate, carried away by the mobile phase. The chemical compounds forming the sample will move with the mobile phase, but as different chemicals have different tendency to stick to the mobile phase, they will cover different distances along the stationary phase. The different heights achieved by the different substances would allow you to identify those chemicals. A component of the mixture with high affinity to the mobile phase will migrate more than a component with a higher affinity to the stationary phase.

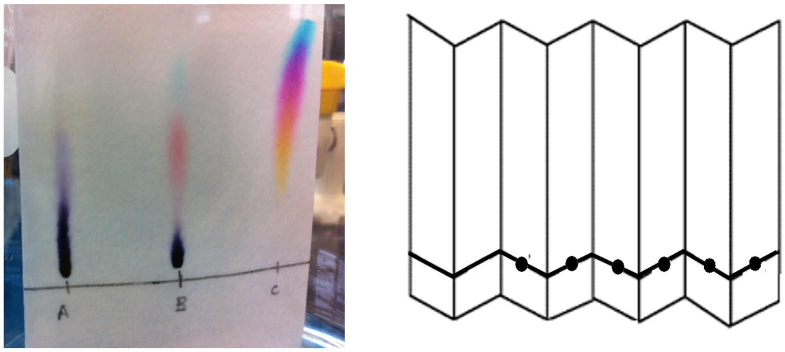


Figure 19: (Left panel) Chromatogram with several spots. (Right panel) Accordion-folded chromatography paper.

Retention factor, R_f

Each chemical will have a specific affinity towards the mobile phase, and as a consequence, it will move further in the stationary phase. The distance traveled by a component referred to the distance traveled by the mobile phase in the experiment is a measure of that affinity between the chemical and the mobile phase. We define the retention factor R_f of a given chemical as:

$$R_f = \frac{\text{distance traveled by the chemical}}{\text{total distance traveled by mobile phase}}$$

The R_f value of a substance is characteristic of that substance. When dealing with mixtures one has to calculate the R_f for each pure component separately to then compare the retention factors with the ones obtained in the mixture.

How to reveal the spots

Most of the chemicals leave no mark when being carried by the mobile phase. For those cases, one needs to reveal the spots by means of a revealing agent. The revealing agent reacts with the components of the mixture giving a colored spot. The revealing is done after the paper chromatography is finished, in order to make the spots visible. Each spot is a chemical that was carried by the mobile phase a certain distance. If the sample is a pure chemical, you will see a single spot. If the sample is a mixture of chemicals you will see one spot for each component. Often times the spots can overlap each other.

Empirical rules for polarity

The affinity between a chemical and the mobile phase is connected to a concept called polarity. Molecules contain electrons and depending on the electron distribution within the molecules, molecules can be polar or non-polar.
Molecules with an even electron distribution are non-polar. An example of this is H_2 molecule, which is non-polar. Differently, HF is a polar molecule, as F likes more the electrons than H and these will spend a longer time along fluorine (you will learn more about this effect at the end of the semester). The polar nature of substances is related to its miscibility and molecules with similar polar character will mingle and mix together creating a single visible phase. As an example, water (polar) and methanol (polar) will mix together. Differently, water (polar) and oil (non-polar) are immiscible due to its different polar nature and they will not mix. Even if the rules or polarity are based of the nature and structure of the molecule, one can use very simple empirical rules to classify molecules as polar or non-polar. These rules work in general well for the case of diatomic and very large molecules:

- Diatomic molecules made of the same element are non-polar.
- Diatomic molecules made of different elements are polar.
- Poliatomic molecules (with more than four atoms) made of C and H are in general non-polar.
- Poliatomic molecules (with more than four atoms) containing C, H and a different atom are in general polar.

Example

Classify the following molecules as polar or nonpolar: H_2, *HCl,* CH_3CH_3, *and* CH_3CH_2Cl.

Answer: H_2 is a non-polar molecule, being a diatomic molecule containing two atoms of the same element. Differently HCl is polar. CH_3CH_3 is a non-polar poliatomic molecule made of C and H atoms only, whereas CH_3CH_2Cl is polar.

References

(1) Bergstrom, B. A.E.M. A., *Chemistry in Your Life Lab Manual*; W. H. Freeman: 2006.

(2) Girard, P. B. J., *Laboratory Manual for Principles of General Chemistry*; Jones and Bartlett Learning: 2002.

(3) Ebbing, R. W.D. D., *Experiments in General Chemistry*; Houghton Mifflin Harcourt Publishing Company: 2004.

(4) Selfe, I. B.G.O. S., *General, Organic, and Biochemistry Lab Manual*; Freeman: 2006.

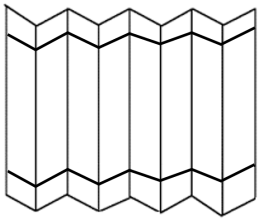
Procedure

Prepare the setup & add the chemicals

☐ *Step 1:* – Obtain the materials.

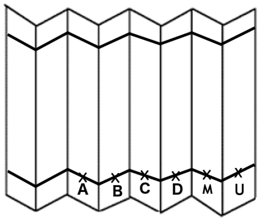
☐ *Step 2:* – Take the piece of filter paper and use a pencil–never a pen–and a ruler to draw a line 2 cm from one of the longer edges. Draw another line 2 cm from from the the opposite edge.

☐ *Step 3:* – Fold the paper in half. Fold the paper again in the half of the half. And once more before you unfold it, to obtain an accordion-like flexure as shown in the figure below:



☐ *Step 4:* – You should now have the 2 pencil lines divided into 8 equal segments. If you are not sure about the result, show the folded paper to your instructor.

☐ *Step 5:* – Mark the center of each of the six inner segments of one line with an X (use only pencil), and write the labels: *A*, *B*, *C*, *D* and *U* under the marks as shown below. *A*-*C* are the blanks (references), *M* is a mixture of the four blanks and *U* is for the unknown.



☐ *Step 6:* – With a clean capillarity tube, add a single drop of each chemical on its mark. Use a clean capillary for each blank so that the blanks do not mix. The drop should be small, barely visible, result of applying the capillary for 2 seconds on the filter paper.

☐ *Step 7:* – Now you will make the mobile phase by mixing about 30 mL of the acetone (or an equivalent organic solvent) with about 10 mL of concentrated HCl ⚠solution in the 600 mL beaker. Cover the beaker with a watch glass so that the solvents do not evaporate.

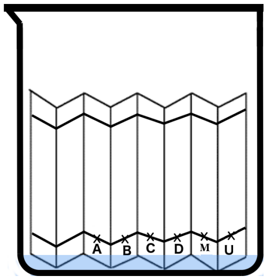
Good Lab Practice

Obtain directions from your instructor before discharging the chemicals in this experiment.

Starting the chromatography

☐ *Step 8:* – Very gently, place the folded paper standing inside the beaker with the X-marked line in the lower part. The liquid should not reach the line with the marks (see image below). Cover the beaker with the watch glass again. Make sure the paper is not coming out of the beaker. If the paper touches the lit of the walls of the beaker you will have to repeat all the steps above.

☐ *Step 9:* – Once the chromatography has started you are not allowed to move the beaker until the experiment is finished.



Good Lab Practice

You are not allowed to move the beaker after the chromatography has started

☐ *Step 10:* – The accordion paper should be flat on the liquid and the beaker must be absolutely stationary through the experiment.

☐ *Step 11:* – Now the mobile phase will move up passing by the bottom line. Make sure you see the line of the wet paper moving upwards.

☐ *Step 12:* – When the liquid has reached the top line open the beaker and take out the paper.

Revealing the marks

☐ *Step 13:* – This part should be done in the hoods and you will only take the paper out when it is completely dry.

☐ *Step 14:* – Wearing plastic gloves, spray the paper with a solution of NH_3 ⚠. The paper should never be wet but just moist.

☐ *Step 15:* – Spray the paper with a solution of NaS_2 ⚠.

☐ *Step 16:* – Let the the paper dry in an adjacent hood and only when it is totally dry you can return it to your station.

☐ *Step 17:* – Circle each spot with a pencil.

☐ *Step 18:* – Measure the vertical distance that the approximate center of each of these spots has traveled from the bottom 2-cm-line. Record the distances in the Results table.

☐ *Step 19:* – Measure the vertical distance travelled by the mobile phase (the front of the liquid) from the bottom 2-cm-line. Record the distance in the Results table.

☐ *Step 20:* – Calculate the R_f value for each spot.

☐ *Step 21:* – Use the R_f values to identify each component in the unknown mixture.

☐ *Step 22:* – Dispose ✍ the different chemicals in the waste disposal.

⚠CAUTION!

⚠ Sodium sulfide (NaS_2) is very toxic and should never be tasted or inhaled.

⚠ Nitric acid (HNO_3) and Ammonia (NH_3) can cause chemical burns.

Calculations

① Measure the vertical distance that the approximate center of each of these spots has traveled from the bottom 2-cm-line.

② Measure the vertical distance travelled by the mobile phase (the front of the liquid) from the bottom 2-cm-line.

③ Divide each ② by ①

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Empirical rules for polarity and Paper Chromatography

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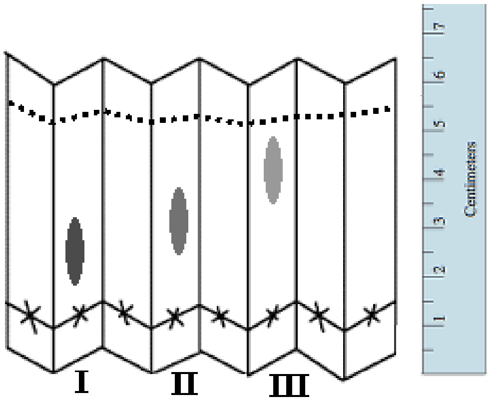
1. Define mobile and stationary phase.

2. A pair of chemicals are separated by means of paper chromatography using methanol (CH_3OH) as carrier solvent. The chemicals that were separated are ethanol (CH_3CH_2OH) and Benzene (C_6H_6). Which substance will have the larger R_f value? Explain your reasoning.

3. The figure below reports a chromatogram for the separation of three compounds (I, II and III), where the solvent level is marked with the dotted line. Calculate R_f for each compound.



	I	II	III
R_f			

STUDENT INFO

Name: ____________________ Date: __________

Results

EXPERIMENT

Empirical rules for polarity and Paper Chromatography

---

① Distance traveled by the mobile phase (cm): ____________

	Blank name	Distance travelled (cm) ②	R_f values ③
A			
B			
C			
D			
Unknown	use as many rows as spots reveal in the unknown sample		

Components of the unknown mixture: ____________

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Empirical rules for polarity and Paper Chromatography

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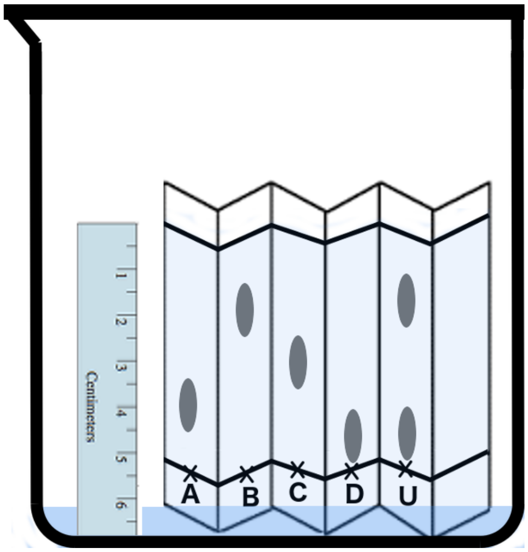
1. For the following experiment, identify the chemicals in the unknown and calculate R_f for each compound.



EXPERIMENT

Periodic properties of elements and compounds

Goal

The goal of this experiment is to qualitatively look for *atomic trends* in a series of atoms and molecules. On one hand, you will quantitatively study the change of several atomic properties such as the atomic radius, the ionization energy and electron affinity on the periodic table. On the other hand, you will compare the solubility of a series of compounds across the periodic table. Overall, this study demonstrates the different trends that atomic and molecules properties follow in the periodic table.

Materials

On one hand, this is a modeling-based experiment, and you will need internet connection and a computer to complete the first part of the experiment. All properties in this experiment will be taken from 🌍 *www.webelements.com*. On one hand, you will need a series of reagents and materials for the second part of the experiment:

☐ A 5 mL graduated cylinder (or 10 mL)

☐ 5 small test tubes

☐ the following solutions: (a) 0.1M $Mg(NO_3)_2$, $Ca(NO_3)_2$, $Sr(NO_3)_2$, $Ba(NO_3)_2$, and $Pb(NO_3)_2$

☐ the following solutions: (a) 0.1M NaOH, NaCl, NaBr, NaI, Na_2SO_4, Na_2CO_3, and $Na_2C_2O_4$

Background

The periodic table can be used to estimate atomic properties, as the elements in the table are arranged in order of increasing atomic number (Z). At the same time, the elements are arranged in such a way that the electron configuration displays a periodic variation: an equivalent kind of electron configuration occurs within a group (a vertical column) and varies along each period (horizontal row) on the table. For example, F and Cl are in the same group and its electronic configuration is $[He]2s^2 2p^5$ and $[Ne]3s^2 3p^5$, respectively. while H and He are located in the same period (period 1) and its configuration is $1s^1$ and $1s^2$, respectively.

Learn more about periodic trends in the periodic table in this 🌍 Video Link

Visual

The Periodicity of Atomic Properties

Atom do not have well-defined boundaries because the electrons in the electronic cloud are delocalized in space. Still, when they form solids or molecules, the distance between atoms can be measured. This measurement allow us to define an *atomic radius* as half the distance between the neighboring atoms. Atoms can loose or gain electrons forming cations or anions, respectively. The radius of cations and anions–*the ionic radius*–differ from the radii of their parent atoms. On one hand, all cations are smaller than their parent atoms, whereas anions are larger than their parent atoms.

The *ionization energy*, I, is the energy needed to remove an electron from a neutral atom in the gas phase. Elements with low ionization energies can easily form cations, whereas elements with high ionization energies are unlikely to form cations. I is normally expressed in electron-volts (eV), a unit that is related to joules by the following conversion factor:

$$\frac{1\,eV}{1.602 \times 10^{-19}\,J}$$

Sometimes I is expressed in kJ/mol that can also be converted to eV:

$$\frac{1\,kJ/mol}{1.04\times 10^{-2}\,eV}$$

The *electron affinity*, EA, of an element is the energy released when a single electron is added to a gas-phase atom. Positive electron affinities means that energy is released when an electron is added to an atom, whereas negative EA values means that energy must be supplied to attach an electron to an atom. Similar to ionization energies, EA values are reported in eV for a single atom. All atomic properties–atomic radius, ionization energy and electron affinity–can be found on web-based periodic tables such as 🌍 www.WebElements.com.

Example

Find the values of the atomic radius (in pm), the first ionization energy and the electron affinity (in eV) for Chlorine in 🌍 www.WebElements.com (1 $pm = 1\cdot 10^{-12}$ m).

Answer: Look for the Chlorine section in www.WebElements.com, and click on the 'More Properties → Atoms Sizes' menu, you will find the 'empirical' atomic radius, the one calculated experimentally, with a value of 100 pm. In order to find the first ionization energy of Cl, navigate to 'More Properties → Electron shell data' menu. It is 1251.2 kJ/mol. We can convert this value to eV:

$$1251.2\,kJ/mol \times \frac{1.04\times 10^{-2}\,eV}{1\,kJ/mol} = 13.01\,eV$$

The electron affinity of Cl is 349 kJ/mol that is 3.6 eV.

Finally, the electronegativity of an element i is a measure of the tendency of an atom to attract a bonding pair of electrons. The Pauling scale is the most commonly used.

The Periodicity of Molecular Properties

Mendeleev periodic table of atoms is one of the most important principles in natural science. Even if there is not such a thing for molecules, one can arrange some molecular properties according to the periodic table. For example, HF is a covalent molecule with a molecular H-F distance of 92pm. The bond distance for HCl, HBr and HI is 127, 141 and 161 pm, respectively. These results show that for the HX distance, with X standing for F, Cl, Br and I, the H-X molecular distance increases as we decrease in the group. In a similar way the molecular bond of BeH_3 is 119pm, whereas the same parameter for CH_4, NH_3, H_2O and HF is 109, 96 and 92 pm. These results indicate that the molecular H-X bond decreases when going from left to right of a period.

Solubility and double replacement reactions

Learn more about solubility in this 🌍 Video Link

Visual

Precipitation reactions occur when a dissolved substances react to form a solid product. Many reactions of this type involve the exchange of ions between ionic compounds in aqueous solution and are sometimes referred to as double replacement reactions. For example, alkaline earth metals and lead form nitrates, hydroxides, chlorides, bromides, and iodides with the general formulas $M(NO_3)_2$, $M(OH)_2$, MCl_2, MBr_2 and MI_2, respectively, where M stands for the alkaline earth and lead cation. These compounds react with sodium hydroxides, chlorides, bromides, and iodides in water by means of the reaction:

$$M(NO_3)_2 \quad + \quad 2\,NaX \quad \longrightarrow \quad MX_2 \quad + \quad 2\,NaNO_3$$

where X stands for hydroxide (OH^-), chloride (Cl^-), bromide (Br^-), or iodide (I^-). In this set of reactions, M replaces Na in NaX, and NO_3 replaces X in NaX, producing a double replacement. $NaNO_3$ is soluble in water and hence if a precipitate appears it can only be due to MX_2. We will use this trick to determine the qualitative solubility of the hydroxides, chlorides, bromides, and iodides in water.

In a similar way, alkaline earth metals and lead also form sulfates, carbonates, oxalates and chromates with the general formula MSO_4, MCO_3, MC_2O_4, and $MCrO_4$, respectively, where M stands for the alkaline earth and lead cation. These compounds react with sodium hydroxides, chlorides, bromides, and iodides in water by means of the reaction:

$$\mathrm{M(NO_3)_2} \quad + \quad \mathrm{Na_2Y} \longrightarrow \mathrm{MY} \quad + \quad 2\,\mathrm{NaNO_3}$$

where Y represent sulfate, carbonate, oxalate and chromate. The solubility of the MY compounds can be determined in exactly the same way as those of the MX compounds.

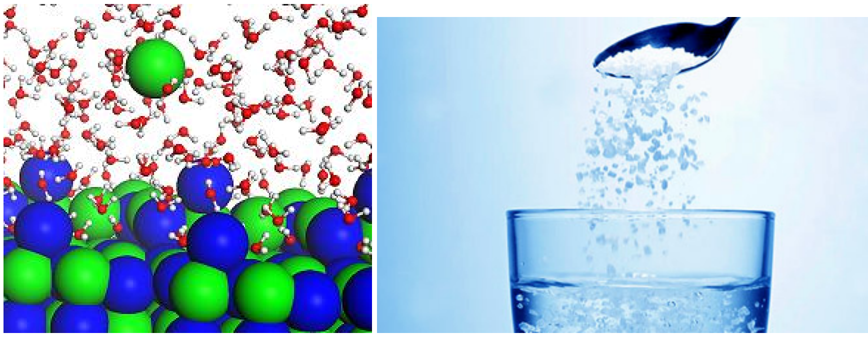


Figure 25: (Left Panel) Representation of solubility. (Right Panel) the dissolution of salt.

Procedure

Periodic properties of atoms

☐ *Step 1:* – Record from 🌍 www.WebElements.com the atomic radius, first ionization energy, the electron affinity and the Pauli's electronegativity of the elements in the *Results* sections. Convert the energy units to eV.

☐ *Step 2:* – Plot the atomic radius, the first ionization energy in eV and the electron affinity in eV vs. the atomic number in the graph paper provided.

Periodic properties of compounds

☐ *Step 1:* – Obtain 5 small test tubes, marking them with *identification numbers*.

☐ *Step 2:* – In the following steps you will create several *1mL marks* on the test tubes. To do this: first use a 5mL (or 10mL) graduated cylinder to place 1mL of distilled water in each of these test tubes. Mark the height of the water in each test tube with a small piece of tape or marking pencil.

☐ *Step 3:* – Now add an additional 1mL to each test tube and mark the new height of the water in each.

☐ *Step 4:* – After creating all marks, pour the water into the sink.

Good Lab Practice

✍ Obtain direction from your instructor for discharging the solutions

Determining qualitatively the solubility of a set of hydroxides $X(OH)_2$

☐ *Step 1:* – Using the lower marks on the test tube as a guides, add 1mL of 0.1M $Mg(NO_3)_2$ to the first test tube. Then add 1mL of 0.1 M $Ca(NO_3)_2$ to the second, 1mL of 0.1M $Sr(NO_3)_2$ to the third, 1mL of 0.1M $Ba(NO_3)_2$ to the fourth and 1mL of 0.1M $Pb(NO_3)_2$ to the fifth test tube. Make sure you are aware of the color and clarity of each solution.

☐ *Step 2:* – Using the upper marks as guides, add 1mL of 1M NaOH to each test tube. Shake each tube gently and wait for about 30s.

☐ *Step 3:* – Notice the colors of all precipitates. A precipitate may come across as a change in the clarity of the solution. A precipitate may appear as 'cloudy' or milky. Record your observations on the results table.

☐ *Step 4:* – Discard the contents of each of the test tubes as directed by your laboratory instructor. Wash the test tubes carefully and rinse them with distilled water.

⚠CAUTION!

⚠ Lead, Barium and oxalate are toxic. Mind to wash your hands throughly after using those compounds.

!

Determining qualitatively the solubility of a set of chlorides (XCl_2), bromides (XBr_2) and iodides (XI_2)

☐ *Step 1:* – Using the lower marks on the test tube a guide, add 1mL of 0.1M $Mg(NO_3)_2$ to the first test tube. Then add 1mL of 0.1 M $Ca(NO_3)_2$ to the second, 1mL of 0.1M $Sr(NO_3)_2$ to the third, 1mL of 0.1M $Ba(NO_3)_2$ to the fourth and 1mL of 0.1M $Pb(NO_3)_2$ to the fifth test tube. Make sure you are aware of the color and clarity of each solution.

☐ *Step 2:* – Using the upper marks as guides, add 1mL of 1M NaCl to each test tube. Shake each tube gently and wait for about 30s.

☐ *Step 3:* – Notice the colors of all precipitates. A precipitate may come across as a change in the clarity of the solution. A precipitate may appear as 'cloudy' or milky. Record your observations on the results table.

☐ *Step 4:* – Discard the contents of each of the test tubes as directed by your laboratory instructor. Wash the test tubes carefully and rinse them with distilled water.

☐ *Step 5:* – Repeat the previous steps six more times in turn adding each of the following solution instead of the solution of NaCl: 0.1M NaBr and 0.1M NaI.

Determining qualitatively the solubility of a set of sulfates ($Na(SO_4)_2$), carbonates ($Na(CO_3)_2$) and oxalates ($Na_2C_2O_4$).

☐ *Step 1:* – Using the lower marks on the test tube as guides, add 1mL of 0.1M $Mg(NO_3)_2$ to the first test tube. Then add 1mL of 0.1 M $Ca(NO_3)_2$ to the second, 1mL of 0.1M $Sr(NO_3)_2$ to the third, 1mL of 0.1M $Ba(NO_3)_2$ to the fourth and 1mL of 0.1M $Pb(NO_3)_2$ to the fifth test tube. Make sure you are aware of the color and clarity of each solution.

☐ *Step 2:* – Using the upper marks as guides, add 1mL of 0.1M $Na(SO_4)_2$ to each test tube. Shake each tube gently and wait for about 30s.

☐ *Step 3:* – Notice the colors of all precipitates. A precipitate may come across as a change in the clarity of the solution. A precipitate may appear as 'cloudy' or milky. Record your observations on the results table.

☐ *Step 4:* – Discard the contents of each of the test tubes as directed by your laboratory instructor. Wash the test tubes carefully and rinse them with distilled water.

☐ *Step 5:* – Repeat the previous steps six more times in turn adding each of the following solution instead of the solution of $Na(SO_4)_2$: 0.1M $Na(CO_3)_2$, and 0.1M $Na_2C_2O_4$.

Calculations

① In 🌍 www.WebElements.com for each element, click on the ’Atoms Sizes’ menu, you will find the ’empirical’ atomic radius R, the one calculated experimentally. Write this value on the *Results* table.
Plot R vs. the atomic number Z. R, in pm, goes in the vertical axis, whereas Z goes in the horizontal axis.

② In 🌍 www.WebElements.com for each element, you will find the ’first ionization energy’ I in kJ/mol. Write this value on the *Results* table.

③ To convert the ionization energy I from kJ/mol to eV do:

$$\textcircled{2} \times \frac{1.04 \times 10^{-2}\, eV}{1\, kJ/mol}$$

Plot I vs. the atomic number Z. I in eV goes in the vertical axis, whereas Z goes in the horizontal axis.

④ In 🌍 www.WebElements.com for each element, selecting ’all properties’ you will find the electron affinity EA in kJ/mol. Write this value on the *Results* table.

⑤ To convert the electron affinity EA from kJ/mol to eV do:

$$\textcircled{4} \times \frac{1.04 \times 10^{-2}\, eV}{1\, kJ/mol}$$

Plot EA vs. the atomic number Z. EA in eV goes in the vertical axis, whereas Z goes in the horizontal axis.

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Periodic properties of elements and compounds

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1. The first ionization potential of Carbon is 1086.45 kJ/mol. Convert this value into eV.

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2. The atomic radius of Carbon is 70pm. Convert this value into meters.

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3. Using 🌍 www.WebElements.com fill the following table with the properties of the atom indicated.

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	Atomic Number	Atomic radius	First Ionization energy		Electron affinity	
Element	Z	(pm)	(kJ/mol)	(eV)	(kJ/mol)	(eV)
H						

4. Using 🌍 www.WebElements.com fill the following table with the properties of the atom indicated.

This table in the pdf is writable only with Acrobat Reader...

Element	Atomic Number Z	Atomic radius (pm)	Electronegativity
H			

5. Name or give the formula for the following compounds:

$Mg(NO_3)_2$

$Ca(NO_3)_2$

$Sr(NO_3)_2$

$Ba(NO_3)_2$

$Pb(NO_3)_2$

Sodium chloride

Sodium bromide

Na_2SO_4

Na_2CO_3

$Na_2C_2O_4$

6. Give the formula for the compound resulting of mixing the following chemicals:

$Mg(NO_3)_2 + NaOH$

$Mg(NO_3)_2 + NaCl$

$Mg(NO_3)_2 + NaBr$

$Mg(NO_3)_2 + NaI$

$Mg(NO_3)_2 + Na_2SO_4$

$Ca(NO_3)_2 + Na_2CO_3$

$Ca(NO_3)_2 + Na_2C_2O_4$

$Sr(NO_3)_2 + NaOH$

$Ba(NO_3)_2 + NaI$

$Pb(NO_3)_2 + Na_2SO_4$

STUDENT INFO

Name: ______________________ Date:______

Results
EXPERIMENT

Periodic properties of elements

Element	Atomic Number	Atomic radius	First Ionization energy		Electron affinity		Electronegativity
	Z	(pm) ①	(kJ/mol) ②	(eV) ③	(kJ/mol) ④	(eV) ⑤	
H							
He							
Li							
Be							
B							
C							
N							

Element	Atomic Number	Atomic radius	First Ionization energy		Electron affinity		Electronegativity
	Z	(pm) ①	(kJ/mol) ②	(eV) ③	(kJ/mol) ④	(eV) ⑤	
O							
F							
Ne							
Na							
Mg							
Al							

Element	Atomic Number Z	Atomic radius (pm) ①	First Ionization energy (kJ/mol) ②	First Ionization energy (eV) ③	Electron affinity (kJ/mol) ④	Electron affinity (eV) ⑤	Electronegativity
Si							
P							
S							
Cl							
Ar							
K							

Element	Atomic Number	Atomic radius	First Ionization energy		Electron affinity		Electronegativity
	Z	(pm) ①	(kJ/mol) ②	(eV) ③	(kJ/mol) ④	(eV) ⑤	
Ca							
Ga							
Ge							
As							
Se							
Br							
Kr							

Figure 26: Atomic Radius Plot: Atomic radius (Y), atomic number(X)

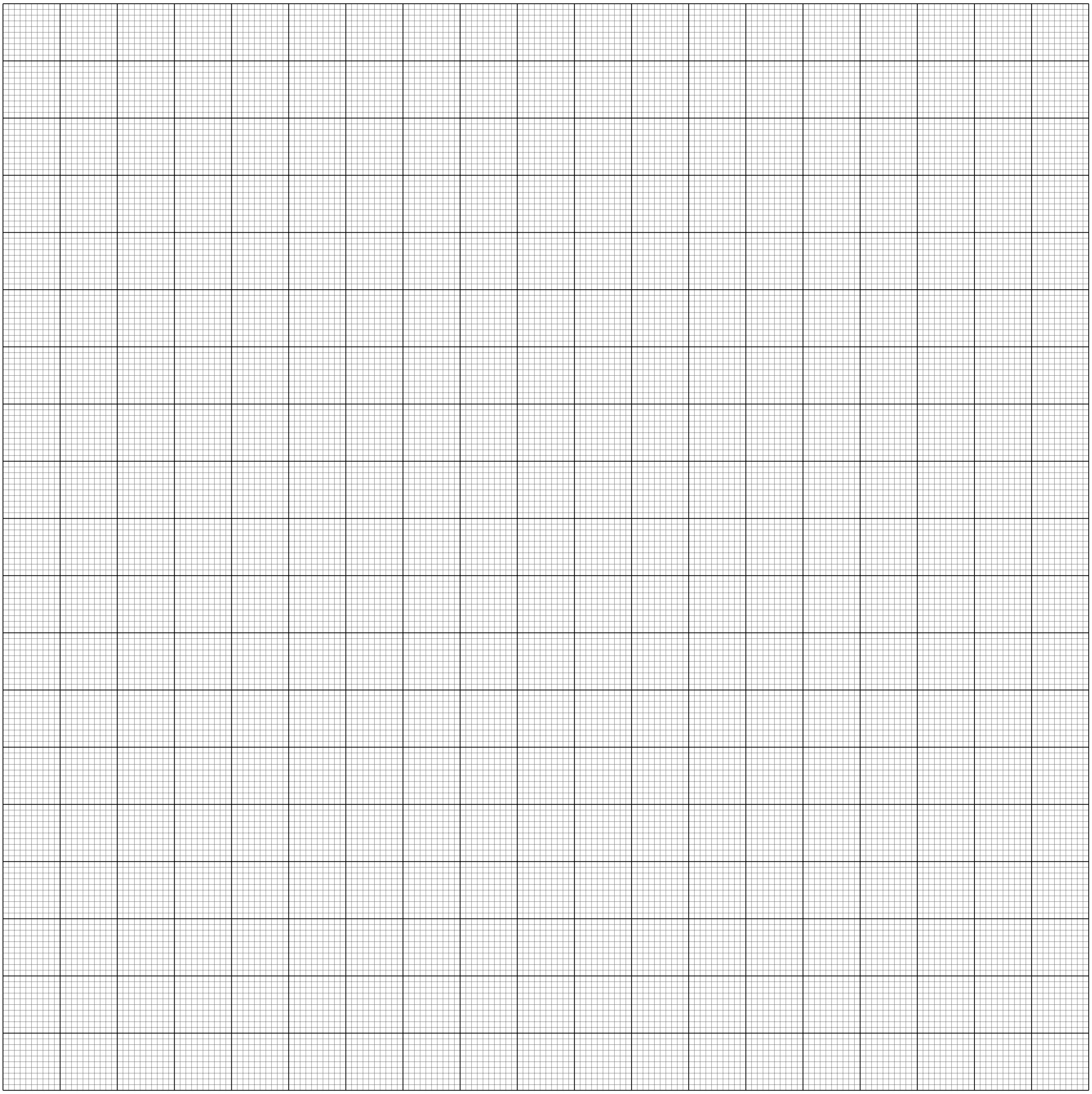

Figure 27: Ionization Potential Plot: Ionization potential (Y), atomic number(X)

Figure 28: Electron Affinity Plot: electron affinity (Y), atomic number(X)

Figure 29: Electronegativity: electronegativity (Y), atomic number(X)

	$Mg(NO_3)_2$	$Ca(NO_3)_2$	$Sr(NO_3)_2$	$Ba(NO_3)_2$	$Pb(NO_3)_2$
NaOH					
NaCl					
NaBr					
NaI					
Na_2SO_4					
Na_2CO_3					
$Na_2C_2O_4$					

(write *S* for soluble product and *I* for insoluble product. If insoluble, indicate the color of the precipitate)

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Periodic properties of elements and compounds

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1. From your data, indicate how the atomic radius changes across a period and across a row of the periodic table.

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2. From your data, indicate how the ionization energy changes across a period and across a row of the periodic table.

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3. From your data, indicate how the electron affinity changes across a period and across a row of the periodic table.

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4. From your data, indicate how the electronegativity changes across a period and across a row of the periodic table.

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5. What can you conclude from your data about the trends in the solubility of the halide compounds as the atomic number increases?

6. What can you conclude from your data about the trends in the solubility of the hydroxides as the atomic number increases?

7. What can you conclude from your data about the trends in the solubility of the sulfates, carbonates and oxalates as the atomic number increases?

8. Compare the solubility of the lead compounds and the alkali earth metals. Give and explanation for these differences.

EXPERIMENT

Geometric isomers

Goal

The goal of this experiment is to understand the concept of **isomerism**. During this practice the student will identify two different molecules based on their properties. Later, a simple reaction will be carried out in an attempt to transform each molecule into it′s corresponding isomer. The student should be able to identify the new products based on the first observations (i.e. the properties of each molecule), and to validate the success of the transformation. The experiment will also familiarize the students with the separation technique known as suction filtration.

Materials

- ☐ 2 small test tubes
- ☐ 2 suitable test tube stoppers
- ☐ 2 test tube holders
- ☐ Bunsen burner
- ☐ 250 mL Erlenmeyer
- ☐ Ring stand
- ☐ Iron ring
- ☐ Wire gauze
- ☐ Büchner funnel
- ☐ Rubber tubing
- ☐ Clamp
- ☐ Suction flask
- ☐ Round filter paper
- ☐ Weighing boats
- ☐ Glass funnel
- ☐ Fumaric Acid, *white powder*
  Maleic Acid, *white powder*
  Hydrochloric Acid HCl (aq) 12 M, *transparent liquid*
  Ice

Background

Learn about 🌍 *Suction filtration*

Visual

Consider a set of Lego® blocks with different sizes and colors. We can combine them in numerous ways to construct many different structures. Similarly, if those block were atoms, we would obtain different molecules from the same set of atoms. Isomers are molecules with the same number and nature of atoms, and therefore the same molecular formula, but with different structures. Diverse structures arise from variation in the bonds between atoms and lead to new molecules, with different names and properties.

Geometric isomers are a subclass of isomers where not only the number and nature of the atoms is the same, but also the bonds. In this case it is the spatial distribution or 3D structure that differs. This little change can lead to considerable differences in basic properties such as melting points and solubilities. In molecules with four ligands distributed around a central atom or a double bond the isomerism can arise from the distribution of those four ligands around the molecule, leading to *cis*- and *trans*- isomers. This is the case in the proposed experiment.

The process of **isomerization** is the transformation from one molecule into its corresponding isomer. This transformation does not take place spontaneously, but typically requires external energy and, in some cases, specific conditions. Energy is necessary to break certain bonds in order to allow the molecule to restructure into the corresponding isomer.

Procedure

Part A. Identify the products by their melting point.

☐ *Step 1:* – Obtain the two unknowns. One unknown is fumaric acid (*white powder*), and the other is maleic acid (*white powder*).

☐ *Step 2:* – Label two small test tubes as A and B.

☐ *Step 3:* – Put a pea-size amount of one product in test tube A, and a similar amount of the other product in test tube B.

☐ *Step 4:* – Using a striker, light the Bunsen burner. Using a test tube holder for each sample, hold test tubes closely together, in the same hand. The samples must be heated in similar conditions.

Good Lab Practice

- ✍ The burner should stand stable with no additional support. Coil the tubing at each end to stabilize the burner.
- ✍ Work safely and avoid burns.
- ✍ Always ensure easy access to the gas valve. Remove any items around the valve, and do not set the flame on your way.
- ✍ Never leave the flame unattended. In case of accident or any risk, close the gas immediately.

☐ *Step 5:* – Swing the test tubes over the flame. Stop as soon as one of the two powders melts. Pay attention to the test tubes′ label. Was it test tube A or test tube B?

☐ *Step 6:* – Refer to the know properties of maleic acid and fumaric acid (from the prelab) to determine which product is unknown A and which is unknown B.

☐ *Step 7:* – Allow the test tubes to cool down. Discard the contents in the indicated waste containers. Clean the test tubes.

Part B. Identify the products by their solubility in water.

☐ *Step 1:* – Label two small test tubes as A and B.

☐ *Step 2:* – Put a pea-size amount of one product in test tube A, and a similar amount of the other product in test tube B.

☐ *Step 3:* – Add similar amounts of distilled water to each test tube. Approximately 2 mL of water should be enough.

☐ *Step 4:* – Cover the tube using the stoppers and shake them until one of the 2 powders dissolves completely in the liquid.

☐ *Step 5:* – Refer to the known properties of maleic acid and fumaric acid (from the prelab) to determine which product is unknown A and which is unknown B.

☐ *Step 6:* – Your result should be consistent with Part A. If not, first repeat part B to double check if the labeling was correct. If the mistake was not in part B, then repeat part A.

Part C. Isomerization reaction.

☐ *Step 1:* – Obtain **about** 1.0 grams of maleic acid (the most soluble sample in Part B) using the scale and a weighing boat.

☐ *Step 2:* – Transfer the 1.0 gram maleic acid powder into a 250 mL Erlenmeyer flask and add **around** 10 mL of distilled water.

☐ *Step 3:* – Prepare a setup to boil the solution that will remove the vapors from the reaction. Set up a ring stand and the iron ring with the wire gauze on top. Do not put the Erlenmeyer with the solution yet! Make sure that the height of the platform is appropriate for the size of the flame. Use a second, larger ring above the platform as protection. To remove the vapors improvise a little hood using a glass funnel. Connect the funnel to the vacuum line using the rubber tubing and set it with a clamp, up side down, at a certain distance over the platform.

⚠CAUTION!

⚠ Handle concentrated acid with care to avoid chemical burns.

⚠ Are you wearing your goggles? Do you know where the eye-washer is?

!

☐ *Step 4:* – Light the burner first. Then, place the Erlenmeyer with the maleic acid solution.

☐ *Step 5:* – While waiting for the solution to boil, put 40 mL of distilled water in a clean 50 mL beaker. Cool down this water using an ice bath. Get ice in a 100 mL beaker and dip the 50 mL beaker in the ice.

☐ *Step 6:* – Obtain 10 mL (aprox.) of concentrated hydrochloric acid, HCl, in a graduated cylinder.

☐ *Step 7:* – When the water starts boiling open the vacuum valve to start the ventilation.

☐ *Step 8:* – Add the acid **carefully and slowly** to the solution and wait until the boiling is reestablished.

☐ *Step 9:* – Let the reaction boil for 1 minute and then turn off the gas.

☐ *Step 10:* – Allow the Erlenmeyer to cool down for 10 minutes. In the meantime, obtain ice in a 600 mL beaker and add some water to the ice.

☐ *Step 11:* – After the 10 minutes cooling down, transfer the Erlenmeyer carefully and safely into the large ice bath.

☐ *Step 12:* – Get the suction filtration setup ready. Connect the suction flask to the vacuum and put the the Büchner funnel on top. Obtain a flat, round filter paper with diameter slightly smaller than the bottom of the funnel. Start the vacuum and add some distilled water with a wash bottle to test the suction.

☐ *Step 13:* – Filter the solution. Stir the mixture to help the precipitate to suspend and add the liquid carefully over the filter paper. Use the cold distilled water in the 50 mL beaker to remove any solid left in the Erlenmeyer and to wash the filtrate.

☐ *Step 14:* – Keep the suction on for another 5 minutes. This will help remove all the water and to partially dry the filter.

☐ *Step 15:* – Turn off the vacuum and remove the filter from the funnel carefully.

☐ *Step 16:* – Identify your product following the steps in Part B.

☐ *Step 17:* – Will isomerization work in the opposite direction? Discover it yourself by repeating Part C for the fumaric acid sample.

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Geometric isomers

---

1. Complete the table for the molecules Fumaric Acid and Maleic Acid

	Fumaric Acid	Maleic Acid
Molecular formula		
Draw the structure		
Solubility in water		
Melting point		
What is it used for?		
Where is it found?		

2. Watch a video about Suction Filtration: Youtube → "Suction Filtration" by "Teaching & Learning Centre"
3. Define:
   (a) Isomers

   (b) Geometric Isomers

(c) Isomerization

(d) Suction filtration

(e) Filtrate

4. What is the difference between a single bond and a double bond

5. What is a *cis* isomer and what is a *trans* isomer

6. Draw the setup for suction filtration (Büchner funnel, suction flask, tubing...)

STUDENT INFO

Name: ______________________ Date: __________

Results

EXPERIMENT

Geometric isomers

This experiment is qualitative, which means that the results are based on observations. Record your observations for each sample during and after each part of the activity.

	Sample A	Sample B
Part A. Melting point		
Part B. Solubility in water		
Part C. Solubility after isomerization		

STUDENT INFO

Name: ____________ Date: ________

Post-lab Questions

Geometric isomers

---

1. Similar amount of both unknowns are used for their characterization in Parts A and B. Why is it important to use similar amounts? What would change if the amounts used were very different?

2. How can you prove that the isomerization process was successful?

3. Why did you cool down the solution after the isomerization?

4. What would happen if hot water was used to wash the filtrate?

EXPERIMENT

Conductivity of electrolyte solutions

Goal

The terms of **conductivity** and **electrolytes** will be introduced and related to the concept of dissociation. The use of the conductivity-meter, very similar to the pH-meter, will help to differentiate between strong and weak electrolytes. In the last part of the experiment, the presence of chemical reactions will be identified by measuring changes in conductivity.

Materials

- ☐ 2 small test tubes
- ☐ 2 suitable test tube stoppers
- ☐ 2 test tube holders
- ☐ Bunsen burner
- ☐ 250 mL Erlenmeyer
- ☐ Ring stand
- ☐ Iron ring
- ☐ Wire gauze
- ☐ Büchner funnel
- ☐ Rubber tubing
- ☐ Clamp
- ☐ Suction flask
- ☐ Round filter paper
- ☐ Weighing boats
- ☐ Glass funnel
- ☐ Fumaric Acid, *white powder*
  Maleic Acid, *white powder*
  Hydrochloric Acid HCl (aq) 12 M, *transparent liquid*
  Ice

Background

Playing with water near electrical appliances can be risky. Surprisingly, however, water does not conduct electricity. For electricity to propagate through media there must be free charges that move along the electric field. Pure water does not contain a significant number of charges and does not conduct electricity. However, when there is something dissolved in the water, therefore making an aqueous solution, it can be conductive.

Conductivity and aqueous solutions

Whether the aqueous solution is conductive or not depends on the nature of the solute. If the solute dissociates into ions when dissolved in water, the ions will conduct the electricity. These type of chemicals are called electrolytes. If the solute dissolves as a molecule without dissociation, the solution will not conduct electricity and the solute will be called nonelectrolyte.

Units

The unit for conductivity is Siemens per centimeter, S cm^{-1}. Pay attention to the units during this experiment because the tool will often display prefixes like milliSiemens ($\text{mS} = 10^{-3}$ S) or microSiemens ($\mu = 10^{-6}$ S) per meter.

Temperature affects conductivity and must be recorded during the measurement of this property.

Types of electrolytes and factors affecting their conductivity.

Salt and sugar are soluble in water. When sugar dissolves in water its molecules remain intact because water can not break apart the strong covalent bonds between their atoms. As a neutral molecule, sugar in solution will not conduct electricity and it is an example of a **nonelectrolyte**. Salt, sodium chloride or NaCl, is an ionic compound that gives place to two ions, Na^+ and Cl^-, when dissolved in water. These ions are responsible for the solution′s conductivity. The number, or more specifically the concentration, and the mobility of these ions will determine how conductive the solution is. Based on these principles we can classify substances as:

- × **Non-electrolytes**. When the aqueous solution of that substance does NOT conduct electricity.
- × **Weak electrolyte**. When the aqueous solution of that substance conducts electricity poorly.
- × **Strong electrolyte**. When the conductivity of the aqueous solution of that substance is high.

The **mobility** of the ions is related to their size. Small ions move fast which enhances conductivity, while large ions encounter more resistance to move and show lower conductivities. The effect of ions′ mobility is easily appreciated in isomolar solutions. **Isomolar solutions** are solutions with the same molarity.

The more ions in the solution, the higher the conductivity will be. For the same substance, the number of ions is determined by the **concentration**; more concentration means more ions and more conductivity. It is important to recall that only ions contribute to conductivity. While some molecules dissociate completely, others dissociate only partially. The latter are weak electrolytes.

Interestingly, when two solutions containing ions are mixed, the conductivity of the resulting solution will be equal to the sum of the two separate solutions with the same concentration as in the mixture. This is called the **additivity rule** and will hold true as long as the ions do not react with each other.

Example

The conductivity of a 0.050 M solutions of HNO_3 is measure to be 19.6 mS/cm.
That of a 0.050 M solution of KCl is 9.6 mS/cm.
A new solution is prepared mixing 10 mL of HNO_3 0.10 M and 10 mL of KCl 0.10 M. The conductivity of the resulting mixture is measure to be 29.6 mS/cm.
Calculate the concentration of each spices in the mixture and determine if their compounds have reacted or not.

Answer: To calculate the new concentration use the formula for dilutions:

$$V_1 \times M_1 = V_2 \times M_2$$

where V_1 and M_1 are the initial volume and concentration and V_2 and M_2 are the values after the dilution.
We look for M_2, therefore

$$M_2 = V_1 \times \frac{M_1}{V_2}$$

for HNO_3,

$$M_2 = \frac{10\,mL \times 0.10\,M}{(10\,mL + 10\,mL)} = 0.05 \;\; M$$

The same values apply for KCl.
We can see that the addition of the separate conductivities is very close to the total conductivity of the mixture with the same concentrations for each electrolyte.

$$\text{Cond. } (HNO_3 \text{ 0.05 M}) + \text{Cond } (\text{KCl 0.05 M}) = 19.6 \text{ mS/cm} + 9.6 \text{ mS/cm} = 29.2 \text{ mS/cm}$$
$$\text{Cond. } HNO_3 \text{ 0.05M} + \text{KCl 0.05M} = 29.6 \text{ mS/cm}$$

The additivity rules applies, meaning that there is no reaction between HNO_3 and KCl.

When two electrolytes undergo a chemical reaction, new products will form and the additivity rule will not apply. A large deviation from the sum of separate conductivities is an indication of a chemical reaction.

Procedure

Part A. Getting started.

☐ *Step 1:* – Install the conductivity-meter. The probe of the instrument is immersed in a clean solution that must be kept aside until the end of the experiment.

☐ *Step 2:* – Cleaning and measuring. It is very important to follow the cleaning steps before each measurement. You will need a 100 mL beaker labeled as waste, a wash bottle and a 250 mL beaker with distilled water.

1. Rinse the probe using the wash bottle over the waste beaker.
2. Dip the probe 3-4 times into the 250 mL beaker with clean distilled water.
3. Dip the probe into the solution to be measured. Use a 50 mL beaker with at least 25 mL of solution.
4. Repeat cleaning steps 1 and 2.

☐ *Step 3:* – Following the procedure outlined in step 2, measure the following solutions:

- × Distilled water
- × Tap water
- × Ethyl alcohol (C_2H_5OH)

Record all the values.

Good Lab Practice

✍ To ensure the probe is well cleaned and the water in the 250 mL beaker remains clean, the conductivity meter should display a value close to that of the distilled water during the cleaning.

✍ If not, replace the water in the 250 mL beaker with fresh distilled water, and thoroughly clean the probe.

Part B. Series of Dilutions.

☐ *Step 4:* – Starting from the concentrated 6 M solution of HCl, prepare the following dilutions; 0.10 M, 0.050 M and 0.020 M.

☐ *Step 5:* – Measure the conductivity of the three HCl dilutions and record the values.

Part C. Dilutions 1:1.

☐ *Step 6:* – In a clean 50 mL beaker, add 10 mL of 0.10 M HNO_3 and 10 mL of distilled water. Mix well and measure the conductivity.

☐ *Step 7:* – In a clean 50 mL beaker, add 10 mL of 0.10 M KCl and 10 mL of distilled water. Mix well and measure the conductivity.

☐ *Step 8:* – In a clean 50 mL beaker, add 10 mL of 0.10 M KNO_3 and 10 mL of distilled water. Mix well and measure the conductivity.

☐ *Step 9:* – In a clean 50 mL beaker, add 10 mL of 0.10 M $Ca(NO_3)_2$ and 10 mL of distilled water. Mix well and measure the conductivity.

Part D. Ammonia and acetic acid.

☐ *Step 10:* – Prepare the following dilutions; acetic acid ($HC_2H_3O_2$) 0.10 M and ammonia (NH_3) 0.10 M. Carefully read the concentration of the initial concentrated solutions indicated in the labels.

☐ *Step 11:* – In a clean 50 mL beaker, add 10 mL of the 0.10 M NH_3 dilution and 10 mL of distilled water. Mix well and measure the conductivity.

☐ *Step 12:* – In a clean 50 mL beaker, add 10 mL of 0.10 M $HC_2H_3O_2$ and 10 mL of distilled water. Mix well and measure the conductivity.

Part E. Mixtures.

☐ *Step 13:* – In a clean 50 mL beaker, add 10 mL of 0.10 M HCl and 10 mL of 0.10 M KNO_3. Mix well and measure the conductivity.

☐ *Step 14:* – In a clean 50 mL beaker, add 10 mL of 0.10 M HNO_3 and 10 mL of 0.10 M KCl. Mix well and measure the conductivity.

☐ *Step 15:* – In a clean 50 mL beaker, add 10 mL of 0.10 M NH_3 and 10 mL of 0.10 M $HC_2H_3O_2$. Mix well and measure the conductivity.

You need to recalculate the concentrations of all the solutions after the dilutions and mixtures.

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Conductivity of electrolyte solutions

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1. Find the definition of the following concepts:
   - (a) Aqueous solution
   - (b) Ionic compound
   - (c) Molecular compound
   - (d) Concentration
   - (e) Molarity
   - (f) Concentrated
   - (g) Diluted
   - (h) Ionic compound
   - (i) Electrolyte

2. Starting with a 6 M solution of NaCl, calculate the volume necessary to prepare 10 mL of a 3 M solution of NaCl.

(a) How much water do you have to add?

(b) From 6 M to 3 M the solution is diluted into half of the initial concentration. How will you dilute the 3 M solution into a 1.5 M solution?

STUDENT INFO

Name: ______________________ Date: __________

Results

EXPERIMENT

Conductivity of electrolyte solutions

Solution	Concentration (M)	Conductivity	Units
Distilled water			
Tap water			
C_2H_5OH			
HCl			
HCl			
HCl			
HNO_3			
KOH			
KCl			
KNO_3			
$Ca(NO_3)_2$			
HNO_3			
$HC_2H_3O_2$			

Mixture		Concentration (M)	Conductivity	Units
$HCl+KNO_3$	HCl			
	KNO_3			
HNO_3+KCl	HNO_3			
	KCl			
$HC_2H_3O_2+NH_3$	$HC_2H_3O_2$			
	NH_3			

STUDENT INFO

Name:		Date:	

Post-lab Questions

Conductivity of electrolyte solutions

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1. What is the difference between a strong electrolyte and a weak electrolyte? Use examples from the experiment.

2. How does concentration affect conductivity? Use your data to illustrate your answer.

3. How does ion mobility affect conductivity? Use your data to illustrate your answer.

4. Which solution′s conductivity was higher, distilled water or tap water? Why ?

5. Based on the conductivity measurements from the experiment , indicate weather each of the following is a weak electrolyte, a strong electrolyte or a nonelectrolyte.

Sample	Strong, Weak or Non-electrolyte
H_2O	
C_2H_5OH	
HCl	
HNO_3	
$HC_2H_3O_2$	

Sample	Strong, Weak or Non-electrolyte
KNO_3	
KOH	
$Ca(NO_3)_2$	
NH_3	
KCl	

6. Acids and bases are classified as strong or weak according to their degree of dissociation when dissolved in water. A weak acid/base will dissociate only partially, while a strong acid/base will dissociate completely. Relate this statement to your results and indicate weather each of the following is a weak acid, a strong acid, a weak base or a strong base.

Acids	Strong or Weak
HCl	
HNO_3	
$HC_2H_3O_2$	

Bases	Strong or Weak
KOH	
NH_3	

EXPERIMENT

Softening of hard water

Goal

The goal of this experiment is to understand the concepts of hard and soft water and their implications in daily life. The student will also experiment the **dissociation of ionic compounds** in water, from which the **hydrogen-donor** nature of the acids and the notion of **pH** are emphasized. Other principles, such as equilibrium and ion exchange, are also presented.

Background

The term **"hard water"** refers to the concentration of certain minerals in the water. This topic concerns **aqueous solutions** where the water is the solvent and the minerals are the solutes. These minerals are ionic compounds that dissociate in aqueous solutions and therefore what we really have is their ions. The hardness is an indication of the concentration of Ca^{2+} and Mg^{2+} ions.

Dissolving the insoluble

Ca^{2+} and Mg^{2+} ions originate in limestone ($CaCO_3$, the material of the pyramids) and chalk ($MgCO_3$, used to write on the chalkboard) deposits. However, if you throw a chalk in water it does not dissolve. Neither do the pyramids dissolve in the rain. How is dissolving these carbonates possible?

The answer is in the air

Water in contact with the air absorbs CO_2 in the atmosphere according to the equation:

$$CO_2\,(g) + H_2O\,(l) \longrightarrow H_2CO_3\,(aq)$$

Where H_2CO_3 is carbonic acid, which is reacts with the carbonate to form a hydrogen bicarbonate. For the case of $CaCO_3$

$$CaCO_3\,(s) + H_2CO_3\,(aq) \longrightarrow Ca(HCO_3)_2\,(aq)$$

Calcium hydrogen carbonate is also called calcium bicarbonate. Bicarbonates are soluble in water (baking soda is sodium bicarbonate, $NaCO_3$) and, as ionic compounds, they are dissociated into its ions;

$$Ca(HCO_3)_2\,(aq) \longrightarrow Ca^{2+} + 2\,HCO_3{}^-$$

The overall reaction is the 3 previous equations added together:

$$CaCO_3\,(s) + CO_2\,(g) + H_2O\,(l) \longrightarrow Ca^{2+} + 2\,HCO_3{}^-$$

and this is how carbonates dissolve in water.

Implications in daily life

The presence of these ions in water has several implications in routine actions in our lives. The problem arises when the latter reaction takes place in the opposite direction, releasing CO_2 gas in the process and depositing solid carbonate. This direction is favored when water is heated and explains the white residue in bathrooms and kettles. This is specially a problem for boilers, that suffer from scaling when operating hard water.

The experiment

The experiment consists of four parts:

- Part A. Preparing the ion exchange resin.
- Part B. Measuring the hardness of a sample of water.
- Part C. Reducing the hardness, also called softening, of the sample.
- Part D. Testing the hardness of the softened sample.

Additionally, a series of simple reactions will be done in order to test the reactivity of Ca^{2+} and Mg^{2+} ions.

Part B. Measuring the hardness of the sample

The concentration of Ca^{2+} ions, or any other metallic cation, is usually measured by a titration of the aqueous sample using a sodium salt as the titrant, a solution of known concentration added to the sample. First an indicator will be added to the sample which will color the solution pink. The indicator chosen is Eriochrome Black T, which will turn blue when the Ca^{2+} ion is removed from the cation-indicator complex:

$$\underset{pink}{\overset{indicator\ complex}{\mathbf{Ca(In)^{2+}}(aq)}} -> \mathrm{Ca}^{2+}\,(\mathrm{aq}) + \underset{blue}{\overset{indicator\ alone}{\mathbf{In}(aq)}}$$

To achieve the color change H_2EDTA^{2-} will be used. H_2EDTA^{2-} (dihydrogen ethylendiaminetetraacetate) is an anion with a stronger affinity for the Ca^{2+} cations than the indicator. As H_2EDTA^{2-} is added to the solution it will replace the indicator forming a new complex, leaving the indicator alone, which in turn will become blue.

$$\mathrm{H_2EDTA^{2-}}\,(\mathrm{aq}) + \underset{pink}{\mathbf{Ca(In)^{2+}}(aq)} <-> \mathrm{Ca(EDTA)}^{2+}(aq) + \underset{blue}{\mathbf{In}(aq)} + 2\,\mathrm{H}^+(aq)$$

The example is for calcium ions, but the same chemistry applies for magnesium ions. It is important to note down the exact number of drops required to achieve the end point of the titration.

Part C. Softening the hard water

After measuring the hardness of the water sample, the next step is to soften the water. **Softening** is the process where ions, particularly calcium and magnesium ions, are removed from the water. When a hard water sample is softened we can refer to it as **softened water**.

A cation-exchange resin will be used to soften the water. This resin contains many sulfonic groups, $(-SO_3H)$. Every two hydrogens in these acidic groups will be replaced by a metallic cation, Ca^{2+} and Mg^{2+}, in a mechanism called **cation exchange** and displayed in Figure 1.

Part D. Evaluating the result.

Finally, it is necessary to evaluate whether the water sample has been softened. As such the titration performed in Part A will be repeated, but this time with the softened sample. The most relevant parameter is the amount of EDTA added to achieve the endpoint of the titration, in order to compare this result to the previous titration.

This experiment aims to be a qualitative analysis of the water treatment. For that reason, and to simplify the process, the titration will be done using droppers instead of the more precise burette typically used in titrations.

Procedure

Part A. Preparing the resin.

☐ *Step 1:* – Obtain a portion of the ion exchange resin in a 250 mL beaker. Remember that this resin can be recycled. Do not discard it at the end of the experiment.

⚠CAUTION!

⚠ Handle concentrated acid with care to avoid chemical burns.

⚠ Are you wearing your goggles? Do you know where the eye-washer is?

☐ *Step 2:* – Add enough 6 M HCl as to cover the resin. By adding the acid, the resin is cleaned of metallic cations. In order to use the ion exchange resin, it must be clean, also known as "in H^+ form". See Figure 1.

☐ *Step 3:* – After 2 minutes, add 200 mL of distilled water to the resin-acid mixture. Do not move it, but rather allow the resin to settle in the bottom of the beaker.

☐ *Step 4:* – Carefully, pour the liquid off (decant) into another beaker. This liquid is a HCl solution and should be handled and discarded as such.

☐ *Step 5:* – To ensure that all the acid is gone, wash again the resin with 200 mL of distilled water. Allow the resin to settle in the bottom of the beaker and poor off the liquid into another beaker.

☐ *Step 6:* – Before you discard the liquid you can test the pH. Dip the tip of a glass stirring rod in the liquid and poke the wet end of the rod to a piece of Blue Litmus paper.

☐ *Step 7:* – If the Litmus paper turns pink, there is still acid, repeat steps 5 and 6. If it does not turn pink, the resin is ready! Reserve the resin for Part C.

Part B. Measure the hardness of the sample

☐ *Step 8:* – Obtain a little bit more than 150 mL of hard water in a beaker.

☐ *Step 9:* – Using a graduated cylinder, add 20.0 mL of hard water to a 125 mL Erlenmeyer. Reserve the rest of the hard water for Part C.

☐ *Step 10:* – Add 5 mL of a $NH_3{}^-NH_4Cl$ buffer solution (not to be confused with Na_2H_2(EDTA)) to the Erlenmeyer, and the tip of a spatula (less than a pea-size amount) of the indicator. Adding too much indicator will ruin the experiment.

Good Lab Practice

✍ The burner should stand stable with no additional support. Coil the tubing at each end to stabilize the burner.

✍ Work safely and avoid burns.

✍ Always ensure easy access to the gas valve. Remove any items around the valve, and do not set the flame on your way.

✍ Never leave the flame unattended. In case of accident or any risk, close the gas immediately.

☐ *Step 11:* – Obtain about 20 mL of 0.01 M Na_2H_2(EDTA) in a 50 mL beaker and a medicine dropper (or a plastic Pasteur pippete). Add the Na_2H_2(EDTA) drop by drop to the mixture while stirring. This is the titration, so count the drops needed for the analyte to turn blue.

Part C. Softening the hard water

☐ *Step 12:* – Transfer the remaining hard water to the beaker with the resin and let it stand while observing reactions with Ca^{2+} and Mg^{2+} ions.

Observing reactions with Ca^{2+} and Mg^{2+} ions

- × Obtain 4 test tubes and the dropper bottles with solutions of 0.10 M $Ca(NO_3)_2$, 0.10 M Na_2CO_3, 0.10 MNaHCO_3, 6 M HCl and 0.10 M $Mg(NO_3)_2$.
- × Test tube 1. Add 10 drops of $\mathbf{Ca(NO_3)_2}$. Add 10 drops of Na_2CO_3 and write down your observations.
- × Test tube 2. Add 10 drops of $\mathbf{Ca(NO_3)_2}$. Add 10 drops of Na_2CO_3 and write down your observations.
- × Test tube 3. Add 10 drops of $\mathbf{Ca(NO_3)_2}$. Add 10 drops of 0.1 M $NaHCO_3$ and write down your observations.
  Heat test tube 3 in the flame of a Bunsen burner. Write down your observations.
  Cool down test tube 3. Add 2 drops of HCl and write down your observations.
- × Test tube 4. Add a pea-sized amount of solid $CaCO_3$. Add 2 drops of HCl and write down your observations.
- × Repeat each step above but instead of starting with 10 drops of $\mathbf{Ca(NO_3)_2}$, use $\mathbf{Mg(NO_3)_2}$.

☐ *Step 13:* – Pour off about 40 mL of the water sample with the resin into a clean beaker. This is the softened water.

Part D. Evaluating the results

☐ *Step 14:* – Using a graduated cylinder, add 20.0 mL of the softened water to a 125 mL Erlenmeyer.

☐ *Step 15:* – Add 5 mL of an $NH_3{}^{-}NH_4Cl$ buffer solution (not to be confused with Na_2H_2(EDTA)) to the Erlenmeyer, and the tip of a spatula (less than a pea-size amount) of the indicator. Adding too much indicator will ruin the experiment.

☐ *Step 16:* – Obtain about 20 mL of 0.01 M Na_2H_2(EDTA) in a 50 mL beaker and a medicine dropper (or a plastic Pasteur pippete). Add the Na_2H_2(EDTA) drop by drop to the mixture while stirring. This is the titration, so count the drops needed for the analyte to turn blue.

Before you go

☐ *Step 17:* – Clean the resin like in Part A by repeating steps 2 to 7.

☐ *Step 18:* – Return the resin to the instructor.

STUDENT INFO

Name: ____________________ Date: ________

Pre-lab Questions

Softening of hard water

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1. Find the definition of the following concepts:
    (a) Titration
    (b) Titrant
    (c) Indicator
    (d) Endpoint
    (e) Litmus paper
    (f) Ion exchange resin
2. Where is the term **hard water** derived from?
3. Find in the manual, and write here a comparison in size for:
    (a) Pea-size
    (b) Tip of a spatula

STUDENT INFO

Name: ______________________ Date: __________

Results
EXPERIMENT

Softening of hard water

Parts B and D. Comparing titrations

Indicate how many drops you used to achieve the endpoint for:

Hard water sample (Part B): __________ drops of Na_2H_2(EDTA)

Softened water sample (Part D): __________ drops of Na_2H_2(EDTA)

Observing reactions with Ca^{2+} and Mg^{2+} ions

	$Ca(NO_3)_2$	$Mg(NO_3)_2$
Na_2CO_3		
Na_2CO_3+HCl		
$NaHCO_3$		
$NaHCO_3$+heat		
$NaHCO_3$+heat+HCl		
Solid $CaCO_3$+HCl		

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Softening of hard water

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1. Based on your measurements. Did the ion exchange resin actually soften the hard water sample? Explain.

2. Write the balanced equations for the chemical reactions observed for Ca^{2+} and Mg^{2+} ions.

EXPERIMENT

Freezing point depression

Goal

The goal of this laboratory is to put into practice the theory seen for **properties of solutions** and to experimentally see the **freezing point depression** effect. The added value is to play detective by relating the nature of **colligative properties** to the molar mass, in order to identify the unknown powder.

Background

During winter, salt is poured over the street to melt the ice on the floor. General knowledge says that the salt melts the ice, while some venture to claim that the salt lowers the freezing temperature of the water. Both assertions might be in the right direction but are, if not false, at least incorrect. When salt is mixed with water, there is no water or salt anymore. Instead there is a **solution** of salt in water. Since the freezing point of the solution is lower than that of the pure water, the solution is no longer solid at outdoors temperatures. Notice that there is no need to state the nature of the solution or the solute because it is irrelevant. The depression of the freezing point is not a property specific to the salt, (i.e. the sodium chloride, NaCl), but rather a general effect of any solute. Salt is used because it is widely available, dissolves fast and is inexpensive.

The formula

The formula governing the change (Δ) in the freezing point (T_f) is:

$$\Delta T_f = K_f m$$

where m is the concentration in **molality** and K_f is the *molal freezing-point depression constant*, which is characteristic of the solvent and is determined experimentally.

ΔT_f is calculated simply from the difference between the freezing point of the pure solvent and that of the solution. Both freezing points will be determined experimentally in Parts A and B below, respectively.
K_f is a parameter specific to each solvent and it is already tabulated.

The mysterious white powder

The **concentration**, m, will be the only unknown in this formula. However, it can be solved. Once the concentration is known, the **number of moles** in the solution can be calculated by simply multiplying the concentration by the volume. Knowing the number of moles it would be possible to find out the **molar mass** of the solute, by using the the mass of the solute that must be measured carefully in advance.
The molar mass will be obtained indirectly from the amount of solute added and the amount of solvent used. For this reason, it is very important to precisely measure the mass of the both the solvent and solute used.

The solvent

The solvent to be used in this activity will be **cyclohexane**, a very **volatile** organic solvent. Special precautions must be taken to avoid evaporation of the solvent .

The experiment

The experiment today is divided in four parts:

Part A. Finding the freezing point of the pure solvent.

Part B. Finding the freezing point of the solution.

Part C. Finding the freezing point after increasing the concentration.

Part D. Solving the molar mass of the solute.

Part A. Finding the freezing point of the pure solvent.

During a phase change all the energy transferred is used to reorganize the molecules and none is used to change the temperature of the substance. A liquid at a higher temperature than its surroundings gives up the heat and its temperature falls until it reaches the freezing point. The liquid continues to give up heat to the surroundings but its temperature remains constant. Only when the all the liquid is frozen will the temperature start to fall again.

If one plots temperature vs time for the freezing of the pure liquid, a negative slope line will represent the cooling down, while the flat horizontal line (plateau) corresponds to the freezing process. Notice that the transition from the cooling sloped line to the horizontal line is not sharp. For that reason the best way to determine the freezing point is to find the horizontal line that fits best the points forming the plateau. That will be the freezing point (temperature). See the left panel on Figure 1.

Parts B and C. Finding the solutions freezing points.

The graph corresponding to a cooling process of the solution, depicted in the left panel of Figure 1, will show the negative slope like in the case of the pure solvent, but will not show the plateau for the phase change. Instead, the slope will simply become less negative. In this case, the freezing point will be obtained by extrapolating both lines and using the temperature at which they intersect.

Part D. Calculating the molar mass.

See an example on how to calculate the molar mass below.

Example

The freezing point of pure benzene is 5.50 Â°C. A solution is prepared by dissolving 0.450 g of an unknown substance in 27.3 g of benzene. The new freezing point is determined to be 4.18 Â°C. What is the molar mass of the unknown substance? The freezing point constant, K_f, for benzene is 5.12 Â°C kg/mol.

Answer: calculate the concentration first:

$$\Delta T_f = K_f \times m \qquad m = \frac{\Delta T_f}{K_f} \qquad m = \frac{5.50°C - 4.18°C}{5.12°C\frac{kg}{mol}} = 0.258\frac{mol}{kg}$$

then, calculate the molar mass:

$$0.258\frac{mol}{kg} \times 27.3g \times \frac{1kg}{1000g} = 7.04 \times 10^{-3} mol \qquad \text{molar mass} = \frac{0.450g}{7.04 \times 10^{-3} mol} = 63.9\frac{g}{mol}$$

Procedure

Part A. Freezing point of the pure solvent.

☐ *Step 1:* – Prepare the set-up to cool down the solution. Set up a ring stand and an iron ring with wire gauze on top. The height of the platform will normally be very low. Use a second, larger ring above the platform as protection. Attach a small clamp above the second iron ring.

☐ *Step 2:* – Obtain a large test tube that is clean and dry, a suitable rubber stopper and a 100 mL beaker. The beaker will be used to hold the test tube vertically. Get the mass of the test tube with stopper and the beaker in the scale.

☐ *Step 3:* – Find another two beakers of 50 mL and 250 mL respectively and a 20 mL pipet. Put ice in the 250 mL beaker and set it on the platform.

☐ *Step 4:* – Pour no more than 25 mL of cyclohexane into the 50 mL beaker. Using the pipet, transfer exactly 20.0 mL of cyclohexane from the 50 mL beaker to the test tube and close it with the stopper immediately.

☐ *Step 5:* – Using the scale get the mass of the test tube with the cyclohexane, the stopper and the 100 mL beaker.

Good Lab Practice

✍ Never add the liquid or solid to be weighed to the container while on the scale. Take the container **off** the scale prior to adding the product. A spill on the scale will ruin both your measurement and the scale.

☐ *Step 6:* – Insert the test tube into the ice and hold it vertically using the clamp. Find the perforated stoppers with the stirring wire and the thermometer. The round wire end must encircle the thermometer probe. Exchange the stopper on the test tube with this gadget.

☐ *Step 7:* – As soon as the temperature is 15°C (or lower) start stirring the solvent with the wire and record the temperature every 15 seconds. Use all the digits given by the thermometer, i.e. down to the tenth of the degree Celsius.

Good Lab Practice

✍ Stirred, not shaken. It is very important to stir the cyclohexane constantly and gently until it is completely frozen.

✍ Make sure the wire stays around the thermometer and that the latter is centered in the test tube.

✍ The thermometer should not touch the walls of the test tube!

✍ A bad stirring method will lead to a partial freezing and the graph will not show straight lines.

☐ *Step 8:* – You will know when to stop the experiment when the temperature starts dropping after several minutes of constant temperature. Do not stop until the temperature reaches the 3°C.

☐ *Step 9:* – Lift the test tube with the clamp above the ice and allow the cyclohexane to melt. Do not use your hands to accelerate the process! Before the cyclohexane gets to 15°C get a second measurement by repeating step 6 to step 8.

☐ *Step 10:* – The same cyclohexane will be used in the Part B.

Part B. Freezing point of the solution.

☐ *Step 11:* – Lift the test tube with the clamp above the ice and allow the cyclohexane to melt. Do not use your hands to accelerate the process! During the next steps be careful to avoid the cyclohexane getting above 15°C.

☐ *Step 12:* – Tare the scale with the weighing boat and obtain between 0.20 g and 0.25 g of the unknown solute. Remove the boat with the solute, tare the empty scale and record the mass of weighing boat together with the solute. Use the balance maximum precision.

☐ *Step 13:* – Carefully add the solute into the test tube with the pure solvent. Make sure no solute is spilled. There is no need to add all the solute from the weighing boat, the amount added will be calculate by difference.

☐ *Step 14:* – Weight the weighing boat with the solute leftovers. Record the mass in the results page.

☐ *Step 15:* – Mix the solution well and insert the test tube in the ice while holding it vertically using the clamp.

☐ *Step 16:* – Start stirring the solution with the wire and record the temperature every 15 seconds. This time the temperature should not reach constant. Stop the experiment when the temperature reaches 1°C.

Part C. Freezing point of a higher concentration solution.

☐ *Step 17:* – Lift the test tube with the clamp above the ice and allow the cyclohexane to melt. Do not use your hands to accelerate the process! During the next steps be careful to avoid the cyclohexane getting above 15°C.

☐ *Step 18:* – The same boat can be reuse without cleaning. Tare the scale with the boat inside. Obtain between 0.10 g and 0.15 g of the unknown solute and record the mass of weighing boat together with the solute using the balanceÂŽs maximum precision.

☐ *Step 19:* – Add carefully the solute into the test tube with the pure solvent. Make sure no solute is spilled. There is no need to add all the solute from the weighing boat, the amount added will be calculate by difference.

☐ *Step 20:* – Weight the weighing boat with the solute leftovers. Record the mass in the results page.

☐ *Step 21:* – Mix the solution well and insert the test tube in the ice while holding it vertically using the clamp.

☐ *Step 22:* – Start stirring the solution with the wire and record the temperature every 15 seconds. This time the temperature should not reach a constant. Stop the experiment when the temperature reaches 0.5°C.

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Freezing point depression

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1. What does volatile mean?
2. What is the empirical formula of a compound?
3. What is the difference between empirical and molecular formulae?
4. What are the colligative properties?
5. What is molality? What are the units?
6. Why is molality used in this experiment instead of Molarity?
7. What is the density of cyclohexane? Use the right units.

STUDENT INFO

Name: ____________________ Date: __________

Results
EXPERIMENT

Freezing point depression

Parts A. Freezing point of the pure solvent.

Mass of the cyclohexane: __________ g

Density of the cyclohexane: __________ g/mL (obtained in the prelab)

Volume of the cyclohexane used: __________ mL

Trial 1				Trial 2			
Time (sec)	Temp. (°C)	Time (sec)	Temp. (°C)	Time (sec)	Temp. (°C)	Time (sec)	Temp. (°C)

(continue in next page)

_______	_______	_______	_______	_______	_______	_______	_______
_______	_______	_______	_______	_______	_______	_______	_______
_______	_______	_______	_______	_______	_______	_______	_______
_______	_______	_______	_______	_______	_______	_______	_______

Freezing point of the pure cyclohexane. Trial 1: __________ °C Trial 2: __________ °C

Mean: __________ °C

Parts B and C. Freezing point of solutions.

Low concentration

Mass of solute+boat: ________ g

Mass of boat (after adding): ________ g

Mass of solute: ________ g

High concentration

Mass of solute+boat: ________ g

Mass of boat (after adding): ________ g

Mass of solute: ________ g

Low concentration				High concentration			
Time (sec)	Temp. (°C)	Time (sec)	Temp. (°C)	Time (sec)	Temp. (°C)	Time (sec)	Temp. (°C)

Part D. Calculating the molar mass

Low concentration

Freezing point: __________ °C

ΔT_f: __________ °C

Molar mass: __________ g/mol

High concentration

Freezing point: __________ °C

ΔT_f: __________ °C

Molar mass: __________ g/mol

Mean molar mass: __________ g/mol

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Freezing point depression

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1. If the empirical formula of the solute used in this experiment is C_3H_2Cl, what is the molecular formula?

2. What would be the effect on the molar mass in the following situations. Explain why.
   (a) A defect in the mass measured for the solute.
   (b) Dirt from the stopper added to the solution.
   (c) Allowing a small amount of cyclohexane to evaporate.

3. How much would the molar mass change if 1.000 g of the cyclohexane evaporates during part B? What would be the induced error?

EXPERIMENT

The iodine clock

Goal

The goal of this laboratory is to use the **method of initial rates** to calculate the rate of a chemical reaction. On top of that, the student will also visualize experimentally the effect of **temperature** and of a **catalyst** on the rate of a chemical reaction.

Background

The rate law expresses the progress of a chemical reaction in function of the concentration of the reactant

$$\text{Rate} = k[\text{A}]^{a}[\text{B}]^{b},$$

where k is the a proportionality constant called the **rate constant**, [A] and [B] are the respective concentrations for reactants A and B, and the exponents are the **orders of the reactants**, a way to establish their corresponding relevance in the reaction rate.

Initial rates method

The initial rates method is used to measure the rate for a chemical reaction and to find out the rate constant k and the order of the reactions. This method is based on the assumption that at the beginning of a chemical reaction there is no reverse reaction, i.e. forming reactants from the products, can take place, and therefore the rate is solely affected by the reactants concentrations.

The time for the reactants concentrations to change by the same amount is measured. Several experiments are carried out using all the same starting concentrations but for one reactant.

The beauty of this experiment is the method used to make the reactant's concentration changed by the same amount for all the experiments. This is probably the most often misunderstood part of this laboratory and it is key to understand it thoroughly.

The Iodine Clock

A watch system is a method used in the navy to schedule regular periods aboard. A ship's bell is used to indicate the end of each period, allowing the crew to know when it is time to switch their task, to make breaks or to have lunch.

A **chemical clock** is a chemical reaction with a sudden change in one of its properties, the tempo of which can be be adjusted by adjusting the concentrations. The **iodine clock** is a chemical clock marked by a sudden, almost magical, color change. For the present experiment the iodine clock will be connected to the reaction being studied and will indicate the moment our reaction reaches the change in concentration desired. It is very important to understand the mechanism of the two competing reactions involved in this experiment to fully understand this activity.

The reaction setting the time is:

$$\text{I}_2(aq) + 2\,\text{S}_2\text{O}_3{}^{2-}(aq) \longrightarrow \text{S}_4\text{O}_6{}^{2-}(aq) + 2\,\text{I}^-(aq)$$

This reaction consumes iodine, I_2. Iodine reacts with starch turning the solution dark. The starch will be used in this experiment in a way similar to the indicator used in a titration.

$$\underset{\text{black-dark blue}}{\overset{\text{starch with iodine}}{\textbf{starch} + \mathbf{I_2}(aq)}} + 2\,S_2O_3{}^{2-}(aq) \longrightarrow \underset{\text{transparent}}{\overset{\text{starch , no } I_2}{S_4O_6{}^{2-}(aq) + 2\,I^-(aq) + \text{starch}}}$$

The iodine clock reaction will be running in parallel with another reaction that produces iodine. In the presence of starch, iodine becomes dark blue, almost black. Therefore, as long as there is thiosulfate ($S_2O_3{}^{2-}$) in the beaker the iodine will react with the latter and keep the solution transparent. We could say that the timer is adjusted by the concentration of thiosulfate ($S_2O_3{}^{2-}$).

The reaction to be studied

In this experiment the student will determine the reaction rate of the reduction of potassium persulfate ($K_2S_2O_8$) with sodium iodide (NaI).

$$K_2S_2O_8(aq) + 2\,NaI(aq) \longrightarrow 2\,Na_2SO_4(aq) + I_2(aq)$$

As ionic compounds dissociate in aqueous solutions it is more accurate to write the **net ionic equation**. These are the ionic net equations for the two reactions, the reduction of potassium persulfate with sodium iodide and the iodine clock:

$$S_2O_8{}^{2-}(aq) + 2\,I^-(aq) \longrightarrow 2\,SO_4{}^{2-}(aq) + \mathbf{I_2(aq)} \qquad \leftarrow \textit{rate to be measured}$$

$$\Downarrow$$

$$\textit{iodine clock} \rightarrow \qquad \mathbf{I_2(aq)} + 2\,S_2O_3{}^{2-}(aq) \longrightarrow S_4O_6{}^{2-}(aq) + 2\,I^-(aq)$$

To summarize: The first reaction will be running producing I_2, which would turn dark blue in the presence of starch. However, the second reaction will avoid the color change by reducing the iodine into iodide. The solution will remain transparent as long as there is thiosulfate ($S_2O_3{}^{2-}$) available. Once the thiosulfate is finished, iodine will finally remain in the solution turning it dark blue. Since the amount of thiosulfate will be constant for all the experiments to be performed, the amount of iodine produced in the first reaction, and therefore the rectant's concentration change, will be the same each time the solution turns dark. The iodine clock is being used to schedule a constant change in the reactant's concentrations.

The experiment

The experiment today is divided in 3 parts:

Part A. Obtaining the reaction rate constant.

Part B. Evaluating the influence of the temperature on the reaction rate.

Part C. Evaluating the influence of a catalyst on the reaction rate.

Part A. Obtaining the reaction rate constant.

The reaction to be studied and its rate law are:

$$K_2S_2O_8(aq) + 2\,NaI(aq) \longrightarrow 2\,Na_2SO_4(aq) + I_2(aq),$$

$$\text{Rate} = k[S_2O_8{}^{2-}]^a[I^-]^b,$$

By means of the **initial rates method** the reactants orders, a and b, can be solved. Find the suitable combinations of experiments where only one initial concentration of $S_2O_8{}^{2-}$ or I^- changes and compare the rates using a fraction. The initial

concentrations must be recalculated using the volumes and their original concentrations.
Once the reactant orders are known and the initial concentrations of $S_2O_8{}^{2-}$ and I^- for each reaction have been recalculated, the reaction constant, k, can be solved.
See the example for more details. Three experiments are performed using different initial volumes:

	Volumes (mL)					⏱
	NaI	NaCl	$Na_2S_2O_3$	Starch	K_2SO_4	$K_2S_2O_8$
Exp. 1	2.0	2.0	2.0	1.0	2.0	2.0
Exp. 2	2.0	2.0	2.0	1.0	0	2.0
Exp. 3	4.0	0	2.0	1.0	2.0	4.0

Example

The reaction starts from a solution containing A (1 M) and B (0.5) M and D (1 M). There are two reactions occurring in the solution. A slow reaction were C is produced and a fast reaction where C is reduced to E ions by D:

$$A + B \longrightarrow 2\,C$$
$$C + D \longrightarrow E$$

An indicator is added, such that the solution will turn dark in the presence of substance C.
Three experiments are performed using different initial volumes and the following times for the solution to turn dark are measured;

	Volumes (mL)				*time*
	A	*B*	*D*	*Water*	*(s)*
Exp. 1	*4.0*	*2.0*	*2.0*	*2.0*	*20*
Exp. 2	*4.0*	*1.0*	*2.0*	*2.0*	*40*
Exp. 3	*2.0*	*1.0*	*2.0*	*4.0*	*160*

Calculate the reactant orders and the reaction rate constant.

Answer: First write the reaction rate expression: Rate = $k[\mathrm{A}]^a[\mathrm{B}]^b$
Recalculate the concentrations for each reactant. Use the formula for dilutions $M_1V_1 = M_2V_2$, where V_2 is the total volume $V_A + V_B + V_D + V_{water}$.
For example, to calculate the concentration of A in Experiment 1:

$$M_1V_1 = M_2V_2 \qquad M_2 = \frac{M_1V_1}{V_2} \qquad M_2 = \frac{1M4mL}{10mL} = 0.4M$$

	[A]	[B]	[D]	rate (Ms^{-1})
Exp. 1	0.4 M	0.2 M	0.2 M	$5\cdot10^{-3}$
Exp. 2	0.4 M	0.1 M	0.2 M	$2.5\cdot10^{-3}$
Exp. 3	0.2 M	0.1 M	0.2 M	$6.25\cdot10^{-4}$

The reaction rate is the change in concentration per unit time. The change in concentration is determined by the second reaction (the clock). For the first experiment it takes 20 seconds to turn black. In those 20 seconds all of D has reacted, i.e. 0.2 M in 20 seconds.
Following the stoichiometry of the reaction, if 1 mole of D reacts with 1 mole of C and every 2 moles of C are produced

when 1 mole of A reacts. The relation between D and A is 2:1.

$$1 \text{ mole of A} = 1 \text{ mole A} \frac{2 \text{ moles C}}{1 \text{ moles A}} \frac{1 \text{ mole D}}{1 \text{ mole C}} = 2 \text{ moles of D}$$

Therefore, 0.2 M of D corresponds to 0.1 M of A reacting in 20 s. The rate for experiment 1 is $5 \cdot 10^{-3}$ M/s.
The order of the reactants are calculated by comparing the rate laws. Each time, choose two experiments were only one parameter is different.

$$\frac{\text{Rate 2}}{\text{Rate 3}} = \frac{k\,[A]^a[B]^b}{k[A]^a[B]^b} = \frac{k\,0.4^a 0.1^b}{k\,0.2^a 0.1^b} = \frac{2.5 \cdot 10^{-3}}{6.25 \cdot 10^{-4}}$$

$$2^a = 4 \qquad a = 2$$

Similarly, using rate 2 over rate 1, b=1 is obtained.
The reactant order can be used to obtain the reaction rate.

$$Rate = k[A]^a[B]^b \qquad k = \frac{Rate}{[A]^a[B]^b} = \frac{5 \cdot 10^{-3} Ms^{-1}}{(0.4M)^2(0.1M)^1} = 0.15625 M^{-2}s^{-1}$$

Since experimentally it is more likely to obtain one slightly different k for each experiment, an average should be calculated.

Parts B. Evaluating the influence of the temperature on the reaction rate.

According to **Arrhenius equation**, the rate constant is affected by the temperature following the equation:

$$k = Ae^{-E_a/RT}$$

where A is a constant, E_a is the activation energy, R is the universal gas constant, and T is the temperature. A higher temperature would make the negative exponent smaller, and therefore the k would be greater. In other words, an increase in the temperature should accelerate the reaction. It is possible to calculate the activation energy measuring the reaction rates at two different temperatures. After calculating the respective rate constants these are compared in their logarithmic forms:

$$\ln(k) = -\frac{E_a}{RT} + \ln(A)$$

After subtracting the expression above for both temperatures, T_1 and T_2, the following equation is generated:

$$\ln\left(\frac{k_2}{k_1}\right) = \frac{E_a}{R}\left(\frac{1}{T_1} - \frac{1}{T_2}\right)$$

The values of k_1 and k_2 obtained respectively at T_1 and T_2 are used to calculate E_a. The temperatures should be given in Kelvins since $R = 8.3145$ J/K mol.

Part C. Evaluating the influence of a catalyst on the reaction rate.

A catalyst is a substance that without being consumed, accelerates the reaction. The catalyst generates a new reaction path with a lower activation energy. Looking at the Arrhenius equation, the effect of lowering the activation energy works in the same way as an increase of temperature.
The catalyst used in this experiment will be copper sulfate ($CuSO_4$). The Cu(II) cations are responsible for the catalytic effect.

Procedure

Part A. Obtaining the reaction rate constant.

☐ *Step 1:* – Obtain three large test tubes with their the corresponding stoppers, a suitable grid or rack, a Mohr pipet and a thermometer.

☐ *Step 2:* – Label the test tubes with numbers 1, 2 and 3.

☐ *Step 3:* – Read the Good Lab Practice box

Good Lab Practice

- ✍ Pipettes are always used together with a suction bulb or a syringe. **Never** such a chemical using the mouth!
- ✍ **Never** leave a pipette laying on the table **while connected to the syringe**. Remaining liquid can leak into the syringe. The syringe will be damage and the liquid will return contaminated to the pipette.
- ✍ Always wash pipettes with distilled water and rinse them using the same product to be measured. Such a small volume, disconnect the syringe and move the liquid about displacing the distilled water. Discard the rinsing liquid as waste. Rinse twice.

☐ *Step 4:* – Obtain about than 25 mL **0.20 M NaI** in a small beaker.

☐ *Step 5:* – Clean and rinse the pipette with NaI. Pipette 2.0 mL in test tube 1, 2.0 mL in test tube 2 and 4.0 mL in test tube 3.

☐ *Step 6:* – Obtain no more than 20 mL **0.20 M NaCl** in a small beaker.

☐ *Step 7:* – Clean and rinse the pipette with NaCl. Pipette 2.0 mL in test tube 1 and 2.0 mL in test tube 2. Nothing in test tube 3.

☐ *Step 8:* – Obtain no more than 20 mL **0.010 M $Na_2S_2O_3$** in a small beaker.

☐ *Step 9:* – Clean and rinse the pipette with $Na_2S_2O_3$. Pipette 2.0 mL in each test tube.

☐ *Step 10:* – Obtain aproximately 10 mL **2% starch solution** in a small beaker.

☐ *Step 11:* – Clean and rinse the pipette with starch. Pipette 1.0 mL in each test tube.

☐ *Step 12:* – Obtain no more than 20 mL **0.20 M K_2SO_4** in a small beaker.

☐ *Step 13:* – Clean and rinse the pipette with K_2SO_4. Pipette 2.0 mL in test tube 1, nothing in test tube 2 and 2.0 mL in test tube 3.

☐ *Step 14:* – Check that the volumes in the 3 test tubes make sense with the following table. Test tubes 1 and 3 should have the same volume. Test tube 2 should have slightly less volume.

	Volumes (mL)					⏱
	NaI	NaCl	$Na_2S_2O_3$	Starch	K_2SO_4	$K_2S_2O_8$
Exp. 1	2.0	2.0	2.0	1.0	2.0	2.0
Exp. 2	2.0	2.0	2.0	1.0	0	2.0
Exp. 3	4.0	0	2.0	1.0	2.0	4.0

☐ *Step 15:* – Using the thermometer, record the room temperature.

☐ *Step 16:* – Get the stopwatch ready. As soon as you add the first drop of $K_2S_2O_8$ in test tube 1 time should start to be measured.

☐ *Step 17:* – $K_2S_2O_8$ will be added and measured one test tube at a time.

☐ *Step 18:* – Obtain about 25 mL **0.20 M $K_2S_2O_8$** in a small beaker.

☐ *Step 19:* – **IMPORTANT:** Every reaction should be done in **exactly the same way** in order to be comparable. The same steps, in the same order and at the same pace should be followed. After adding the $K_2S_2O_8$, the test tube should be closed with the stopper and the content should be mixed by inverting the test tube 16 times. Do not rush, turn the test tube *Mississippily*. Do exactly the same for the following reactions.

⚠CAUTION!

⚠ The stopper should always be placed before inverting the test tube.

⚠ Failure to place the stopper will cause the spill of the reaction.

⚠ There is a story behind every Caution Box.

☐ *Step 20:* – Clean and rinse the pipette with $K_2S_2O_8$. Start the stopwatch as you pipette 2.0 mL of $K_2S_2O_8$ in test tube 1. Mix the reaction as explained in Step 19.

☐ *Step 21:* – Stop the stopwatch when the solution turns dark. The color will appear suddenly. If more than 5 minutes elapse without the solution turning dark, start doubting about what have you mixed.

☐ *Step 22:* – Get test tube 2 and a stopper. Set the time to zero.

☐ *Step 23:* – Start the stopwatch as you pipette 2.0 mL of $K_2S_2O_8$ in test tube 2. Mix the reaction as explained in Step 19.

☐ *Step 24:* – Stop the stopwatch when the solution turns dark.

☐ *Step 25:* – Get test tube 3 and a stopper. Set the time to zero.

☐ *Step 26:* – Start the stopwatch as you pipette 4.0 mL of $K_2S_2O_8$ in test tube 3. Mix the reaction as explained in Step 19.

☐ *Step 27:* – Stop the stopwatch when the solution turns dark.

☐ *Step 28:* – Discard the solutions in the indicated waste container and clean the test tubes.

☐ *Step 29:* – Do a second trial for the three test tubes.

Part B. Evaluating the influence of the temperature on the reaction rate.

☐ *Step 30:* – Label three large test tubes with the letters “**T**”, “**C**”’ and “**K**”’. “T” stands for “Temperature”, “C” for “Catalyst” and “K” for “$K_2S_2O_8$”.

☐ *Step 31:* – Add to test tubes “T” and “C” the same volumes of NaI, NaCl, $Na_2S_2O_3$, Starch and K_2SO_4 as for test tube 1. Do not add the $K_2S_2O_8$ yet. Reserve test tube “C” in the rack for Part C.

☐ *Step 32:* – Add about 5 mL of $K_2S_2O_8$ to test tube “K”.

☐ *Step 33:* – Obtain a 600 mL beaker and, with the help of the thermometer, add hot and cold water until you have about 400 mL of water between 40 and 45 degrees Celsius.

☐ *Step 34:* – Place the test tubes “T” and “K” in the hot water bath and allow them to temper for at least 5 minutes. Part C can be done during those 5 minutes.

☐ *Step 35:* – Measure the temperature of the water after the tempering time. Start the reaction by pipetting 2 mL of $K_2S_2O_8$ from test tube “K” into test tube “T”. After the 16-Mississippi mixing, return the test tube to the bath and wait to record the time it takes for the solution to turn dark.

Part C. Evaluating the influence of a catalyst on the reaction rate.

☐ *Step 36:* – Add **one drop** of **0.2 M $CuSO_4$** to test tube “C”. Shake the test tube to mix the solution.

☐ *Step 37:* – Start the stopwatch as you pipette 2.0 mL of $K_2S_2O_8$ in the test tube. Mix the reaction as explained in Step 19.

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

The iodine clock

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1. Does the concentration of the reactants affect the rate of the reaction?

2. What is the general expression for the reaction 2A+ 3B → C?

3. What is the Iodine clock reaction used for in this experiment?

4. In the example, is it relevant that the second reaction is faster than the first one?

5. What is starch and where can you find it in nature?

6. Low quality boiled ham often contains potato to make it cheaper. How could you test if one piece of ham is pure meat or it it contains potato?

7. Find out the difference between a **Mohr pipette** and a **transfer pipette**.

STUDENT INFO

Name: ____________________ Date: __________

Results

EXPERIMENT

The iodine clock

Parts A. Obtaining the reaction rate constant.

Trial 1	Time (s)	Trial 2	Time (s)	Average Time (s)
Test tube 1	_________	Test tube 1	_________	_________
Test tube 2	_________	Test tube 2	_________	_________
Test tube 3	_________	Test tube 3	_________	_________

Laboratory Temperature ∘C _______

Parts B. Evaluating the influence of the temperature on the reaction rate.

	Time (s)
Test tube “T”	_________

Parts C. Evaluating the influence of a catalyst on the reaction rate.

	Time (s)
Test tube “C”	_________

Water Temperature _______ ∘C

STUDENT INFO

Name: ______________ Date: __________

Post-lab Questions

The iodine clock

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1. Recalculate the initial concentrations.

	Concentrations (M)					
	NaI	NaCl	$Na_2S_2O_3$	Starch	K_2SO_4	$K_2S_2O_8$
Exp. 1	______	______	______	______	______	______
Exp. 2	______	______	______	______	______	______
Exp. 3	______	______	______	______	______	______

2. Calculate the initial rates for each experiment. Use the average times from both trials for each experiment. Use the correct units.

	Test tube 1	Test tube 2	Test tube 3
$\frac{\Delta[S_2O_8{}^{2-}]}{\Delta t}$			

3. Write the expression for the reaction rate and calculate the reaction orders.

4. Calculate the rate constant *k* for experiments 1, 2 and 3. Find the average value. Use the correct units.

5. Calculate the rate constant *k* for the experiment “T”. Use the correct units.

6. Calculate the Activation Energy for the reaction. Use the correct units.

7. What was the effect of the catalyst?

EXPERIMENT

Le Châtelier's Principle

Goal

The goal of this laboratory is to see the shift in the **equilibrium position** and to connect the observations to **Le Châtelier's Principle**, which states that:

Any change in status quo prompts an opposing reaction in the responding system.

Background

A double arrow in a chemical reaction indicates that the reaction can proceed in both directions. Products can react with each other to generate the original reactant.

$$A + B \rightleftharpoons C + D$$

Since the speed of most reactions depend on the concentration of the combining spices, the forward reaction rate will decrease the more products are formed, and the reverse reaction rate will increase. The later will replenish the concentration of reactants which in turn will accelerate the forward reaction. Eventually forward and backward reaction rates will equal each other and the concentrations of reactants and product will remain constant; equilibrium has been reached.

The experiment

The experiment today is divided in 5 parts:

Part A. Colored complexed ions. $FeSCN^{2+}$

Part B. Colored complexed ions. $Ni(NH_3)_6{}^{2+}$.

Part C. Effect of pH on an indicator.

Part D. Effect of pH on solubility

Part E. Heat as a product

Part A. Colored complexed ions; $FeSCN^{2+}$.

Many metals form colored complex ions with several ligands, such as iron (III) ion Fe^{3+} with thiocyanate SCN:

$$\underset{\text{yellow}}{\mathrm{Fe^{3+}}\,(aq)} + \mathrm{SCN^-}\,(aq) \rightleftharpoons \underset{\text{red}}{\mathrm{FeSCN^{2+}}\,(aq)}$$

The Iron (III) solution is yellow, the thiocyanate is transparent and the Iron (III) thiocyanate is red. The final color of the solution will be determined by the concentrations in the final equilibrium position.

Part B. Colored complexed ions; $Ni(NH_3)_6{}^{2+}$

.

In the case of Hexaaminenickel (II) you will also study the change in equilibrium position when adding an acid. The ligand will act as a base reacting with the acid and detaching from the metal.

$$\underset{\text{}}{\overset{\text{green}}{Ni^{2+}\,(aq)}} + 6\,NH_3\,(aq) \rightleftharpoons \overset{\text{blue}}{\mathbf{Ni(NH_3)_6{}^{2+}\,(aq)}}$$

$$\Downarrow$$

$$\underset{\text{blue}}{\mathbf{Ni(NH_3)_6{}^{2+}\,(aq)}} + 6\,HCl\,(aq) \rightleftharpoons \underset{\text{green}}{NiCl_2\,(aq)} + 6\,NH_4Cl\,(aq)$$

Part C. Effect of pH on an indicator.

pH indicators are substances that change color depending on the medium pH. Typically they are weak bases or acids. As such, they dissociate slightly in water forming ions. Those ions might be colored spices. The general expression for the dissociation of a divalent weak acid indicator (H2In) is:

$$H_2In\,(aq) + H_2O\,(aq) \longrightarrow HIn^-\,(aq) + H_3O^+\,(aq)$$

In the case of Methyl orange, the protonated form is red, while the conjugated base is yellow. Notice that hydronium is a product in this equilibrium. How will adding an acid affect this equilibrium?

Part D. Effect of pH on solubility.

Hydroxide salts of Group II elements (Ca, Sr, and Ba) are slightly soluble. Adding an acid will neutralize the hydroxides in solution, while adding more hydroxide will increase the ion product. In both cases the amount of precipitate will be affected.

$$Ca(OH)_2\,(s) \rightleftharpoons Ca2+\,(aq) + OH^-\,(aq)$$

$$\Downarrow$$

$$OH^-\,(aq) + H^+\,(aq) \rightleftharpoons H_2O\,(aq)$$

Part E. Heat as a product

Lastly, you will experience Le Châtelier's Principle in an exothermic reaction, where the heat can be considered a product of the reaction.

$$\underset{\text{violet}}{CoCl_4{}^{2-}\,(aq)} + 6\,H_2O\,(l) \rightleftharpoons \underset{\text{pink}}{Co(H_2O)_62+\,(aq)} + 4\,Cl^-\,(aq) + \mathbf{HEAT}$$

Example

The following endothermic reaction is allowed to reach equilibrium:

$$\underset{\text{blue}}{A(aq)} + B(aq) + HEAT \rightleftharpoons C(aq) + D(aq)$$

Where A, only colored compound, is blue.
How will the following changes affect the color/equilibrium?
(a) Adding more B reactant. (b) Adding more C product. (c) Heating the mixture.

Answer: (a) less blue/equilibrium shifts to the right. (b) more blue/ equilibrium shifts left. (c) Heat is a reactant; less blue/equilibrium shifts to the right.

Procedure

Part A. Colored complexed ions; $FeSCN^{2+}$

☐ *Step 1:* – Find 3 tests tubes and a 100 mL beaker, clean them and mark the test tubes with letter A, B and R.

☐ *Step 2:* – Read the Good Lab Practice box

Good Lab Practice

✍ Concentrated solutions of ammonia, sodium hydroxide and hydrochloric acid should be handled with care.

✍ Do not shake open test tubes. Do not use your fingers as stoppers, even if you wear gloves.

☐ *Step 3:* – Add about 20 mL of distilled water to the beaker, 20 drops of 0.1 M $Fe(NO_3)_3$ and 20 drops of 0.1 M KSCN. Mixed the solution until the color is homogeneous.

☐ *Step 4:* – Use a 10 mL graduated cylinder to add 3 mL of the mixture to each of the test tubes.

☐ *Step 5:* – Add 20 drops of 0.1 M $Fe(NO_3)_3$ to test tube A. Put a stopper and mix the solution.

☐ *Step 6:* – Add 20 drops of 0.1 M KSCNto test tube B. Put a stopper and mix the solution.

☐ *Step 7:* – Add 20 drops of distilled water to test tube R. Put a stopper and mix the solution.

☐ *Step 8:* – Compare the color of the test tubes A and B to R and write down your observations.

Part B. Colored complexed ions; $Ni(NH_3)_6{}^{2+}$.

☐ *Step 9:* – Obtain 1 test tube and clean it.

☐ *Step 10:* – Add 10 drops of 0.1 M $Ni(NO_3)_2$. Indicate the color in the results page.

☐ *Step 11:* – Add drops of 6 M NH_3 until the color changes.

☐ *Step 12:* – Add drops of 6 M HCl until the color changes.

Part C. Effect of pH on an indicator.

☐ *Step 13:* – Obtain two 50 mL beakers, clean them and mark them with letters A (for Acid) and B (for Base). Find and clean a test tube.

☐ *Step 14:* – Add 10 mL of distilled water to each beaker and 1 mL of distilled water to the test tube.

☐ *Step 15:* – Add 4 drops of 6 M HCl to beaker A and stir it.

☐ *Step 16:* – Add 4 drops of 6 M NH_3 to beaker B and stir it.

☐ *Step 17:* – Add 4 drops of the indicator to the test tube.

☐ *Step 18:* – Add 2 drops from the diluted acid in beaker A to the test tube. Mix kindly the solution. Record the color in the results page.

☐ *Step 19:* – Add drop by drop the diluted ammonia solution from beaker B until the color changes. Homogenize the solution from time to time.

☐ *Step 20:* – Attempt another color change by adding this time the diluted acid solution from beaker A.

Part D. Effect of pH on solubility.

☐ *Step 21:* – Obtain two 50 mL beaker, an small Erlenmeyer, a 10 mL graduated cylinder, a plastic funnel and a filter paper.

☐ *Step 22:* – Add 5 mL of 6 M NaOH to the beaker using the graduated cylinder.

☐ *Step 23:* – Rinse the cylinder with water 3 times. Then, use it to add 5 mL of 1 M $Ca(OH)_2$ to the beaker.

☐ *Step 24:* – Stir the mixture.

☐ *Step 25:* – Make a cone with the filter paper and place it in the funnel, on top of the Erlenmeyer. Filter the solution with the precipitate. After a couple of washings transfer the white solid to a small clean beaker.

☐ *Step 26:* – Add 10 mL of distilled water to the beaker with the white solid. Stir the mixture. Do not expect the solid to completely dissolve. The solution is saturated.

☐ *Step 27:* – Add drops of 6 M HCl until a change is observed. Record your results.

☐ *Step 28:* – Add drops of 6 M NaOH until a change is observed. Record your results.

Part E. Heat as a product.

☐ *Step 29:* – Obtain a bunsen burner, a stand, two iron rings, a wire gauze, two 250 mL beaker and a test tube.

☐ *Step 30:* – Prepare the setup for the bunsen burner, using the second iron ring over the wire gauze to protect the beaker from falling. Put about 100 mL distilled water in the beaker and bring the water to boil.

☐ *Step 31:* – Add 5 drops of 0.1 M $Co(NO_3)_2$ to the test tube.

☐ *Step 32:* – Record the color of the liquid solution after each step.

☐ *Step 33:* – Add drops of 12 M HCl until the color of the solution changes. You might need to stir the test tube to help mixing the reagents.

☐ *Step 34:* – Add 5 drops of distilled water and mix.

☐ *Step 35:* – Place the test tube in the boiling water and wait for another color change.

☐ *Step 36:* – Attempt to reverse the reaction by placing the test tube in a beaker with ice.

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Le Châtelier's Principle

1. Write the equilibrium constant expressions for the main reactions (5) in each part.
2. In a chemical reaction with two reactants, one is added in large excess. How will the concentration of the other reactant become?
3. How is supposed to affect the pH the solubility of $Ca(OH)_2$?
4. An endothermic reaction needs external energy or heat to happen. Will HEAT be considered a product or a reactant?
5. In an exothermic reaction, is HEAT a product or a reactant? Why?

STUDENT INFO

Name: ______________________ Date: __________

Results

EXPERIMENT

Le Châtelier's Principle

Part A. Colored complexed ions; $FeSCN^{2+}$

After step #	Indicate color of	Color
5	Solution in test tube A	____________
6	Solution in test tube B	____________
7	Solution in test tube A	____________

Part B. Colored complexed ions; $Ni(NH_3)_6^{2+}$.

After step #	Indicate color of solution	
10	__________________	
11	__________________	# drops added ___
12	__________________	# drops added ___

Part C. Effect of pH on an indicator.

After step #	Indicate color of solution	
17	__________________	
18	__________________	
19	__________________	# drops added ___
20	__________________	# drops added ___

Part D. Effect of pH on solubility.

After step # Indicate # drops added

27 ___

28 ___

Part E. Heat as a product.

After step # Indicate color of solution

31 ____________________

33 ____________________ # drops added ___

34 ____________________

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Le Châtelier’s Principle

1. In part A. How do the different solutions compare? Explain the different colors based on the equilibrium reaction and Le Châtelier’s Principle.
2. In part B. What did you observe after adding the acid and after adding the base? Explain the different colors based on the equilibrium reaction and Le Châtelier’s Principle.
3. In Part C. What did you observe after adding the acid and after adding the base? Explain the changes based on the equilibrium reaction and Le Châtelier’s Principle.
4. In Part D. What did you observe after adding the acid and after adding the base? Explain the changes based on the equilibrium reaction and Le Châtelier’s Principle?
5. In Part E. What did you observe after heating and after cooling? Explain the changes based on the equilibrium reaction and Le Châtelier’s Principle.

EXPERIMENT

Equilibrium constant

Goal

The goal of this laboratory is to determine **the equilibrium constant** of a chemical reaction. Additionally **Lambert-Beer's law** will be used to measure the concentration of the species in solution.

Background

Often, chemical reactions do not proceed to completion, instead only proceed to a certain point where the concentrations of both reactants and products remain constant over the time. The reason for this behavior is that most reactions are reversible, which means that they can run in both forward and reverse directions. The concentrations of the species will achieve an **equilibrium** state when the reaction rates in both directions equalize. For a certain chemical

$$\mathrm{aA} + \mathrm{bB} \rightleftharpoons \mathrm{cC} + \mathrm{dD}$$

the equilibrium can be characterized by the expression:

$$K = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

where K (capital letter) is the equilibrium constant, the brackets represent the concentrations of the respective species and the exponents are their corresponding stoichiometric coefficients.

Iron (II) thiocianate

The reaction to be studied is the formation of the ion **iron (III) thiocyanate** ($(FeSCN)^{2+}$) from the **iron (III) cation** (Fe^{3+}) and the **thiocyanate anion** (SCN^-). Interestingly, Fe^{3+} is yellow in solution, while it becomes red when forming the complex $(FeSCN)^{2+}$. This coloration allows us to use **spectrophotometry** to measure the concentration of all the species in equilibrium.

Lambert-Beer's law

Light intensity passing through a colored solution is **attenuated** by the liquid. The effect can be comparable to a dirty window. The dirtier the window (higher concentration of dirt), the less one can see through it (because more light is absorbed by the dirt). This effect is simplified in the Lambert-Beer's law:

$$A = k_\lambda c$$

In this formula, A stands for the **absorbance** (light absorbed by the liquid), k_λ (lower case, not to be confused with the equilibrium constant, K) is a **proportionality constant**, which is specific to each compound and at the wavelength (λ) specified by the subindex, and c is the solution concentration.

Lambert-Beer's law is a linear function of the form $y = mx + b$. When graphing the absorbance (y-axes) versus the concentration (x-axes), the slope of the line corresponds to the Lambert-Beer's proportionality constant.

Spectrophotometry

Spectrophotometry is a technique that measures the amount of light absorbed by a chemical substance, typically in solution. The **spectrophotometer** measures the light intensity passing through the sample and establishes the absorbance comparing the measurement to a **reference** called the **blank**.

The **blank** is a sample containing only the solvent and no solute. The spectrometer sets the blank as the baseline of the measurement, eliminating any absorbance caused by the solvent.

The spectrometer allows to measure absorbance at a single wavelength. Interestingly, for each substance there is a wavelength of **maximum absorbance**. This is of great use because it allows a differentiation among chemical substances with different maximum absorbances.

For instance, iron thiocyanate has an absorbance maximum at around λ=450 nm, while the iron (III) cation's absorbance at that wavelength is negligible.

The experiment

The experiment today is divided in 2 parts:

Part A. Determining Lambert-Beer's constant k.

Part B. Determining the equilibrium constant K.

Part A. Determining Lambert-Beer's constant k.

The reaction to be studied is:

$$\mathrm{Fe}^{3+}(aq) + \mathrm{SCN}^{-}(aq) \rightleftharpoons \mathrm{Fe(SCN)}^{2+}(aq)$$

The first part of this experiment requires the determination of Lambert-Beer's constant, k, by representing the absorbance for a series of known concentrations of $\mathrm{Fe(SCN)}^{2+}$. However, and according to the equilibrium, $\mathrm{Fe(SCN)}^{2+}$ in solution will follow the reverse reaction forming iron (III) and thiocyanate ions. The shift due to the equilibrium makes the concentration of $\mathrm{Fe(SCN)}^{2+}$ unknown. How can this problem be avoided?

The answer is in the **Le Châtelier's principle**. According to Le Châtelier's, the equilibrium shifts to the products when one reactant is added. When doing so, the other reactant will be consumed proportionally. If one of the reactants is added in overwhelming excess, the other reactant will be consumed almost to exhaustion. The latter will be the limiting reactant, it can be considered that all the limiting reactant will react and the amount of product formed can be calculated stoichiometrically.

For the present experiment, Fe^{3+} ion will be added in excess, while only small amounts of SCN^{-} will be added. The concentration of $\mathrm{(FeSCN)}^{2+}$ will be calculated using the concentration of the anion.

	KSCN 10^{-3} M	$\mathrm{Fe(NO_3)_3}$ 0.25 M	$\mathrm{HNO_3}$ 0.1 M
Test Tube 1	1.0 mL	5.0 mL	4.0 mL
Test Tube 2	2.0 mL	5.0 mL	3.0 mL
Test Tube 3	3.0 mL	5.0 mL	2.0 mL
Test Tube 4	4.0 mL	5.0 mL	1.0 mL
Test Tube 5	5.0 mL	5.0 mL	0.0 mL

The 10^{-3} M solution of KSCN must be previously prepared by dilution from a 0.0025 M solution.

Each solution is added to a clean cuvette and absorbance at λ=450 nm is measured using the spectrophotometer. With the 5 points obtained and the origin of coordinates, a graph can be plotted and the slope calculated. Notice that in this linear representation, where y-intercept (b) is zero, the origin can be considered as a valid point. The line must pass through the origin by definition, since at concentration=0, absorbance must be 0.

Parts B. Determining the equilibrium constant K.

The following solutions will be prepared and measured by the spectrophotometer:

	$Fe(NO_3)_3$ 0.25 M	KSCN 0.025 M	HNO_3 0.1 M
Test Tube 6	1.0 mL	5.0 mL	1.0 mL
Test Tube 7	1.0 mL	4.5 mL	1.5 mL
Test Tube 8	1.0 mL	4.0 mL	2.0 mL
Test Tube 9	1.0 mL	3.5 mL	2.5 mL
Test Tube 10	1.0 mL	3.0 mL	3.0 mL
Test Tube 11	2.0 mL	4.0 mL	1.0 mL
Test Tube 12	2.0 mL	3.5 mL	1.5 mL
Test Tube 13	2.0 mL	3.0 mL	2.0 mL
Test Tube 14	2.0 mL	2.5 mL	2.5 mL
Test Tube 15	2.0 mL	2.0 mL	3.0 mL

Using the absorbance measured and the value of k_{450nm}, it is possible to calculate the concentration in equilibrium of $Fe(SCN)^{2+}$ for each system.

$$A = k_{450nm}[Fe(SCN)^{2+}] \qquad [Fe(SCN)^{2+}] = \frac{A}{k_{450nm}}$$

The initial-change-equilibrium table allows to calculate the concentration in equilibrium of the rest of the species.

	$[Fe(NO_3)_3]$	[KSCN]	$[Fe(SCN)^{2+}]$
Initial	①	②	0 M
Change	+③	+③	−③
Equilibrium	①−③	②−③	③

③ is calculated from the absorbance. ② and ① are recalculated from the initial concentrations.

Procedure

Part A. Determining Lambert-Beer's constant k.

☐ *Step 1:* – Clean 5 large test tubes and the matching stoppers and set them upside down on a rack. Number the test tubes from 1 to 5.

☐ *Step 2:* – Obtain a 100.0 mL volumetric flask and a Mohr pipette.

☐ *Step 3:* – Obtain about 15 mL of 0.0025 M KSCN in a 50 mL beaker. Rinse the Mohr pipette a couple of times with a small amount of this solution. Always discard the rinsing liquid. Then, transfer 4.00 mL into the volumetric flask.

☐ *Step 4:* – In a clean 100 mL beaker, obtain about 100 mL of distilled water. Carefully add water to the volumetric flask directly from the beaker and STOP adding before you reach the 100.0 mL mark. Use a plastic dropper to level the water to the mark, drop by drop.

☐ *Step 5:* – Rinse the Mohr pipette with the diluted solution of KSCN again.

☐ *Step 6:* – Following the Table in the description of Part A, add 1.00 mL of the diluted KSCN solution to test tube 1; 2.00 mL to test tube 2, 3.00 mL to test tube 3; 4.00 mL to test tube 4 and 5.00 mL to test tube 5.

Good Lab Practice

- Pipettes are always used together with a suction bulb or a syringe. **Never** suck a chemical using the mouth!
- **Never** leave a pipette lying on the table **while connected to the syringe**. Remaining liquid can leak into the syringe. The syringe will be damaged and the liquid will return to the pipette contaminated.
- Always wash pipettes with distilled water and rinse them using the same product to be measured. Suck a small volume, disconnect the syringe and move the liquid about, displacing the distilled water. Discard the rinsing liquid as waste. Rinse twice.

☐ *Step 7:* – Rinse the pipet with distilled water.

☐ *Step 8:* – Obtain about 50 mL of 0.25 M $Fe(NO_3)_3$ in a 100 mL beaker. Rinse the pipette with the solution of $Fe(NO_3)_3$ from the beaker.

☐ *Step 9:* – Add 5.0 mL of 0.25 M $Fe(NO_3)_3$ in each test tube.

☐ *Step 10:* – Rinse the pipette with distilled water.

☐ *Step 11:* – Start the spectrophotometer. The bulb needs time to heat up to a stable temperature.

☐ *Step 12:* – Obtain about 40 mL of 0.1 M HNO_3 in a 50 mL beaker. Rinse the pipette with the solution of HNO_3 from the beaker.

☐ *Step 13:* – Following the Table in the description of Part A, add 4.00 mL of the HNO_3 solution to test tube 1; 3.00 mL to test tube 2, 2.00 mL to test tube 3 and 1.00 mL to test tube 4.

☐ *Step 14:* – At this point, all your test tubes should have exactly the same volume of liquid. If not, repeat the ones that diverge.

☐ *Step 15:* – Put the stoppers in the test tubes and mix the solutions.

Good Lab Practice

- For best results, use the same cuvette, with the same orientation, for all the measurements.
- Keep the outer walls of the cuvette clean. Do not touch them. Wipe them with a piece of optical paper.
- Rinse the cuvette at the same time as you rinse the pipette.

☐ *Step 16:* – Get a 250 mL beaker for waste. Refill the 100 mL beaker with distilled water.

☐ *Step 17:* – Using the plastic dropper, fill the cuvette 3/4 full with distilled water.

☐ *Step 18:* – Insert the cuvette into the spectrophotometer, set the wavelength to 450 nm and press "blank" or "zero".

☐ *Step 19:* – Discard the water. Now rinse the dropper and the cuvette:

1. Fill the cuvette 3/4 full with the solution from test tube 1.
2. Discard the solution
3. Fill the cuvette 3/4 full with the solution from test tube 1.

4. Discard the solution.

☐ *Step 20:* – Insert the cuvette into the spectrophotometer to measure the absorbance. Record the result.

☐ *Step 21:* – Repeat steps 19 and 20 for test tubes 2, 3, 4 and 5.

STUDENT INFO

Name: ______________ Date: ________

Pre-lab Questions

Equilibrium constant

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1. Find the definition of the following concepts:
   - (a) Equilibrium constant
   - (b) Le Châtelier’s Principle
   - (c) Light absorbance
   - (d) Wavelength
   - (e) Spectrophotometer
   - (f) Complex ion
2. Write the expression for Lambert-Beer’s law and identify the different parameters.
3. Write the equilibrium expression for the reaction: $Fe^{3+}(aq) + SCN^{-}(aq) \rightleftharpoons Fe(SCN)^{2+}(aq)$

STUDENT INFO

Name: ____________________ Date: __________

Results
EXPERIMENT

Equilibrium constant

Parts A. Determining Lambert-Beer's constant k.

	A
Test tube 1	
Test tube 2	
Test tube 3	
Test tube 4	
Test tube 5	

Parts B. Determining the equilibrium constant K.

	A
Test tube 6	
Test tube 7	
Test tube 8	
Test tube 9	
Test tube 10	

	A
Test tube 11	
Test tube 12	
Test tube 13	
Test tube 14	
Test tube 15	

STUDENT INFO

Name: ________________________ Date: __________

Post-lab Questions

Equilibrium constant

The questions serve as a guide for solving Parts A and B. If you develop the answer in your Data or Conclusions, there is no need to repeat the answer in this section.

1. Recalculate the initial concentrations for KSCN and $Fe(NO_3)_3$

	$[KSCN]_0$	$[Fe(NO_3)_3]_0$
Test Tube 1	________	________
Test Tube 2	________	________
Test Tube 3	________	________
Test Tube 4	________	________
Test Tube 5	________	________

2. Apply Le Châtelier's Principle to determine $[Fe(SCN)^{2+}]$.

	$[Fe(SCN)^{2+}]$
Test Tube 1	________
Test Tube 2	________
Test Tube 3	________
Test Tube 4	________
Test Tube 5	________

3. Plot the graph A vs [$Fe(SCN)^{2+}$]. Do not forget to include the Origin.

4. Applying the Initial-Change-Equilibrium method, calculate the concentrations in equilibrium and K.

	$[Fe^{3+}]_0$	$[SCN^-]_0$	$[Fe(SCN)^{2+}]$	$[Fe^{3+}]_{eq}$	$[SCN^-]_{eq}$	K
Test Tube 6						
Test Tube 7						
Test Tube 8						
Test Tube 9						
Test Tube 10						
Test Tube 11						
Test Tube 12						
Test Tube 13						
Test Tube 14						
Test Tube 15						

5. Calculate the mean value for K.

EXPERIMENT

Bleach Strength

Goal

The goal of this laboratory is to perform an **oxidation-reduction titration** in order to determine the concentration of sodium hypochlorite (NaOCl) in commercial bleach samples.

Background

The properties of chlorine for bleaching fabrics were discovered in the 18th century by the French scientists Claude Berthollet. He was able to make an aqueous solution of sodium hypochlorite which was named after the quarter in Paris where it was produced, Javel. *Eau de Javel* (Javel water) is still used in French to name bleach.

In the beginning of the 18th century, the pharmacist Antoine Germain Labarraque discovered the disinfecting ability of another hypochlorite solution, this time calcium hypochlorite. The so called "Eau de Labarraque" was employed in the disinfection of animal guts to make musical instrument strings. Hypochlorite was the first antiseptic product, helping to treat gangrene and putrescent wounds in people in the 1820s, and ending with the propagation of "cadaveric particles" in morgues.

Nowadays bleach is a common product at home, used to whiten clothes, remove stains and as a disinfectant. Because of its strong bactericidal properties, hypochlorite is also used for disinfecting and to prevent the proliferation of algae in swimming pools.

Chlorine

Chlorine forms four different polyatomic anions with oxygen:

$$\begin{aligned}
&\text{ClO}^- hypochlorite\\
&\text{ClO}_2{}^- chlorite\\
&\text{ClO}_3{}^- chlorate\\
&\text{ClO}_4{}^- perchlorate
\end{aligned}$$

These are called **oxoanions** and illustrate 4 possible oxidation states of chlorine atom; +1, +3, +5 and +7 respectively. When the oxidation state of the chlorine decreases, for instance going from $\text{ClO}_4{}^-$ to $\text{ClO}_3{}^-$ (+7 → +5), it reduces. For reduction to happen another species must be oxidized. Chlorine is then acting as the **oxidizing agent**. Perchlorate it is the strongest oxidizing agent of the 4 oxoanions, but even ClO^- can still be reduced to Cl- (Cl oxidation state -1).

The first oxidation

In order to determine the concentration of ClO^-, two consecutive oxidation-reduction reactions will be performed. In the first reaction, all the ClO^- present in the sample will be oxidized into Cl^-. The **oxidizing agent** will be I^-, which will be in turn reduced into I_2. The reaction, carried out in acidic medium, will read as:

—

$$\text{H}^+(aq) + \underline{\ \text{ClO}^-}(aq) + \underline{\ \text{I}^-} \longrightarrow \underline{\ \text{I}_2}(aq) + \underline{\ \text{Cl}^-}(aq) + \underline{\ \text{H}_2}\text{O}(aq)$$

⚠ Note: The reaction is NOT balanced. Determine the number of electrons being transferred using the oxidation states (in the prelab) to help you balance the reaction.

An excess amount of I^- will be used to ensure that all the ClO^- is oxidized. The iodine (I_2) formed in this reaction will be therefore limited by the amount of hypochlorite present in the sample. Keep in mind that the amount of hypochlorite is what we want to determine, and it can be stoichiometrically calculated from the amount of iodine produced.

In the following step the amount of I_2 will be determined using an oxidation-reduction titration.

The solution might turn slightly brownish due to the presence of a Lewis complex between iodine and water.

The oxidation-reduction titration

The titration to be performed accounts for the following redox reaction:

—

$$I_2(aq) + \underline{\;}S_2O_3{}^{2-}(aq) \longrightarrow \underline{\;}I^-(aq) + \underline{\;}S_4O_6{}^{2-}(aq)$$

⚠ Again, this reaction must be balanced.

Iodine (I_2) is reduced back to iodide I^-. Thiosulfate $S_2O_3{}^{2-}$ is the reducing agent, being itself oxidized into tetrathionate ions ($S_4O_6{}^{2-}$).

The titration endpoint will be determined by adding starch as an indicator at the beginning of the reaction. Remember from previous experiments that starch and I_2 react turning the solution into a dark-blue color, therefore, the titration should be stopped when all the iodine has been reduced and the solution becomes transparent.

The experiment

The experimental part consists of a trial titration and an exact titration.

Part A. Trial titration.

Part B. Exact titration.

Part A. Trial titration.

The trial titration will be used to estimate the volume added at the endpoint of the titration. It is conducted quickly, adding portions of 1 mL until the endpoint is exceeded. The final result is not used for calculations.

Part B. Exact titration.

In the exact titration an initial amount is added rapidly. This amount is calculated by subtracting 5 mL from the volume used in the trial titration. Like this it is possible to stop before the endpoint of the titration and to continue carefully, adding the titrant in a drop-by-drop basis until the precise endpoint.

If time allows, the exact titration should be repeated and the mean volume used in calculations.

Procedure

Part A. Trial titration.

☐ *Step 1:* – Obtain a buret, a stand, and a clamp. Rinse the buret with distilled water and make sure the valve works properly.

☐ *Step 2:* – In a 100 mL beaker, obtain 70 mL of 0.0250 M solution of $Na_2S_2O_3$. Use a small portion of the solution to rinse the buret. Discard the rinsing liquid and fill up the buret above the zero-mark. Bring the volume in the buret to zero or below, discarding the excess liquid trough the tip of the buret.

☐ *Step 3:* – In a 125 mL Erlenmeyer, obtain about one inch of a spatula of KI.

⚠CAUTION!

⚠ Bleach is an acid.

⚠ Acids can cause burns and clothes color loss.

⚠ Always wear eye protection and handle carefully.

!

☐ *Step 4:* – In a 50 mL beaker, obtain about 30 mL of a 10% bleach solution.

☐ *Step 5:* – Obtain a 10 mL transfer pipet. Pipet 10.0 mL of the diluted bleach into the KI containing Erlenmeyer. Add 20 mL of distilled water from a graduated column and 20 drops of HCl 2 M. Homogenize the mixture by stirring the flask.

☐ *Step 6:* – Record the buret initial reading to the closest hundredth of mL.

☐ *Step 7:* – Place the Erlenmeyer under the buret and start the titration by adding 1 mL portions until the solution turns from brown to yellow.

☐ *Step 8:* – Add 40 drops of a 0.2% starch solution to the yellow solution in the Erlenmeyer.

☐ *Step 9:* – Proceed with the titration adding the $Na_2S_2O_3$ in 1 mL portions until it turns from dark color to colorless.

☐ *Step 10:* – Record the buret final reading and calculate the volume added.

☐ *Step 11:* – Discard the solution in the Erlenmeyer in the corresponding waste container.

Part B. Exact titration.

☐ *Step 12:* – Refill the buret with $Na_2S_2O_3$ 0.0025 M. At this point you know how much volume you will need from the trial titration, you don not need to fill the buret up to zero but add a few milliliters more than the expected volume.

☐ *Step 13:* – In a 125 mL Erlenmeyer, obtain about one inch of a spatula of KI.

☐ *Step 14:* – Pipet 10.0 mL of the diluted bleach into the KI containing Erlenmeyer. Add 20 mL of distilled water from a graduated column and 20 drops of HCl 2 M. Homogenize the mixture by stirring the flask.

☐ *Step 15:* – Record the buret initial reading to the closest hundredth of mL.

☐ *Step 16:* – Place the Erlenmeyer under the buret and add 5 mL less than the volume used in the trial titration.

☐ *Step 17:* – Add 40 drops of a 0.2% starch solution to the solution in the Erlenmeyer.

☐ *Step 18:* – Continue the titration adding the titrant drop by drop until the solution turns transparent.

☐ *Step 19:* – Record the buret final reading in your notes.

☐ *Step 20:* – Discard the solution in the Erlenmeyer. If the time allows, do a second exact titration.

☐ *Step 21:* – When finished empty the buret and rinse it with plenty of distilled water. The buret should be stored empty, open and upside down.

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Bleach Strength

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1. Find the definition for the following concepts:
   (a) Oxidation-Reduction reaction.

   (b) Reducing agent.

   (c) Oxidizing agent.

   (d) Limiting reactant.

   (e) Cadaveric particles.

2. What is the proportion of bleach to water necessary to make water safe to drink?

3. What is the active ingredient in liquid bleach?

4. What is the active ingredient in bleach powder?

5. Indicate the oxidation state of all elements in the following species:
   × H^+
   × H_2O
   × Cl^-
   × ClO^-
   × I^-
   × I_2
   × $S_2O_3{}^{2-}$
   × $S_4O_6{}^{2-}$

STUDENT INFO

Name: ____________________ Date: __________

Results
EXPERIMENT

Bleach Strength

Bleach dilution

Final buret reading __________ mL

Intial buret reading __________ mL

NaClO added __________ mL

Trial titration

Final buret reading __________ mL

Intial buret reading __________ mL

$Na_2S_2O_3$ added __________ mL

Exact titrations

Titration	1	2
Final buret reading (mL)	________	________
Initial buret reading (mL)	________	________
$Na_2S_2O_3$ added (mL)	________	________
Moles $Na_2S_2O_3$	________	________
Moles NaOCl	________	________
[NaOCl] (diluted) (M)	________	________
[NaOCl] (original) (M)	________	________
Mean Molarity	________	

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Bleach Strength

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1. Commercial bleach claims a 5.25% NaClO by weight. Assuming a density of 1.0 g/mL for NaClO and water and considering your results in the laboratory, is the claim true?

2. Considering the bleach:water ratio from question 2 in the Prelab Question. What is the concentration, in molarity, of NaClO in the safe to drink dilution?

EXPERIMENT

Acids Strength

Goal

The goal of this laboratory is to learn three different techniques to measure pH.

Background

Acids are considered dangerous substances and while it is true that special care must be taken when handling some of these products, many acids are used in daily life with great simplicity. For example, aspirin's active principle ingredient is acetylsalicylic acid, vinegar is a 5% solution of acetic acid, and Vitamin-C is ascorbic acid.

Acids are defined by Brønsted-Lowry as proton (H^+) donors, because when dissolved in water, acids dissociate by releasing protons. Strong acids dissociate completely into two ions: the anion is the acid conjugate base and the the cation is a proton. For stoichiometric reasons, the concentration of each ion in the solution equals the initial concentration of the acid. Weak acids do not dissociate completely but, instead, only do so to a certain extent. The strength of weak acids depends on how much they dissociate, a ratio characterized by the acid dissociation constant, K_a.

For a certain weak acid

$$\mathrm{HA} \rightleftharpoons \mathrm{H^+} + \mathrm{A^-}$$

the equilibrium can be characterized by the expression:

$$K_a = \frac{[H^+][A^-]}{[HA]}$$

where K_a (capital letter) is the dissociation constant and the brackets represent the concentrations of the respective species.

pH

The acidity depends on the concentration of H^+ in the solution, which will always be a value between 10^{-14} and 1. To avoid using such a large and complex range, a logarithmic scale is preferred, ranging from 0 to 14.

This scale is called the pH, defined by Danish biochemist Sørensen in 1909 as the "power of Hydrogen".

$$pH = -log[H^+]$$

Considering that the strength of acids is their power to generate H^+, and in order to compare the strength of different acids, pH should be measured.

In this experiment, pH for different solutions will be measured using three different approaches; indicators, pH-paper, and a pH-meter. Each of these techniques has a different accuracy. While using **pH indicators** it is only possible to determine a range for the pH, **pH paper** will provide an estimated unit, and the **pH-meter** will read down to the decimals.

The experiment

The experiment today is divided in 3 parts:

Part A. Determining pH-range using indicators.

Part B. Determining pH using pH-paper.

Part C. Determining pH using a pH-meter.

Part A. Determining pH-range using indicators.

A pH indicator is a chemical substance that, when added to a solution, changes its color depending on the solution's pH. The pH at which color changes is called the **equivalence point**. Some indicators have one equivalence point while others have more. By using an indicator it will be possible to determine a pH-range only. Combining different indicators the range can be narrowed down.

Example

A solution is divided into three samples and each sample is tested using a different indicator, each one with a different equivalence point.
- Indicator A is colorless for pH values lower than 10 and turns pink for a pH greater than 10.
- Indicator B is blue up to pH 3, then it turns colorless.
- Indicator C is colorless for a pH lower than 6, when it turns red. The red color changes to yellow between pH 9 and pH 10, and stays yellow for higher pH values.
Estimate the pH range for the solution if the three samples are colorless.

Answer: Indicator A tells that the solution's pH is lower than 10. Therefore, the range must be $0 < pH < 10$.
Indicator B shows that the solution's pH is greater than 3 (3 to 14). This information, combined with that of indicator A, narrows the pH range to $3 < pH < 10$.
Indicator C tells that the solution's pH is lower than 5. The final range is $3 < pH < 5$.

Parts B. Determining pH using pH-paper.

pH paper, or **universal pH paper** is a strip of paper impregnated with a mixture of indicators. The paper changes color smoothly for each pH value and the pH is determined by matching the color of the paper with that in a chart attached to the pH paper. In the best case, by using pH paper it will be possible to select one value for the pH, although in some situations, when the color is not clearly one in the chart, it will be possible to indicate a range of two numbers only.

Parts C. Determining pH using a pH-meter.

pH meters are scientific instruments that measure the pH of solutions. The tool has a probe that is inserted in the liquid and measures the concentration of protons based on the voltage between two electrodes. Their use is similar to that for a conductivity-meter, and some instruments can perform both tasks. Probes need to be cleaned after each measurement to avoid contamination and should never be touched with the hands or paper towels. A calibration of the tool with solutions of known pH ensures the best accuracy. After use, probes are best kept moist, typically in distilled water. A pH meter can offer a precision down to the hundredth of the unit of pH.

Procedure

Part A. Determining pH-range using indicators.

☐ *Step 1:* – Obtain 5 large test tubes and a rack. Clean the test tubes and let them drain upside down in the rack.

☐ *Step 2:* – Obtain a 100 mL beaker with plenty distilled water and transfer 4.0 mL in a graduated cylinder.

☐ *Step 3:* – Label one test tube with the letter R, and pour the 4.0 mL of distilled water into it. This test tube will be used as a volume reference for the following test tubes.

☐ *Step 4:* – In a 50 mL beaker get 40 mL of one of the following solutions.

1. 0.1 M H_3PO_4
2. 0.1 M $HC_2H_3O_2$
3. 0.1 M NaH_2PO_4
4. 0.1 M $Zn(NO_3)_2$
5. 0.1 M NH_4NO_3
6. Distilled water
7. Tap water
8. Unknown solution

⚠CAUTION!

⚠ Handle concentrated acids with care to avoid chemical burns.

⚠ Are you wearing your goggles? Do you know where the eye-washer is?

!

☐ *Step 5:* – Obtain a 250 mL beaker for waste, a medicine dropper of plastic pipette

☐ *Step 6:* – Label the rest of the test tubes from 1 to 4. Use test tube R as a reference for the volume to add in the following test tubes. Rinse the dropper with the distilled water (discarding it in the waste beaker) between different solutions. Add 4.0 mL of the solution chosen to each test tube.

☐ *Step 7:* – Add two drops of Thymol Blue to test tube 1.

☐ *Step 8:* – Add two drops of Methyl Orange to test tube 2.

☐ *Step 9:* – Add two drops of Methyl Red to test tube 3.

☐ *Step 10:* – Add two drops of Bromothymol Blue to test tube 4.

☐ *Step 11:* – Record the colors in the results page.

☐ *Step 12:* – Reserve the solution in the 50 mL beaker for Parts B and C. Set the beaker on a paper towel and label it.

☐ *Step 13:* – Discard the solutions and clean the test tubes.

☐ *Step 14:* – Repeat steps 4 to 12 for another solution. Make sure to rinse the dropper.

Part B. Determining pH using pH-paper.

☐ *Step 15:* – Obtain a stirring rod and the universal indicator paper. Be careful not to contaminate the paper with dirty gloves or bare hands. Always set the (small) pieces of paper on a clean paper towel.

☐ *Step 16:* – Rinse the stirring rod with the wash bottle over the waste beaker. Rinse thoroughly between different solutions.

☐ *Step 17:* – Immerse the tip of the stirring rod in one solution and touch a piece of pH-paper with the rod.

☐ *Step 18:* – Match the color of the drop on the paper with the color chart on the plastic case. Record the pH value in the results section.

☐ *Step 19:* – Repeat the same process; rinsing, dipping and touching, for the rest of the solutions.

☐ *Step 20:* – Reserve the solutions in the beakers for part C.

Part C. Determining pH using a pH-meter.

☐ *Step 21:* – Install the pH-meter. The probe of the instrument is immersed in a clean solution that must be kept aside until the end of the experiment.

☐ *Step 22:* – Cleaning and measuring. It is very important to follow the cleaning steps before each measurement. You will need a 100 mL beaker labeled waste a wash bottle and a 100 mL beaker with distilled water.

1. Rinse the probe using the wash bottle over the waste beaker.

2. Dip the probe 3-4 times into the 250 mL beaker with clean distilled water.
3. Dip the probe into the solution to be measured. Use a 50 mL beaker with at least 25 mL of solution.
4. Repeat cleaning steps 1 and 2.

☐ *Step 23:* – Following the procedure outlined in step above, measure the eight solutions.

☐ *Step 24:* – Record all the values.

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Acids Strength

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1. Find the definition of the following concepts:
    (a) Acids (according to Brønsted-Lowry definition).
    (b) pH.
    (c) Acid dissociation constant.
    (d) Equivalence point.
2. Name the three methods to determine pH to be used in this laboratory.
3. Order the three methods from the most accurate to the less accurate, according to your expectations.
4. Read the example box. What would be the pH range for the same example if sample C had turned reddish? Briefly reason your answer.

STUDENT INFO

Name: ____________ Date: ________

Results
EXPERIMENT

Acids Strength

Part A. Determining pH-range using indicators.

Indicate the colors of the solutions.

Solution	Thymol Blue	Methyl Orange	Methyl Red	Bromothymol Blue
H_3PO_4				
$HC_2H_3O_3$				
NaH_2PO_4				
$Zn(NO_3)_2$				
NaH_4NO_3				
Distilled water				
Tap water				
Unknown				

Part B. Determining pH using pH-paper.

Solution	pH
H_3PO_4	________
$HC_2H_3O_3$	________
NaH_2PO_4	________
$Zn(NO_3)_2$	________
NaH_4NO_3	________
Distilled water	________
Tap water	________
Unknown	________

Part C. Determining pH using a pH-meter.

Solution	pH
H_3PO_4	________
$HC_2H_3O_3$	________
NaH_2PO_4	________
$Zn(NO_3)_2$	________
NaH_4NO_3	________
istilled water	________
Tap water	________
Unknown	________

STUDENT INFO

Name: ______________________ Date: __________

Post-lab Questions

Acids Strength

---

1. Use the following indicator's colors chart to determine the pH-range for the solutions.

Solution	pH-range	Solution	pH-range
1. H_3PO_4	__________	5. NaH_4NO_3	__________
2. $HC_2H_3O_3$	__________	6. Distilled water	__________
3. NaH_2PO_4	__________	7. Tap water	__________
4. $Zn(NO_3)_2$	__________	8. Unknown	__________

2. Which of the three techniques is the most accurate?

3. Calculate the concentration of H_3O^+ ions in the most acidic solution, using the pH-meter value.

4. Calculate the concentration of H_3O^+ ions in the less acidic solution, using the pH-meter value.

EXPERIMENT

Entropy

Goal

The goal of this laboratory experiment is to calculate the minimum entropy change for a chemical reaction involving the dissolution of sodium nitrate.

Materials

- ☐ A coffee-cup calorimeter
- ☐ A thermometer and a lit for the calorimeter
- ☐ $NaNO_3(s)$

Background

Learn more about heat and reactions in this Video Link

Visual

Chemical kinetics is a discipline that deals with the pace of chemical reactions by using activation energies, concentrations, and temperatures. It also deals with the different pathways of reaction. However, chemical kinetics does no provide information about the reasons why chemical reactions happen. Thermodynamics informs about the direction in which a chemical reaction occurs spontaneously and the reasons for this reaction to happen, without informing about the speed at which the process occurs. It uses thermodynamic functions–enthalpies, entropies, and Gibbs free energies–to understand the factors that favor the spontaneity of a reaction. Overall, kinetics and thermodynamics are complementary disciplines in chemistry as it is convenient to gain insight into the pace and spontaneity of chemical reactions.

Spontaneity

A spontaneous process happens naturally, without any help or external input. For example, ice spontaneously melts at room temperature, without any help. Still, spontaneity happens under a certain set of conditions. For example, ice melts at room temperature but does not melt at -5°C, as the process in these conditions is not spontaneous. Other spontaneous processes are the rusting of iron at room temperature, or the freezing of water below 0°C. Spontaneous processes have a natural tendency to occur. Still, this tendency might not be actualized in practice. At the same time, this natural tendency to occur is not linked in any way to the speed or rate of the process. For example, diamonds have a natural tendency to become graphite. Still, the rate of this change is so slow that in human years it does never occur. Remember that thermodynamics addresses spontaneity whereas kinetics address the rate of occurrence of a process.

The meaning of entropy

Entropy is represented with the letter S. Qualitatively, it is a measure of the spreading in space of the energy of a system, or simply the system. For example, in a glass full of hot water, the energy is concentrated in the glass as the water molecules initially have high thermal energy. With time, the system evolves naturally so that heat spreads between the room and the container. The initial state (glass with hot water) has lower entropy than the final state (glass with thermalized water). Another example, think about a concentrated solution separated from a more diluted system by means of a membrane. Initially, the entropy of the system is low, as one side of the container has a high density of molecules whereas the other side has a lower density. As the molecules diffuse and the concentration equalizes. In this case, the entropy of the initial state (a concentrated and a diluted solution) is lower than the entropy of the final system (two solutions of equal concentration). These two examples represent how energy and matter spread as entropy increases in spontaneous processes. On one hand, entropy, just like enthalpy, is a *state function* in thermodynamics which means that the change of entropy of a system when going from an initial into a final state is independent of the path taken to arrive from the initial into the final state. On the other hand, entropy is an *extensive property* that depends on the size of the system and for example the entropy change when heating 2 moles of gas is larger than when heating 1 mole of gas.

Standard molar entropy

The standard molar entropy of a substance is the absolute entropy of one more substance at 1 atm. Entropies are absolute properties in contrast to other relative thermodynamic functions like the enthalpy in which a set of molecular states act as a reference. The tabulated standard entropy values are listed at the end of the chapter at one bar and 25 °C, the conventional temperature of reporting thermodynamic data. The units of molar entropy are J/molK. Molar entropies tend to be small values–as the unit joule is a small unit of energy. Finally, entropies can be positive or negative values.

Factors affecting entropy

We can qualitatively rationalize the trends between the standard molar entropy values for different substances and for different conditions (temperature, volume, pressure) by using some simple rules described below:
⚘ **The state of matter:** The standard entropy of gases is larger than the standard entropy of liquids as gases present large degree of freedom (possible configurations) than liquids. The standard entropy of liquids is larger than the standard entropy of solids as liquids present large degree of freedom than solids. The molecules of a liquid molecules can jiggle more in comparison with the molecules of a solid. Overall, we can assume that liquids and solid has almost null entropy whereas gases have very large entropy. For example, the entropies of ice, water and steam are respectively 41, 69.95, and 69.95 J/molK.
⚘ **Molar mass:** For monoatomic substances (e.g. Ne, Ar, etc.) the larger the atomic weight the larger entropy. This is because the energy levels of larger molecules are more packed, and hence are easier to access at a given temperature.
⚘ **Molecular complexity:** For substances with comparable molar mass (O_3 and F_2), the more complex the molecule the larger entropy, as the number of degrees of freedom are directly correlated with the number of microstates.
⚘ **Temperature:** Temperature increases entropy as the system have more accessible microstates. The plot below represents the entropy change with temperature for a substance. ⚘ **Number of particles:** The larger the number of particles of a system, the larger entropy, as the more particles the more microstates or possible configurations.
⚘ **Volume:** The larger the volume of a system, the larger entropy, as the larger volume the more microstates or possible configurations.

Calculating entropy changes in reactions

We can calculate the standard entropy of a reaction in a similar way as we calculate the standard enthalpy of a reaction:

$$\boxed{\Delta S^\circ_R = \Delta S^\circ_{products} - \Delta S^\circ_{reactants}} \quad \text{Entropy change} \tag{1}$$

where:

ΔS°_R is the standard entropy change of the reaction

$\Delta S^\circ_{products}$ is the standard entropy of all products

$\Delta S^\circ_{reactants}$ is the standard entropy of all reactants

It is important to take into account the stoichiometric coefficients. For example, for the reaction:

$$2\,CH_3OH_{(g)} \quad + \quad 3\,O_{2\,(g)} \quad \longrightarrow \quad 2\,CO_{2\,(g)} \quad + \quad 4\,H_2O_{(l)}$$

We have the entropy values of: $S^\circ(CH_3OH_{(g)})$=239.7 J/K· mol, $S^\circ(O_{2\,(g)})$=161.1 J/K· mol, $S^\circ(CO_{2\,(g)})$=213.79 J/K· mol, and $S^\circ(H_2O_{(l)})$=69.95 J/K· mol. We can calculate ΔS°_R:

$$\begin{aligned} \Delta S^\circ_R = \Delta S^\circ_{products} - \Delta S^\circ_{reactants} = \\ \Big(2 \cdot S^\circ(CO_{2\,(g)}) + 4 \cdot S^\circ(H_2O_{(l)})\Big) - \Big(2 \cdot S^\circ(CH_3OH_{(g)}) + 3 \cdot S^\circ(O_{2\,(g)})\Big) \\ = \Big(2 \cdot 213.79 + 4 \cdot 69.95\Big) - \Big(2 \cdot 239.7 + 3 \cdot 161.1\Big) = -255.32 J/K \end{aligned}$$

Stimating the sign of the entropy changes in reactions

Often times were are more interested in predicting the sign of the entropy change of a reaction than to compute the exact value. This is because the sign can be used in order to estimate wether a reaction proceeds spontaneously. Two basic rules are used in order to estimate the entry change sign:

- **The state of matter:** solids and liquids have very low entropy in comparison to gases. For example, in the case below, we have the the production of liquid water from ice produces entropy as liquids have more entropy than solids:

$$H_2O_{(s)} \longrightarrow H_2O_{(l)} \qquad \Delta S^\circ_R > 0$$

On the other hand, the condensation of water vapor to produce a liquid consumes entropy, as liquids have less entropy than gases:

$$H_2O_{(g)} \longrightarrow H_2O_{(l)} \qquad \Delta S^\circ_R < 0$$

- **The number of molecules:** the larger the number of molecules of gas the larger entropy. For example, in the reaction below we have that we produce three molecules from two molecules. Hence, the entropy increases.

$$2\,SO_{3\,(g)} \longrightarrow 2\,SO_{2\,(g)} + O_{2\,(g)} \qquad \Delta S^\circ_R < 0$$

However, this rules only works if we only take into account the number of gas molecules. For example, in the reaction below we produce two gas molecules from three gas molecules and hence we lose entropy:

$$2\,H_2S_{(g)} + SO_{2\,(g)} \longrightarrow 3\,S_{(s)} + 2\,H_2O_{(g)} \qquad \Delta S^\circ_R < 0$$

Gibbs free energy

The Gibbs free-energy is just a combination of enthalpy and entropy for a given temperature:

$$G = H - T \cdot S \tag{2}$$

Working at constant temperature, we can compute the change in Gibbs free energy as:

$$\boxed{\Delta G^T = \Delta H - T\Delta S \quad \text{Gibbs free-energy}} \tag{3}$$

where:

ΔG^T is the Gibbs free-energy change at constant temperature

ΔH is the enthalpy change

ΔS is the entropy change

T is the temperature

Gibbs free-energy is a state function that only depends on the final and initial state and not the path followed. At the same time, Gibbs free-energy depends on temperature and pressure–we will discuss more about this at the end of the section. More importantly, the change in Gibbs free-energy is associated with the spontaneity of the process. Gibbs free-energy decreases with temperature and the decrease is sharper for gases in comparison to liquids, and at the same time, the decrease is also sharper for liquids in comparison to solids.

Gibbs free energy and spontaneity

The Gibbs free-energy change of a reaction is associated with the spontaneity of the process or with its state of equilibrium. In particular, reactions that produce Gibbs free-energy are nonspontaneous. Differently, reaction consuming free-energy are indeed spontaneous. Finally, reaction without a change in free-energy are in equilibrium.

- $\Delta G < 0$**:** The reaction is spontaneous
- $\Delta G = 0$**:** The reaction is in equilibrium
- $\Delta G > 0$**:** The reaction is nonspontaneous

This is because the expression of Gibbs free energy is related to the total entropy change of the universe given by the second law of thermodynamics:

$$\Delta G^T = \Delta H - T\Delta S \quad \text{and } T\Delta S^{T,P}_{univ} = -(\Delta H - T\Delta S) \tag{4}$$

Hence, we have that

$$\Delta G^T = T\Delta S^{T,P}_{univ} \tag{5}$$

This last expression means that an increase in total entropy is accompanied by a decrease in Gibbs free energy at constant pressure and temperature. At the same time the expression indicates that the direction of spontaneous change corresponds to a decrease of Gibbs free energy, at constant temperature and pressure.

Procedure

Measuring ΔH for the $NaNO_3$ dissolution

☐ *Step 1:* – Obtain a coffee-cup calorimeter with a thermometer and a plastic lid.

☐ *Step 2:* – Weight the $NaNO_3$ mass needed to prepare 100mL of a 1M solution of this salt. The calculations were covered in the prelab.

☐ *Step 3:* – Use a 100mL graduated cylinder to add 100mL of distilled water in the calorimeter.

☐ *Step 4:* – Measure the temperature of water in the calorimeter. Record the mean of the two temperatures as your initial temperature t_i.

☐ *Step 5:* – Get the timer ready. Time should be taken as soon as the solid reactant is added.

☐ *Step 6:* – Add the $NaNO_3$ to the calorimeter. If you are using a magnetic stirrer set the speed to a medium speed. Then add the base. Start the timer immediately and place the lid on the calorimeter.

☐ *Step 7:* – Record the temperature for 4 min. Do not stop the timer during the experiment. Record the results in the results table.

☐ *Step 8:* – Plot the temperature of the calorimeter against time and use a straight line to extrapolate your results against the starting time (time = 0 s). Record the extrapolated temperature as t_f.

☐ *Step 9:* – Calculate the change of enthalpy, ΔH.

☐ *Step 10:* – Repeat the procedure twice in order to calculate the average enthalpy change.

Good Lab Practice

✍ Make sure you always use the same thermometer all through the experiment so that the results are consistent.

Calculations

① This is the initial temperature before you add the solid.

② This is the final temperature after the dissolution reaction stops. You need to plot the data and extrapolate to get this value.

③ This is the maximum temperature during the experiment

④ This is the number of moles of salt: $c_{acid} * 50/1000$

⑤ You can calculate the enthalpy change (ΔH) by means of the following formula:

$$\Delta H = -\frac{1}{\textcircled{3}} \cdot \left(4.184 \cdot 1.0 \cdot 100.0 \cdot (\textcircled{2} - \textcircled{1}) + 10 \cdot (\textcircled{2} - \textcircled{1})\right)$$

⑥ Calculate the minimum entropy change for equilibrium using the maximum temperature and the average enthalpy (remember enthalpy is expressed in KJ/mol and entropy in J/molK)

$$\Delta S_{min} = \frac{\Delta H}{T_m} = -\frac{\textcircled{5} \times 1000}{\textcircled{3}}$$

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Entropy

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1. Ammonium nitrate (NH_4NO_3) is dissolved in a coffee-cup calorimeter. We recorded the temperatures with time and the data is given below. The initial temperature, t_i was 25.0 °C.

t (s)	30	60	90	120	150	180	210	240
T (°C)	40.2	40.1	40.3	40.4	40.5	40.6	40.7	40.8

   (a) Plot T vs. t. Obtain t_f, the final temperature after the reaction starts, by extrapolating to the time of mixing (time = 0 s) with a straight line.

t_f = ______________________

   (b) Indicate whether this is an exothermic or an endothermic reaction.

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(c) Explain the rise and fall of temperature inside the calorimeter.

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2. Predict the sign of the entropy change for the dissolution of NH_4NO_3.

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3. Predict the spontaneity of the dissolution of NH_4NO_3 using the signs of the entropy and enthalpy estimabed above.

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STUDENT INFO

Name:

Date:

Results

EXPERIMENT

Entropy

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Trial 1															
t (s)	0	30	60	90	120	150	180	210	240		T_i (°C) ①	T_f (°C) ②	T_m (°C) ③	n (moles) ④	$\Delta H (kJ/mol)$ ⑤
T (°C)															
Trial 2															
t (s)	0	30	60	90	120	150	180	210	240		T_i (°C)	T_f (°C)	T_m (°C)	n (moles)	$\Delta H (kJ/mol)$
T (°C)															
											Average ΔH (kJ/mol) =				
											Average T_m (°) =				
											ΔS_{min} $(J/molK)$ = ⑥				

STUDENT INFO

Name: ____________________ Date: __________

Post-lab Questions

Entropy

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1. Predict the sign of the entropy change for the dissolution of $NaNO_3$

2. Use the data below to compute the entropy change in the dissolution process of sodium nitrate at 298K. Compare your experimental result with the calculated giving reasons for the difference.

	$NaNO_{3(s)}$	$\longrightarrow$	$Na^+_{(aq)}$	+	$NO_3{}^-{}_{(aq)}$
$S°(J/molK)$	116.32		59.0		146.4
n					
$S° \cdot n$					
$\Delta S°(J/molK)$					

3. Use the data below to compute the enthalpy change in the dissolution process of sodium nitrate at 298K. Compare your experimental result with the calculated giving reasons for the difference.

	$NaNO_{3(s)}$	$\longrightarrow$	$Na^+_{(aq)}$	+	$NO_3{}^-{}_{(aq)}$
$H°(KJ/mol)$	-466.68		-240.1		-207.36
n					
$H° \cdot n$					
$\Delta H°(KJ/mol)$					

4. Was the dissolution of sodium nitrate spontaneous in the conditions of the experiment?

5. Calculate the value of Gibbs free energy at 298K using the entropy and enthalpy calculated above.

EXPERIMENT

Solubility product of an iodate salt

Goal

The goal of this laboratory experiment is to calculate the solubility poduct of a iodate salt.

Materials

- ☐ A 10mL pipet, a 150mL beaker, stiring rod
- ☐ Filter paper, finel, stand, ring
- ☐ A burete, stand and fisher clamp
- ☐ A 0.001M $KIO_{3\,(aq)}$, $KI_{(s)}$, 2M $HCl_{(aq)}$, 2/
- ☐ A series of 125mL (or 250mL) erlenmeyers

Background

Learn more about heat and reactions in this 🌍 Video Link

Visual

When insoluble compounds dissolve in water an equilibrium between the solid and the dissolved ions establishes. Based on the stoichiometry of the compound we will obtain a different number of moles of ions in solution.

Solubility equilibrium

Insoluble compounds dissolve in small amounts in water. Solubility equilibrium is the equilibrium that describes the dissolution of an insoluble compound to produced an aqueous solution of ions. For example, for the case of silver chloride, we have:

$$AgCl_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$$

The solubility equilibrium above represents the dissolution of the insoluble salt–a solid–to produce ions, silver and chloride in solution. In general terms, solubility equilibriums start with a solid and generate ions in solution. In order to break down the insoluble compound into ions, we just need to take into account the stoichiometry of the compound. For example, for the case of Cobalt(II) phosphate, we have:

$$Co_3(PO_4)_{2\,(s)} \rightleftharpoons 3\,Co^{2+}_{(aq)} + 2\,PO_4{}^{3-}{}_{(aq)}$$

Solubility product in terms of molarities, K_{sp}

Silver chloride, AgCl is an insoluble compound in water which means this compound will not fully dissolve in water. Still, small quantities of the salt will indeed dissolve leading to a small amount of silver and chlorine ions. As this compound is insoluble the ion concentration will be very small. The dissociation equilibrium involved in the dissolution, indicate below, is characterized by an equilibrium constant K_{sp} called *solubility product constant* or simply *solubility product*:

$$AgCl_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \qquad K_{sp} = [Ag^{+}] \cdot [Cl^{-}]$$

As pure solids are not included in any equilibrium constant, the formula above does not include $AgCl_{(s)}$. Hence, all solubility products simply result from the product of the molarities of the ions involved in the dissociation with the corresponding powers based on the stoichiometry of the salt. This is why the term *product* is included on its name. K_{sp} will have a different explicit expression depending on the formula of the insoluble compound–deppending on its stoichiometry. For example, K_{sp} for calcium fluoride–a 1:2 compund–would be:

$CaF_{2\,(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2\,F^{-}_{(aq)}$ $\qquad K_{sp} = [Ca^{2+}] \cdot [F^{-}]^2$

Solubility products are related to the solubility of chemicals. However, the relation is not one-on-one. In other words, a larger K_{sp} does not necessarely imply a larger solubility. For example, K_{sp} for $PbBr_2$ is 6.6×10^{-3}M whereas K_{sp} for $MgCO_3$ is 4.0×10^{-5}M. We have that K_{sp} for $PbBr_2$ is smaller than K_{sp} for $MgCO_3$. However, the solubility of $PbBr_2$ is indeed larger than the solubility of $MgCO_3$. Table **??** reports numerous solubility product constant values.

Relating solubility to the concentration of ions

We have that insoluble compounds dissolve to produce small amounts of ions in solutions. The concentration of these ions is directly related to solubility s by means of the stoichiometry of the compound. For example, for silver chloride–a salt with a 1:1 stoichiometry–we have that solubility is has a 1 and 1 relationship with the concentration of ions:

$$[Ag^+] = s \text{ and } [Cl^-] = s$$

Similarly, for barium sulfate we have that solubility has a 1:1 relationship with the concentration of ions:

$$[Ba^{+2}] = s \text{ and } [SO_4{}^{-2}] = s$$

When the stoichiometry of the compound is not 1:1 we need to include the stoichiometry coefficients in the relationship between solubility and ion concentration. For example, for Ag_2SO_4–a compound with 2:1 stoichiometry, we have that

$$[Ag^+] = 2 \cdot s \text{ and } [SO_4{}^{-2}] = s$$

This is because for every mole of silver sulfate we produce two moles of silver–and hence the factor two–and one mole of sulfate in solution. Similarly, for $Nd_2(CO_3)_3$ we have

$$[Nd^{+3}] = 2 \cdot s \text{ and } [CO_3{}^{-2}] = 3 \cdot s$$

Relating solubility to the concentration of ions

The solubility product is directly related to molar solubility s. We will demonstrate how to obtain this relationship by means of three examples. First, the solubility product of AgCl is:

$AgCl_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Cl^{-}_{(aq)}$ $\qquad K_{sp} = [Ag^+] \cdot [Cl^-]$

As the concentration of each ion, Ag^+ and Cl^-, is related to the molar solubility of the salt, we have

$$K_{sp} \quad = \quad [Ag^+] \ \cdot \ [Cl^-] \quad = \quad (s) \ \cdot \ (s) \quad = \quad s^2$$

Second, for silver sulfide, an insoluble compound with a more complex stoichiometry, we have:

$Ag_2S_{(s)} \rightleftharpoons 2\,Ag^{+}_{(aq)} + S^{2-}_{(aq)}$ $\qquad K_{sp} = [Ag^+]^2 \cdot [S^{2-}] = (2s)^2 \cdot (s) = 4s^3$

Third, the solubility equilibrium and K_{sp} expression for $Nd_2(CO_3)_3$ is

$Nd_2(CO_3)_{3\,(s)} \rightleftharpoons 2\,Nd^{3+}_{(aq)} + 3\,CO_3{}^{2-}{}_{(aq)}$ $\qquad K_{sp} = [Nd^{3+}]^2 \cdot [CO_3{}^{2-}]^3$

As the ion concentrations are related to solubility, we have

$$K_{sp} \quad = \quad (2s)^2 \ \cdot \ (3s)^3 \quad = \quad 108s^5$$

K_{sp} in terms of molar solubility: general formula

For any insoluble salt A_xB_y, we have that K_{sp} is related to s by means of a general formula:

$$\boxed{K_{sp} = a \cdot s^b} \tag{6}$$

where:

a is $x^x \cdot y^y$

b is $x + y$

For example, for Ba_1F_2 the constant a would be $1^1 \cdot 2^2$, that is four, whereas the constant b will be $1 + 2$ that equals to three. As such, the expression of K_{sp} in terms of s would be: $K_{sp} = 4s^3$. This approach is useful when the we need to compute the solubility product constant given the molar solubility.

Molar solubility in terms of K_{sp}: general formula

We previously explored the relationship between K_{sp} and molar solubility. Here we will explore the relationship between molar solubility and K_{sp}, simply solving for s in Equation 6. Again, for any insoluble salt A_xB_y, we have:

$$\boxed{s=\left(\frac{K_{sp}}{a}\right)^{\frac{1}{b}}} \tag{7}$$

where:

a is $x^x \cdot y^y$

b is $x+y$

For example, for Ba_1F_2 the constant a would be $1^1 \cdot 2^2$, that is four, whereas the constant b will be $1+2$ that equals to three. As such, the expression of s in terms of K_{sp} would be: $s=\left(\frac{K_{sp}}{4}\right)^{\frac{1}{3}}$. This approach is useful when the solubility product constant is given and we need to calculate the molar solubility. As the relation between molar solubility and K_{sp} is not a one-to-one relationship, K_{sp} are not directly related to solubility. For salts with a similar stoichiometry, it would be safe to compare solubilities in terms of K_{sp}. For example:

CuS	K_{sp}=8 × 10^{-37}	$s = 6 \times 10^{-19}$
PbS	K_{sp}=3 × 10^{-28}	$s = 1 \times 10^{-14}$
LiF	K_{sp}=2 × 10^{-3}	$s = 3 \times 10^{-2}$

and we have that the lower K_{sp} the lower solubility. When the salt stoichiometry differs

MgF_2	K_{sp}=5 × 10^{-11}	$s = 2 \times 10^{-4}$
Li_3PO_4	K_{sp}=2 × 10^{-4}	$s = 7 \times 10^{-2}$
Li_2CO_3	K_{sp}=8 × 10^{-4}	$s = 6 \times 10^{-2}$

then an increase in K_{sp} do not necessarily follows an increase in solubility.

Procedure

Preparation of a saturated solution

☐ *Step 1:* – In a 150mL beaker mix 50mL of a 0.2M KIO_3 solution and 20mL of a 1M $Ca(NO_3)_2$ solution. Make sure you pick the reagents with the indicated concentration as in the lab there could be more diluted potasium iodate also. Stir the resulting mix with a stirring rod. A precipitate will appear.

☐ *Step 2:* – Using gravity filtration, filter the precipitate. Make sure you rinse the beaker with water, to collect all precipitate leftovers.

☐ *Step 3:* – Wash the precipitate on the filter paper with three different volumes of water.

☐ *Step 4:* – Use a spatula to separate the precipitate into two halves. Dispose of each part in a clean labeled 100mL beaker. Put one of the beakers aside to be used in case of an accident. Use the other beaker in the rest of the experiment. If you do not use the first half of the precipitate, dispose it at the end of the experiment following your instructor's indications.

☐ *Step 5:* – Add 40mL of distilled water to the beaker containing the precipitate. Use a stirring rod to stir the mixtures.

☐ *Step 6:* – This is the saturated $Ca(IO_3)_2$ solution. Let this solution rest for 30 minutes, occasionally stirring.

☐ *Step 7:* – It is important not to add additional water to this solution.

Preparing three mixtures of the saturated solution

☐ *Step 1:* – The goal now is to prepare three samples of an acidified saturated solution containing iodide. To do this we will first have to filter the saturated solution to eliminate the solid. After that, we will add acid and iodine to the exact volume of this solution.

☐ *Step 2:* – Filter the saturated solution obtained earlier in the experiment. Make sure the funnel and the paper are dry. Do not wash the precipitate with water. The filtrate is the solution to be titrated.

☐ *Step 3:* – Rinse a 10mL pipet with distilled water making sure the pipet is free of water at the end. Now rinse the same pipet two times with the filtrate collected in the previous step, discarding the solution from the pipet.

☐ *Step 4:* – Pipet 10mL from the filtrate into a 125mL (or 250mL) Erlenmeyer flask. First, add 20mL of water distilled water from a cylinder. Add also 20 drops of 2M HCl and the equivalent of 1mL of solid KI, measured with a 10mL cylinder. Mind you can approximately measure the volume of a solid with a cylinder. Finally, add 40 drops of a 0.2/

☐ *Step 5:* – Repeat the previous step two more times in order to prepare a total of three mixtures.

Titration of the saturated solution mixtures: trial titration

☐ *Step 1:* – The final goal here is to calculate the molarity of the saturated solution, which corresponds to the solubility of the insoluble compound. Here we will estimate the equivalence point of the titration of sodium thiosulfate and sodium iodate.

☐ *Step 2:* – Rinse a buret with a small portion of $Na_2S_2O_3$ solution, discarding the solution used in the rinses.

☐ *Step 3:* – Fill a buret with $Na_2S_2O_3$ solution.

☐ *Step 4:* – Titrate one of the samples adding increments of 1mL of sodium thiosulfate at a time. Swirl the mixture after each addition. You will have reached the equivalency point when the solution becomes colorless. Try placing the flask containing the analyte on a piece of paper. The white background will help see the color change when it occurs.

☐ *Step 5:* – Write down the volume of thiosulfate added to reach the equivalency point to the nearest 0.01mL. This will give an estimate of the volume needed to reach the endpoint.

Titration of the saturated solution mixtures: exact titration

☐ *Step 1:* – After having estimated the volume of titrate needed to reach the endpoint, we will now speed up the titration by rapidly titrating the second thiosulfate sample.

☐ *Step 2:* – Add thiosulfate to a second sample of saturated solution, stopping 3mL before the estimated endpoint.

☐ *Step 3:* – Add $Na_2S_2O_3$ drop by drop until you reach the endpoint and the solution becomes colorless. Record the buret reading to the nearest 0.01mL.

☐ *Step 4:* – Repeat the procedure with the last saturated sample.

Titration of thiosulfate: mixture preparation

☐ *Step 1:* – Pipet 10mL of 0.01M KIO_3 into a 125mL (or 250mL) Erlenmeyer flask. Make sure you pick the reagents with the indicated concentration as in the lab there could be more concentrated potasium iodate also. First, add 20mL of water distilled water from a cylinder. Add also 20 drops of 2M HCl and the equivalent of 1mL of solid KI, measured with a 10mL cylinder. Mind you can approximately measure the volume of a solid with a cylinder. Finally, add 40 drops of a 0.2/

☐ *Step 2:* – Repeat the previous step two more times, preparing a total of three Erlenmeyers containing the iodate mixture.

Titration of thiosulfate

☐ *Step 1:* – Rinse a buret with a small portion of $Na_2S_2O_3$ solution, discarding the solution used in the rinses.

☐ *Step 2:* – Fill a buret with $Na_2S_2O_3$ solution.

☐ *Step 3:* – Titrate one of the iodate samples samples adding increments of 1mL of sodium thiosulfate at a time. Swirl the mixture after each addition. You will have reached the equivalency point when the solution becomes colorless. Try placing the flask containing the analyte on a piece of paper. The white background will help see the color change when it occurs.

☐ *Step 4:* – Write down the volume of thiosulfate added to reach the equivalency point to the nearest 0.01mL. This will give an estimate of the volume needed to reach the endpoint.

☐ *Step 5:* – After having estimated the volume of titrate needed to reach the endpoint, we will now speed up the titration by rapidly titrating the second thiosulfate sample.

☐ *Step 6:* – Add thiosulfate to a second iodate sample, stopping 3mL before the estimated endpoint.

☐ *Step 7:* – Add $Na_2S_2O_3$ drop by drop until you reach the endpoint and the solution becomes colorless. Record the buret reading to the nearest 0.01mL.

☐ *Step 8:* – Repeat the titration with the last iodate sample.

Calculations

⓪ This is the molarity of the potassium iodate solution, $c_{\mathrm{KIO_3}}$.

① This is the initial volume in the buret.

② This is the final volume in the buret.

③ This is the volume of sodium thiosulfate used in the standarization, $v_{\mathrm{Na_2S_3O_3}}$:

$$v_{\mathrm{Na_2S_3O_3}} = \textcircled{2} - \textcircled{1}$$

④ This is the molarity of thiosulfate, $c_{\mathrm{Na_2S_3O_3}}$:

$$c_{\mathrm{Na_2S_3O_3}} = 10 \times c_{KIO3} / v_{\mathrm{Na_2S_3O_3}} = 10 \times \textcircled{0} / \textcircled{3}$$

⑤ This is the average sodium thiosulfate concentration.

⑥ This is the molarity of sodium thiosulfate, $c_{\mathrm{Na_2S_3O_3}}$. It corresponds to ⑤ if you standarized the reagent. If not this molarity should be written in the bottle.

⑦ This is the initial volume in the buret.

⑧ This is the final volume in the buret.

⑨ This is the volume of thiosulfate used, $v_{\mathrm{Na_2S_3O_3}}$:

$$v_{\mathrm{Na_2S_3O_3}} = \textcircled{8} - \textcircled{7}$$

⑩ This is the concentration of iodate IO_3^- in the saturated solution, $c_{IO_3^-}$:

$$c_{IO_3^-} = v_{Na_2S_3O_3} \times c_{Na_2S_3O_3}/60 = \textcircled{9} \times \textcircled{6}/60$$

⑩ This is the solubility of the iodate salt, s.

$$s = c_{IO_3^-}/2 = \textcircled{10}\ /2$$

STUDENT INFO

Name: ______________________ Date: __________

Pre-lab Questions

Solubility product of an iodate salt

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1. In this experiment the iodate (IO_3^-) and iodide ions (I^-) react to produce molecular iodine (I_2). Hence, in this redox reaction, a reactant is converted into products by simultaneous oxidation and reduction reactions. This type of redox reaction is referred to as disproportionation. Balance the corresponding disproportionation reaction.

$$IO_3^- \quad + \quad I^- \quad \longrightarrow \quad I_2$$

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2. In this experiment thiosulfate ($S_2O_3^{-2}$) is used in a redox titration. This compound reacts with molecular iodide (I^-), which has a reddish-brown color, to produce the ion colorless iodide (I_2) and tetrathionate. Hence, when all thiosulfate is finished along with the titration the reaction becomes colored. Balance the corresponding redox reaction.

$$S_3O_3^{-2} \quad + \quad I_2 \quad \longrightarrow \quad I^- \quad + \quad S_4O_6^{-2}$$

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3. Both reaction in the previous questions happen consecutively during the experiment. In the redox titration, first iodate reactions with iodine to produce iodine, which at also reacts with thiosulfate to produce iodine. Combine both reaction to generate the overall reacction happening in the titration.

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4. You need 25mL of sodium thiosulfate 0.01M to reach the equivalency point in the titration of a 10mL sample of iodate. Using the ballanced reaction from previous question, carry stoichiometric calcuations to calculate the molarity of iodate in the solution.

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STUDENT INFO

Name: ______________________ Date: ________

Results
EXPERIMENT

Solubility product of an iodate salt

$Na_2S_3O_3$ standarization

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⓪ Molarity of KIO_3, c_{KIO_3} = ______________________

		1	2	3	4
①	Initial Buret Volume (mL)				
②	Final Buret Volume (mL)				
③	$Na_2S_2O_3$ Volume used (mL)				
④	$Na_2S_2O_3$ Concentration (M)				

⑤ Mean $Na_2S_2O_3$ Concentration (M) = ______________________

Titration of saturated solution

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Molarity of $Na_2S_2O_3$, $c_{Na_2S_2O_3}$ = ______________________

		1	2	3	4
⑥	Initial Buret Volume (mL)				
⑦	Final Buret Volume (mL)				
⑧	$Na_2S_2O_3$ Volume used (mL)				
⑨	IO_3^- Concentration (M)				
⑩	Solubility (M)				

⑪ Mean solubility (M) = ______________________ ⑫ Mean K_{sp} = ______________________

STUDENT INFO

Name: ____________ Date: ________

Post-lab Questions

Solubility product of an iodate salt

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1. Write down the solubility equilibrium of calcium iodate.

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2. Write down the expression of the solubility constant of calcium iodate in terms of ion molarities.

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3. Write down the expression of the solubility constant of calcium iodate in terms of solubilities.

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4. Write down the expression of the solubility of calcium iodate in terms of the solubility product constant.

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EXPERIMENT

Units

Goal

The goal of this experiment is to familiarize with the *conversion of units*. You will carry basic unit conversions.

Background

Let's consider the following measurements: **1 k**m, **2 c**m, 3 m, which can be read as one centimeter, two kilometers and three meters. The word kilo (**k**) and centi (**c**) are called prefixed whereas meter (m) is a simple unit. Kilometer is larger than meter, whereas centimeter is smaller than a meter. Prefixes such as kilo or centi are attached to units in order to make numbers more manageable. For example, the radius of the earth is 6356 km, and this number is easier to handle than 6356000m. At the same time, we can attach any prefix to different units. Hence, we can talk about a centimeter (**c**m) but also about a centisecond (**c**s) or centiliter (**c**L). All these units have the same prefix. The Table below lists some of the metric prefixes, their symbols, and their decimal values.

How to identify prefixes?

Look for example in the measurement 2 cm. Centi (c) is the prefix and means 1×10^{-2} and meter (m) is the unit which refers to length. Another example, 7 kg means kilogram. Kilo (k) is the prefix and means 1×10^{3}, whereas gram (g) is the unit that refers to mass. The prefix refers to the first letter whereas the unit refers to the last letter.
Would you prefer to be paid a kilodollar, a dollar or a centidollar? A unit with a prefix can be bigger or smaller than the plain unit–this is the unit without prefix–, depending on the prefix. The following prefixes make the unit smaller: deci, centi, milli, micro, nano, pico and femto. For example a fs (femtosecond) is smaller than a s (second). Differently, the following prefixes make the unit larger: Tera, Giga, Mega. For example a Tb (terabite) is larger than a b (bite). Bite is a unit used in computer science.

How to write unit equalities?

Unit equalities are simple expressions that relate a unit with a unit with prefix. For example: one centimeter (cm) is 1×10^{-2}m. Hence we can write this as a unit equality:

$$\boxed{1cm = 1 \times 10^{-2} m} \text{ unit equality}$$

How to write conversion factors?

Let's compare cm and m. The first, cm, is a unit with a prefix, whereas m is simple a unit of length without a prefix. In order know how many m are there in a cm we need to write down a conversion factor. Think in the prefixes as synonymous of a number. In this way, centi stands for 1×10^{-2}, so

$$\boxed{\frac{1cm}{1 \times 10^{-2} m} \quad \text{or} \quad \frac{1 \times 10^{-2} m}{1cm}} \text{ conversion factor}$$

Using Conversion Factors

Unit equalities in the form of conversion factors are used to convert a unit into another. Sometimes one wants to get rid of a prefix, such as when we transform centimeter (cm) into meter (m). Sometimes, one wants to convert a prefix into another prefix. An example would be converting centimeters (cm) to millimeters (mm). Let's work on some example.

Removing or adding prefixes

Imagine that you need to remove a prefix from a unit, and convert 3 km (we will call this one the origonal unit) in meters (this is the final unit). First, you would need the conversion factor corresponding to the prefix (centi) from Table 2. Then you need to arrange the conversion factor placing the prefix at the bottom of the fraction. This will cancel out the prefix in the original unit and in the bottom part of the conversion factor, hence leaving the final unit on top of the conversion factor. The arrangement would be:

$$3\cancel{km}\times\frac{1\times10^{3}m}{1\cancel{km}}=3000m$$

Imagine now that you need to add a prefix into a unit, and convert 4000 m in km. The same would apply for this case, but now you will have to arrange the conversion factor so that the prefix is on the top:

$$4000\cancel{m}\times\frac{1\text{ km}}{1\times10^{3}\cancel{m}}=4km$$

Different prefixes

Prefix	Symbol	Meaning	Value
peta	P	1000000000000000	1×10^{15}
tera	T	1000000000000	1×10^{12}
giga	G	1000000000	1×10^{9}
mega	M	1000000	1×10^{6}
kilo	k	1000	1×10^{3}
hecto	h	100	1×10^{2}
deca	da	10	1×10^{1}
–	–	1	1×10^{0}
deci	d	0.1	1×10^{-1}
centi	c	0.01	1×10^{-2}
milli	m	0.001	1×10^{-3}
micro	μ	0.00001	1×10^{-6}
nano	n	0.000000001	1×10^{-9}
pico	p	0.000000000001	1×10^{-12}
femto	f	0.000000000000001	1×10^{-15}

STUDENT INFO

Name: ______________________ Date: __________

Worksheet

Units

1. Fill in the following unit equalities indicating the value of *a* and *b*.

Unit equality	a	b
a cs= b s		
a μm= b m		
a b= b Tb		
a L= b dL		

2. Fill in the following conversion factors indicating the value of *a* and *b*.

conversion factors	a	b
$\frac{a\ mL}{b\ L}$		
$\frac{a\ s}{b\ ps}$		
$\frac{a\ kg}{b\ g}$		

3. Fill in the following conversion factors calculations.

conversion factors	a	b	c
$9km \times \frac{a\ m}{b\ km} = c\ m$			
$100cm \times \frac{a\ m}{b\ cm} = c\ m$			
$1000m \times \frac{a\ km}{b\ m} = c\ km$			

EXPERIMENT

Units with power and switching prefixes

Goal

The goal of this experiment is to familiarize with the *square and cubic units* and to carry prefix switches. You will carry basic square and cubic unit conversions.

Background

In order to switch a prefix into another prefix, such as transforming 30 millimeters (30 mm) into centimeters (cm), you will need two different conversion factors: the first conversion factor will remove the original unit (mm) introducing an intermediate unit, meters (m), whereas the second conversion factor will remove the intermediate meter and introduce the final unit (cm). You will get the conversion factors from Table 2. You will arrange the first conversion factor so that the original unit cancels out with the bottom of the first conversion factor, giving you an intermediate unit. You will arrange the second conversion factor so that the intermediate unit cancels out with the bottom of the second conversion factor giving the final unit. For this example:

$$30\cancel{mm} \times \frac{1 \times 10^{-3}\cancel{m}}{1\cancel{mm}} \times \frac{1cm}{1 \times 10^{-2}\cancel{m}} = 3cm$$

Square or cubic units

How big is your apartment? You might be living in a $750ft^2$ loft in Brooklyn or in a larger house Upstate. Often times we encounter cubic or square units such as cubic centimeter (cm^3) or square feet (ft^2). The equivalencies for cubic or square units should take into account the unit power (power of two or power of three). If $1cm = 1 \times 10^{-2}m$, for square units the relation should be squared and $1cm^2 = 1 \times (10^{-2})^2 m^2 = 1 \times 10^{-4} m^2$. Another example, for the case of mm and mm^3:

$$\boxed{\frac{1mm}{1 \times 10^{-3}m} \quad \text{and} \quad \frac{1mm^3}{1 \times 10^{-9}m^3}}$$

Let us work on an example in which we want to convert $30m^2$ into m^2:

$$30\cancel{m^2} \times \frac{1cm^2}{1 \times 10^{-4}\cancel{m^2}} = 3 \times 10^5 cm^2$$

Different prefixes

Prefix	Symbol	Meaning	Value
peta	P	1000000000000000	1×10^{15}
tera	T	1000000000000	1×10^{12}
giga	G	1000000000	1×10^{9}
mega	M	1000000	1×10^{6}
kilo	k	1000	1×10^{3}
hecto	h	100	1×10^{2}
deca	da	10	1×10^{1}
–	–	1	1×10^{0}
deci	d	0.1	1×10^{-1}
centi	c	0.01	1×10^{-2}
milli	m	0.001	1×10^{-3}
micro	μ	0.00001	1×10^{-6}
nano	n	0.000000001	1×10^{-9}
pico	p	0.000000000001	1×10^{-12}
femto	f	0.000000000000001	1×10^{-15}

STUDENT INFO

Name: ____________________ Date: __________

Worksheet

Units with power and switching prefixes

1. Fill in the following unit equalities involving powers indicating the value of *a* and *b*.

Unit equality	a	b
$a\ cm^2 = b\ m^2$		
$a\ \mu m^3 = b\ m^3$		
$a\ dm^3 = b\ m^3$		
$a\ cm^3 = b\ m^3$		

2. Fill in the following conversion factors calculations involving powers.

conversion factors	a	b	c
$9m^3 \times \frac{a\ dm^3}{b\ m^3} = c\ dm^3$			
$100cm^2 \times \frac{a\ m^2}{b\ cm^2} = c\ m^2$			
$1000cm^3 \times \frac{a\ m^3}{b\ cm^3} = c\ m^3$			

3. Fill in the following prefix switches.

Conversion factor

$$9cm \times \frac{a\ m}{b\ cm} \times \frac{c\ mm}{d\ m} = e\ mm$$

a	b	c	d	e

EXPERIMENT

Classification of matter

Goal

The goal of this experiment is to classify different substances and mixtures in terms of its composition

Background

We can classify chemical substances in terms of what they are made of. Some are made of just one thing, others contain different components. At the same time, some substances look like they are made of one thing and indeed they are made of many components. This section covers the classification of matter according to the purity of substances and mixtures. It also elaborates on the types of mixtures one can find. The final part of the section deals with density.

Pure Substances

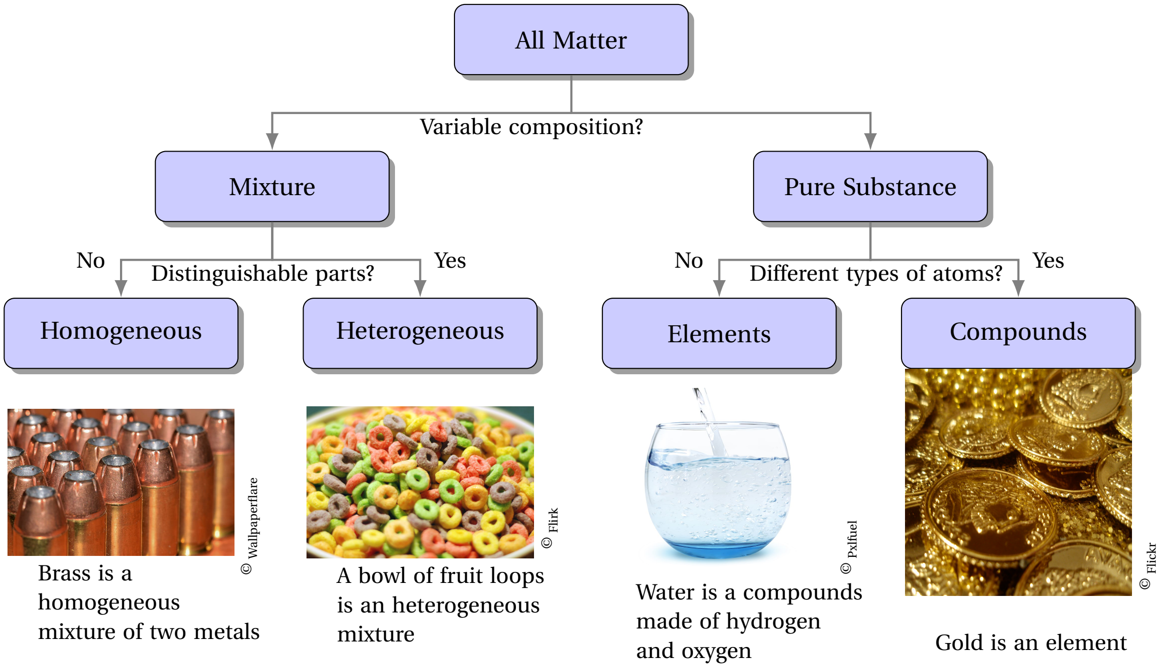
Pure substances have a definite composition, that is, are only made of one thing. There are two different types of pure substances: elements and compounds. *Elements* are composed of only one type of atom. Examples are silver, iron, and aluminum that all contain one type of substance, and iron is only made of iron atoms, for example. *Compounds* are combinations of different elements. For example, water, H_2O is made of a combination of hydrogen and oxygen atoms.

Mixtures

Mixtures are physical combinations of different substances. For example, air is a mixture of oxygen and nitrogen.

Types of Mixture

Mixtures can be homogeneous or heterogeneous. In a *homogeneous mixture*–also known as solutions–the composition is uniform throughout the sample. An example of an homogeneous mixture is air, which contains oxygen and nitrogen or salty water, a solution of salt and water. *Heterogeneous mixtures* are mixtures in which the components are not uniformly distributed throughout the sample. An example would be a chocolate chip cookie in which you can differentiate the dough and the chocolate.



Brass is a homogeneous mixture of two metals

A bowl of fruit loops is an heterogeneous mixture

Water is a compounds made of hydrogen and oxygen

Gold is an element

Figure 30: Classification of the matter

STUDENT INFO

Name: ____________________ Date: __________

Worksheet

Classification of matter

1. Classify the following as mixtures (homogeneous or heterogeneous) or pure substances. Select with an X the correct option.

	Homog. mixture	heterog. mixture	Element	Compound
Vinegar				
Distilled water				
Air				
Table salt				
Apple juice				
Brass				
sand				

EXPERIMENT

The Periodic table

Goal

The goal of this experiment is to familiarize with the *periodic table of the elements*. You will analyze the rows and columns of the tables and the different classifications of elements.

Background

The periodic table contains all elements that form the matter arranged in columns and rows. Every element has a different name and a symbol that represents their name. The tabular arrangement of elements in the form of rows and columns allow further classification according to their properties. This section will cover the different features of the periodic table.

Elements and Symbols

Elements cannot be broken down into simpler substances. For example aluminum is an element only made of aluminum atoms and if you analyze the composition of a piece of this metals you would only find aluminium atoms. Chemical symbols are one- or two-letter abbreviations that represent the names of the elements. Only the first letter is capitalized and if a second letter exist in the element's name, the second letter should be lowercase. For example, the chemical symbol for aluminum is Al with capital A and lowercase l. The periodic table contains the symbols of all elements.

Periods and groups

Learn more about *the periodic table.*

Visual

The periodic table contains elements arranged in rows and columns. The horizontal rows are called *periods* and the vertical columns are called *groups*. For example, the first period contains hydrogen (H) and helium (He), and the second group contains Beryllium (Be), Magnesium (Mg), Calcium (Ca), Strontium (Sr), Barium (Ba) and Radium (Ra). There are seven periods (periods 1-7) and 18 groups. Some of the groups are labeled with an A (e.g. group 8A) whereas others are labeled with a B (e.g. group 8B). The groups numbers can be written with roman numbers and a letters (A or B) or with a modern group numbering of 1-18 going across the periodic table. For example, the group 2 (Mg-Ra) can also be called IIA, and the group 13 (B-Ti) is also known as IIIA.

Classification of groups

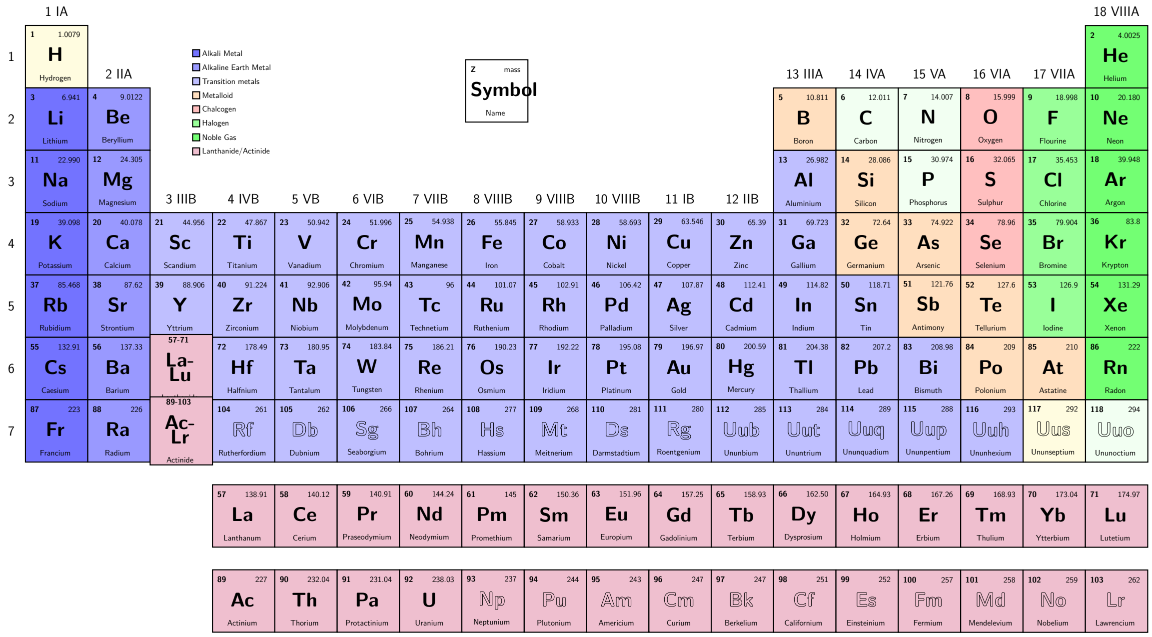
Some of the groups in the periodic table have specific names such as alkali metals, alkaline earth metals, transition metals, halogens or noble gases. Alkali metals are the group 1A elements: lithium (Li), sodium (Na), potassium (K), rubidium (Rb), cesium (Cs), and francium (Fr). Alkali elements are soft and shiny metals, and they are also good conductors of heat and electricity, having low melting points. Alkali earth metals are the group 2A (2) elements: beryllium (Be), magnesium (Mg), calcium (Ca), strontium (Sr), barium (Ba), and radium (Ra). Transition metals are the elements from group 3 to 12 and they are located in the middle of the table. Halogens are the group 7A (17) elements: fluorine (F), chlorine (Cl), bromine (Br), iodine (I), and astatine (At). Halogens are very reactive elements. Finally, noble gases are the group 8A (18) elements: helium (He), neon (Ne), argon (Ar), krypton (Kr), xenon (Xe), and radon (Rn). They are inert and rarely combine with other elements in the periodic table, like novelty: have you ever meet any royal?

Metals, Nonmetals, and Metalloids

Overall, the elements of the periodic table can also be classify as metals, nonmetals, and metalloids. Metals are those elements on the left of the table with a blue color and nonmetals are the elements on the right of the table with a green color. The elements between metals and nonmetals are called metalloids and include only B, Si, Ge, As, Sb, Te, Po, and At. Metals are shiny solids and usually melt at higher temperatures. Some examples of metals are Gold (Au) or Iron (Fe). Nonmetals are often poor conductors of heat and electricity with low melting points. They also tend to be matt (non-shinny), malleable, or ductile. Some examples of nonmetals are Carbon (C) or Nitrogen (N). Metalloids are elements that share some properties with metals and others with the nonmetals. For example, they are better conductors of heat and electricity than the nonmetals, but not as good conductors as the metals. The metalloids are semiconductors because they can act as both conductors and insulators under certain conditions. An example of metalloids is Silicon (Si) that should not be confused from silicone, a chemical employed in prosthetics.

How to classify Hydrogen

At first sight, hydrogen (H) may seem to be put in the wrong spot at the periodic table. Although it is located at the top of Group 1A (1), it is not an alkali metal, as it has very different properties. Thus hydrogen does not belong to the alkali metals, being a nonmetal.



- ■ Alkali Metal
- ■ Alkaline Earth Metal
- ■ Transition metals
- ■ Metalloid
- ■ Chalcogen
- ■ Halogen
- ■ Noble Gas
- ■ Lanthanide/Actinide

Z mass
Symbol
Name

	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 4.0025 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 Flourine	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 65.39 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 83.8 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 95.94 Mo Molybdenum	43 96 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	52 127.6 Te Tellurium	53 126.9 I Iodine	54 131.29 Xe Xenon
6	55 132.91 Cs Caesium	56 137.33 Ba Barium	57-71 La-Lu	72 178.49 Hf Halfnium	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	104 261 Rf Rutherfordium	105 262 Db Dubnium	106 266 Sg Seaborgium	107 264 Bh Bohrium	108 277 Hs Hassium	109 268 Mt Meitnerium	110 281 Ds Darmstadtium	111 280 Rg Roentgenium	112 285 Uub Ununbium	113 284 Uut Ununtrium	114 289 Uuq Ununquadium	115 288 Uup Ununpentium	116 293 Uuh Ununhexium	117 292 Uus Ununseptium	118 294 Uuo Ununoctium

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	70 173.04 Yb Ytterbium	71 174.97 Lu Lutetium
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

STUDENT INFO	
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Worksheet

The periodic table

1. In the first table, give, the chemical symbol, identify the name, group, period, metallic character (metal, nonmetal, metalloid) and group (alkali, alkaline earth, halogen, nobel gas, lanthanide, actinide) class of the element. Write "N/A" for elements not belonging to these defined classes.

Symbol	Chemical Name	Period	Group	Metallic Character	Group Class
Ba					
Cd					
P					
Br					
Ne					
Na					

2. In the second table, give, the chemical name, identify the symbol, group, period, metallic character (metal, nonmetal, metalloid) and group (alkali, alkaline earth, halogen, nobel gas, lanthanide, actinide) class of the element. Write "N/A" for elements not belonging to these defined classes.

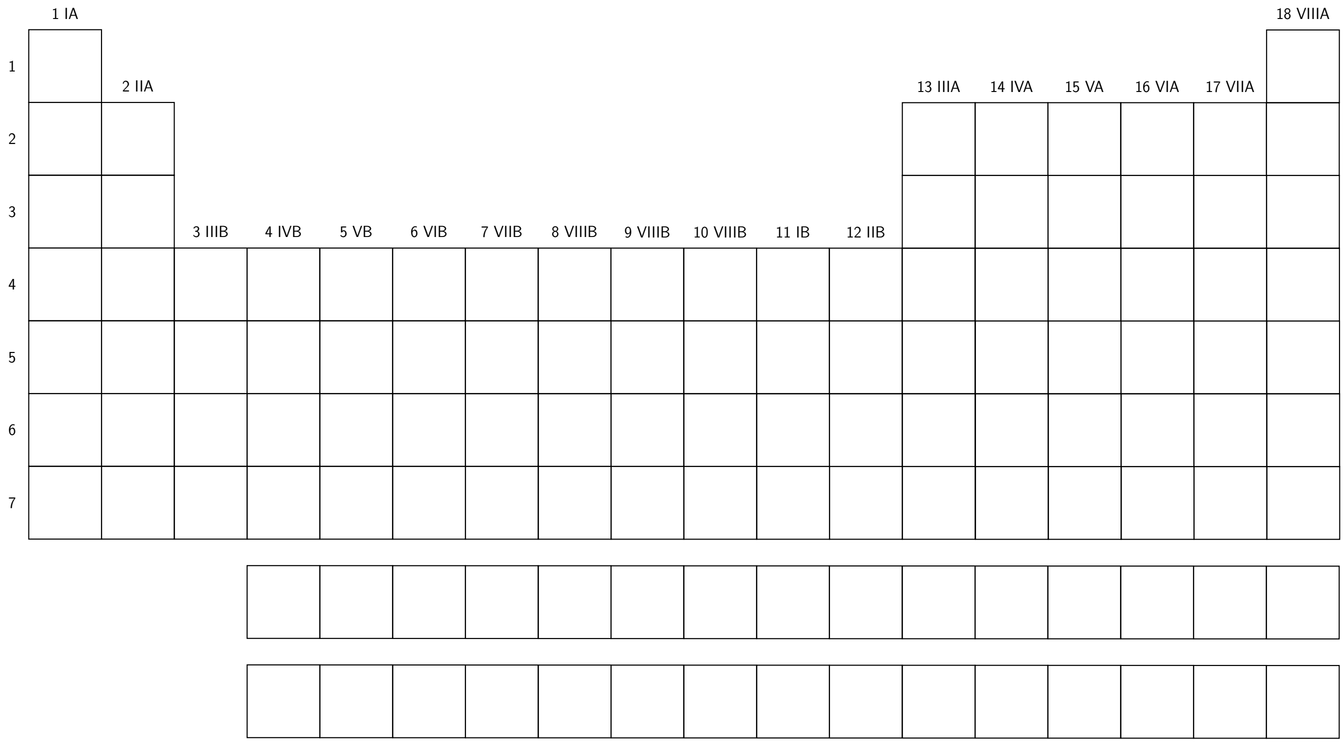
Symbol	Chemical Name	Period	Group	Metallic Character	Group Class
	Iron				
	Potassium				
	Boron				
	Radon				
	Radium				
	Scandium				

3. In the last table, give, the group and period, identify the name, symbol, metallic character (metal, nonmetal, metalloid) and group (alkali, alkaline earth, halogen, nobel gas, lanthanide, actinide) class of the element. Write "N/A" for elements not belonging to these defined classes.

Symbol	Chemical Name	Period	Group	Metallic Character	Group Class
		1	1		
		5	18		
		2	16		
		4	5		
		6	12		
		6	7		

4. Given the following outline of a periodic table, show the location of the alkali, alkaline earth metals, halogens and noble gases. Also draw the line separating the metals and nonmetals from the metalloids. Finally, locate the following elements:

- ✔ tin
- ✔ carbon
- ✔ strontium
- ✔ molybdenum
- ✔ xenon
- ✔ Europium
- ✔ helium
- ✔ Copper



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EXPERIMENT

Electrolytes and insoluble compounds

Goal

The goal of this experiment is to familiarize with *insoluble compounds and electrolytes*. You will identify compounds based on its soluble and electrolyte character.

Background

On one hand, electrolytes are compounds that conduct the electricity once dissolved in water. Differently, nonelectrolytes are compounds that do not conduct the electricity once dissolved in water. On the other hand, insoluble compounds are not soluble in water, whereas soluble compounds can be dissolved in water. This section covers the properties of electrolytes and insoluble (and soluble) compounds. At the end of this section, you should be able to classify a chemical in terms of its electrolyte type and solubility character.

Soluble and insoluble salts

Learn more about 🌍 *the solubility rules.*

Visual

Soluble compounds dissolve in water, whereas insoluble compounds do not. For example, barium chromate ($BaCrO_{4\,(s)}$) is an insoluble salt. How do we know that? Table 5 will help you predict the solubility of a salt. In order to do this, you need to start by assess the right ion (the anion, $CrO_4{}^{2-}$) located on the left column of Table 5. After that you need to assess the left ion (the cation, Ba^{2+}) located on the right column. If you follow this, you will see that chromate is insoluble and barium is not part of any exception. Let us predict for example the soluble/insoluble nature of $CaSO_4$, calcium sulfate. We start by looking for $SO_4{}^{2-}$ in the left column to find out is soluble. Next we continue in the same line as $SO_4{}^{2-}$ and look for the ion in the left Ca^{2+}. In conclusion, even when $SO_4{}^{2-}$ is soluble, when combined with Ca^{2+}, we have that $CaSO_4$ is insoluble, and overall $CaSO_{4\,(s)}$ is insoluble.

Strong electrolytes

Strong electrolytes completely dissociate in water. Hence, in a solution of a strong electrolyte you will only have ions and never molecules. Strong electrolytes are typically ionic compounds such as $MgCl_2$ or NaCl (table salt). We represent the dissociation of a strong electrolyte with a single arrow, meaning that the reaction proceeds to completion and for the example below, in the solution we will only have ions ($Mg^{2+}_{(aq)} + 2\,Cl^{-}_{(aq)}$) and not molecules ($MgCl_{2\,(s)}$):

$$MgCl_{2\,(s)} \xrightarrow{H_2O} Mg^{2+}_{(aq)} + 2\,Cl^{-}_{(aq)}$$

Weak electrolytes

Weak electrolytes partially dissociate in water, and this is indicated by means of a chemical reaction with a double arrow. Hence in a solution of a weak electrolyte you will have ions as well as molecules at the same time. Examples of weak electrolytes are hydrofluoric acid, water, ammonia or acetic acid. The dissociation of hydrochloric acid (HF) proceeds as:

$$HF_{(g)} \xrightleftharpoons{H_2O} H^{+}_{(aq)} + F^{-}_{(aq)}$$

Acetic acid (CH_3COOH) is an important weak electrolyte and its dissociation proceeds somehow in a peculiar way:

$$CH_3COOH_{(l)} \xrightleftharpoons{H_2O} CH_3COO^{-}_{(aq)} + H^{+}_{(aq)}$$

Table 5 Soluble and insoluble compounds

Ions that form *soluble* compounds...	... **except** when combined with
Group I ions (Na^+, Li^+, K^+, etc)	no exceptions
Ammonium ($NH_4{}^+$)	no exceptions
Nitrate ($NO_3{}^-$)	no exceptions
Acetate (CH_3COO^-)	no exceptions
Hydrogen carbonate ($HCO_3{}^-$)	no exceptions
Chlorate ($ClO_3{}^-$)	no exceptions
Halide (F^-, Cl^-, Br^-)	Pb^{2+}, Ag^+ and $Hg_2{}^{2+}$
Sulfate ($SO_4{}^{2-}$)	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, $Hg_2{}^{2+}$ and Pb^{2+}
Ions that form *insoluble* compounds...	**... except when combined with**
Carbonates ($CO_3{}^{2-}$)	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Chromates ($CrO_4{}^{2-}$)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}or ammonium ($NH_4{}^+$)
Phosphates ($PO_4{}^{3-}$)	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Sulfides (S^{2-})	group I ions (Na^+, Li^+, K^+, etc) or ammonium ($NH_4{}^+$)
Hydroxides (OH^-)	group I ions (Na^+, Li^+, K^+, etc) or Ca^{2+}, Mg^{2+}, Sr^{2+}or ammonium ($NH_4{}^+$)

Nonelectrolytes

Nonelectrolytes do not dissociate in water. Hence a solution of a nonelectrolyte will only contains molecules and not ions. Examples of nonelectrolytes are carbon-based chemicals such as methanol, ethanol, urea or sucrose. The dissociation of urea for example CH_4N_2O proceeds as:

$$CH_4N_2O\,(s) \xrightarrow{H_2O} CH_4N_2O_{(aq)}$$

Identify the electrolyte character of a chemical

You can use Table 4 to identify the electrolyte character of a chemical. Ionic compounds are in general strong electrolytes, and most acids as well. There is four important weak electrolytes: water, acetic acid, ammonia and hydrofluoric acid. Covalent compounds are in general nonelectrolytes. Organic compounds, compounds based on carbon atoms (e.g. $C_{12}H_{22}O_{11}$) are in general nonelectrolytes.

Table 4 Different types of electrolytess

Electrolyte Type	Dissociation	Particles in solution	Examples
Strong	Fully	Mostly ions	Ionic Compounds and most acids and bases: NaCl, NaOH, HCl, $MgCl_2$, H_2SO_4, etc·
Weak	Partially	Ions & molecules	NH_3, CH_3COOH (acetic acid), HF, H_2O
Nonelectrolytes	No	molecules	Most covalent compounds: CH_3OH(methanol), CH_3CH_2OH(ethanol), $C_{12}H_{22}O_{11}$ (sucrose), CH_4NO_2(urea)

Breaking down chemicals into ions

Electrolytes–in particular strong electrolytes–dissociate producing ions. This way, a solution of for example NaCl does not contain NaCl molecules but $Na^+_{(aq)}$ cations and $Cl^-_{(aq)}$ anions. Hence it is important to correctly break down electrolytes into ions. In order to do this, you need to revert the combination of ions that produce a given chemical while making sure the charges are balanced. For example, let us beak magnesium chloride $MgCl_{2\,(aq)}$ into ions. This is a strong electrolytes formed by magnesium cations and chloride anions. The valence of magnesium is +II and the valence of chlorine is -I. The $MgCl_2$ formula also tells us we have one magnesium and two chlorines. The overall process is:

$$MgCl_{2\,(aq)} \longrightarrow Mg^{2+}_{(aq)} + 2\,Cl^{-}_{(aq)}$$

Another example, magnesium nitrate $Mg(NO_3)_2$. This strong electrolyte–as this is an ionic salt–is made of lithium with valence +I and nitrate with valence -I. The formula indicated we have one $Mg^{2+}_{(aq)}$ and two $NO_3{}^-{}_{(aq)}$. Hence:

$$Mg(NO_3)_{2\,(aq)} \longrightarrow Mg^{2+}_{(aq)} + 2\,NO_3{}^-{}_{(aq)}$$

Example

Break down the following chemicals into ions, if possible: $K_2CrO_{4\,(aq)}$, $Ba(NO_3)_{2\,(aq)}$, $BaCrO_{4\,(s)}$, *and* $KNO_{3\,(aq)}$.

Answer: We can only break down into ions ionic compounds and oxosalts that are not solid. From the list of chemicals in the example, we will not be able to break down $BaCrO_{4\,(s)}$ into ions as it is a solid. From the other chemicals, $K_2CrO_{4\,(aq)}$ is named potassium chromate and contains $2K^+_{(aq)}$ and $CrO_4{}^{2-}{}_{(aq)}$ ions. Barium nitrate–$Ba(NO_3)_{2\,(aq)}$–will produce $Ba^{2+}_{(aq)}$ and $2NO_3{}^-{}_{(aq)}$. Finally, potassium nitrate– $KNO_{3\,(aq)}$–will produce $K^+_{(aq)}$ and $NO_3{}^-{}_{(aq)}$.

STUDENT INFO

Name: ______________________ Date: __________

Worksheet

Electrolytes and insoluble compounds

1. Classify the following compounds in terms of their electrolyte character by selecting with an X the correct classification: strong electrolyte, weak electrolyte, or nonelectrolyte.

Compound	Strong electrolyte	Weak electrolyte	Nonelectrolyte
KCl			
KOH			
CO_2			
HF			
Li_3N			

2. Classify the following compounds in terms of their soluble character by selecting with an X the correct classification.

Compound	Soluble	Insoluble
KCl		
Ag_2S		
$Be(OH)_2$		
$NaHCO_3$		
$NaCO_3$		

3. Name and dissociate the following compounds into ions.

Compound	Ions	Chemical Name
NaCl		
$Be(OH)_2$		
Li_3N		
$Ca_3(PO_4)_2$		
KNO_3		

EXPERIMENT

Concentrations units

Goal

The goal of this experiment is to familiarize with some of the different *concentrations units* employed in chemistry. You will calculate molarities and mass percent.

Background

Solutions are homogeneous mixtures of two components. The state of the matter of both components of the mixture or their polarity affect the formation of a solution. For example, a solution will not result from mixing oil and water as they have different polar character and it will form from mixing table salt and water as both are polar chemicals. At the same time, the more solute you add to a solution the more concentrated the solution will be. This section covers polarity and the composition of solutions.

What makes a solution?

Solutions are homogeneous mixtures of a solute and a solvent. Homogeneous means that if you look at the mixture you will not be able to differentiate both components and you will only see it as a whole. In a solution, the solute is the component of the mixture in less amount, whereas the solvent is the component in a larger amount. Think about mixing sugar with water. Sugar is sweet and water tasteless. When you mix both, you form a solution of sugar (solute) in water (solvent) and you will not see sugar in the solution as it is dissolved. In this particular example, sugar will be the solute in the solution, as the sugar is in less amount than water. Is important to remember that a solution is the result or mixing a solute and a solvent:

$$\boxed{\text{Solution} = \text{Solute} + \text{Solvent}}$$

Types of solutions

Learn more about *solutions.*

Visual

You can prepare different types of solutions by mixing a solid and a liquid, like when you mix sugar and water, or salt and water. You can create solutions as well by mixing two liquids or two solids. Examples are vinegar–a liquid solution of acetic acid (liquid) in water (liquid)–or steel– a solid solution that contains iron and carbon, both solids.

Concentration of solutions

One of the most important properties of a solution is its concentration. The larger the concentration of a solution the more solute in the solution. There are many different concentration units, such as molarity or mass percent concentration. This section will introduce you to some of the most important concentration units. The concentration of a solution refers to the amount of solute with respect to the amount of solution. The larger concentration the larger number of solute particles with respect to the solution.

Mass percent concentration

The mass percent (m/m) is the amount of solute in grams per grams of solution in percent form

$$\boxed{m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100}$$

Example

A NaCl solution is prepared by mixing 4g of NaCl with 50g of H_2O. Calculate the percent (m/m) of the solution.

Answer: We need the grams of solute and the grams of solution. The grams of solute are given (4g of NaCl), whereas the grams of solution result from adding the grams of solvent and solute: 54 g of solution. Using the formula for the percent (m/m), he have:

$$m/m = \frac{\text{4 g of solute}}{\text{54 g of solution}} \times 100 = 7.4\%$$

This means that by mixing 4g of NaCl with 50g of H_2O you prepare a 7.4% (m/m) solution.

Molarity concentration

The molarity (M) is the moles of solute per L of solution.

$$\boxed{M = \frac{\text{moles of solute}}{\text{L of solution}}}$$

Example

A NaCl solution is prepared by mixing 4g of NaCl (MW=58.4g/mol) with 50 g of water until a final volume of 52mL of solution. Calculate: (a) the mass percent (m/m) concentration; (b) the molarity.

Answer: (a) to calculate the mass percent (m/m) we just need the grams of solute and the grams of solution–that is four plus fifty. Both numbers are already given:

$$m/m = \frac{\text{g of solute}}{\text{g of solution}} \times 100 = \frac{\text{4 g of solute}}{\text{54 g of solution}} \times 100 = 9.2\%$$

(b) To calculate molarity we need the moles of solute and the liters of solution. We have the mL of solution, that can be converted to L: 52mL= $5.2 \times 10^{-2} L$. To calculate the moles of solute, we will use the grams of solute and the molar mass to convert this value into moles: $4g/58.4g/mol$ =0.068moles. Plugging all values into the molarity formula:

$$M = \frac{\text{moles of solute}}{\text{L of solution}} = \frac{\text{0.068 moles of solute}}{5.2 \times 10^{-2}\text{L of solution}} = 1.31M$$

Dilution

Dilution is the process for preparing a diluted solution from a more concentrated solution. Solutions are often times stored in a stock room in concentrated form. These stocks should be diluted before use. In order to dilute a solution we need to take a certain amount of the concentrated solution and add water. When adding water, the number of moles of solute does not change, and the concentration always decreases. We have a concentrated solution (c_1) and we need to prepare a certain volume (V_2) of a more diluted solution (c_2). The question is how much volume of the concentrated solution (V_1) we need to take. In order to answer this we should use the following formula:

$$\boxed{c_1 \cdot V_1 = c_2 \cdot V_2}$$

Example

How many liters of a 3M NaCl solution are required to prepare 2L of a 1M diluted NaCl solution.

Answer: We have a concentrated solution of 3M molarity and we want to prepare a more dilute solution. In particular 2L of a 1M. Hence: $c_1 = 3$ and $c_2 = 1M$ and $V_2 = 2L$. Using the dilution formula:

$$3M \cdot V_1 = 1M \cdot 2L$$

Solving for V_1 we have a volume of 0.66L.

STUDENT INFO

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Worksheet

Concentrations units

1. Sodium hydroxide NaOH is a chemical used in drain cleaners. A solution is prepared by mixing 25g of NaOH in 250g of water. Calculate the solute mass percent (m/m) concentration.

2. Vinegar is a (m/m) 5% acetic acid solution. How many grams of acetic acid are there in 2g of vinegar.

3. An HCl solution is prepared by mixing 4 moles of HCl with water reaching a volume of 250mL. Calculate the molarity of the solution.

4. How many mL of a 3M KCl solution contains 0.06 moles of KCl.

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5. How many mL of a 4M NaCl (MW=58$g \cdot mol^{-1}$) solution contains 5 grams of NaCl.

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6. How many grams of solute are there in 100mL of a 0.01M HNO_3 (MW=63$g \cdot mol^{-1}$) solution.

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36 Redox Numbers

EXPERIMENT

Redox number

Goal

The goal of this experiment is to familiarize with the calculation of redox numbers.

Background

The redox number, also called oxidation number or oxidation state helps compare elements that have lost (or gained) different charge. When comparing two elements with different redox numbers, the larger this number the more electrons the element has lost. Similarly, the smaller this number the smaller this number the more electrons the element has won.

Rules to calculate redox numbers

We indicate redox numbers with roman number on top of the element. For example the redox number of manganese in this compound is +7: $\underline{Mn^{VII}}O_4{}^-$. The redox number can be a non integer number. There is five rules to identify the redox number of an element.

- **Rule 1** Single atoms or elements have zero redox number. Examples are Na or H_2, both with redox zero.
- **Rule 2** Monoatomic ions have redox number equal to their charge. Examples are Na^+ or Cl^- with redox +1 and −1, respectively.
- **Rule 3** The redox number of fluorine is −1
- **Rule 4** The redox number of hydrogen on its covalent (e.g. H_2O) compounds is +1.
- **Rule 5** The redox number of oxygen in normal oxides (e.g. MgO) is normally −2, with the exception of peroxides (e.g. H_2O_2) in which is −1.

Calculating the redox number

How do we calculate the redox number for example of manganese in this chemical: $\underline{Mn}O_4{}^-$, permanganate. IN order to do this, we need to set up a formula so that the redox numbers of all elements in the molecule–taking into account the number of atoms in the molecule–equals to the charge. In the case of permanganate, let us call x to the redox number of manganese. We know the redox of oxygen is −2 and we have four oxygens in the molecule. We also know the charge of the ion is −1. Therefore we have:

$$x + 4 \cdot (-2) = -1 \text{ (redox equation)}$$

If we solve for x we obtain a redox number of manganese of VII.

Example

Calculate the redox number of the elements underlined in the following molecules: (a) $K_2\underline{C}O_3$ *and (b)* $H_2\underline{C}O$.

Answer: Let us set up the redox equation for the first compound, knowing that the redox of oxygen is −2 and potassium +1. The unknown variable x represents the redox number of the underlined element. We have:

$$2 \cdot (+1) + x + 3 \cdot (-2) = 0$$

Mind we have two potassium and three oxygens hence we need to time the redox of K by two and the redox of O by three. If we solve for x we obtain a redox number for carbon of IV. The redox equation for the second example is:

$$2 \cdot (+1) + x + (-2) = 0$$

Mind that according to the redox rules, the redox number of oxygen is $+1$. Solving for x we have a redox number of zero.

STUDENT INFO

Name: ____________________ Date: __________

Worksheet

Redox number

1. Indicate the redox number of the underlined element, indicating the redox equation:

Compound	Redox equation	Redox #
$H\underline{Mn}O_4$		
$\underline{Ni}O_2$		
$Li\underline{Bi}O_3$		
$K\underline{S}_2O_3$		
$\underline{N}O_2{}^-$		
$\underline{Ag}$		
$(NH_4)_2\,\underline{Ce}\,(SO_4)_3$		
$\underline{Fe}\,(OH)_3$		
$\underline{Fe}\,O_4{}^-$		

EXPERIMENT

Redox balancing

Goal

The goal of this experiment is to familiarize with balancing redox reactions.

Background

Redox reaction are different than acid-base or precipitation reaction. They contain the same chemical element in two different states resulting from the loss or win of electrons. For example:

$$2\,Al_{(s)} \quad + \quad 3\,Cu^{2+}_{(aq)} \quad \longrightarrow \quad 2\,Al^{3+}_{(aq)} \quad + \quad 3\,Cu_{(s)} \text{ (redox reaction)}$$

We have that neither $Al_{(s)}$ or $Cu^{2+}_{(aq)}$ are an acid or a base. Also there is no product precipitate. Hence, this reaction is not an acid-base reaction or a precipitation reaction. This is a redox reaction, as we have Al in two different states: as metallic $Al_{(s)}$ and as ionic $Al^{3+}_{(aq)}$, which result from the loss of an electron. In redox reaction there is always elements in the chemicals that lose electrons and chemicals winning electrons.

Redox means oxidation and reduction

By comparing the redox number of the same element in two different compounds we can figure out in what compound the element has lost or gained electrons. Look for example the case of $\underline{Cr^{VI}}_2O_7{}^{2-}$ and $\underline{Cr^{III}}_2O_3$. The same element in two different molecules has two different redox numbers. In the case of dichromate, the redox of Cr is VI, whereas in the case of chromium(III) oxide the redox of Cr is III. The larger the redox number the more oxidized is an element, and that means the element has lost electrons. The smaller the redox number the more reduced is the element and that means it has gained electrons. If we compare both case, we have that Cr in dichromate is oxidized–it lost electrons–and Cr in chromium(III) oxide is reduced–it gained electrons.

Redox numbers in chemical reactions

The goal is to identify the element that undergoes oxidation and reduction in a chemical reaction. We can reach this goal by using the half-reaction method. Every redox reaction is composed of two process, a reduction and the oxidation. These two processes can be separated into two half-reactions so that the combination of both half-reactions lead to the balanced redox. Let us work on a simple unbalanced redox reaction:

$$Al_{(s)} \quad + \quad Cu_{(aq)}{}^{+2} \quad \longrightarrow \quad Al_{(aq)}{}^{+3} \quad + \quad Cu_{(s)}$$

Solid $Al_{(s)}$ on the reactant side has zero redox number, whereas ionic $Al_{(aq)}{}^{+3}$ on the product side has redox number equal to III. Al has undergone oxidation as its redox number increases from zero to three. Al has lost three electrons. We can write the oxidation half-reaction:

$$Al_{(s)} \quad \longrightarrow \quad Al_{(aq)}{}^{+3} \quad + \quad 3\,e^- \text{ (oxidation)}$$

Mind that electrons have negative charge and we add electrons to compensate the charge of $Al_{(aq)}{}^{+3}$. Now let us compare the redox number of Cu. In the reactant side we have $Cu_{(aq)}{}^{+2}$ with redox of II. In the product side we have $Cu_{(s)}$ with zero redox number. Cu has undergone reduction as its redox number has decreases. This means it has gained electrons, in particular two electrons:

$$Cu_{(aq)}{}^{+2} \quad + \quad 2\,e^- \quad \longrightarrow \quad Cu_{(s)} \text{ (reduction)}$$

Balancing simple redox reactions

The goal here is to balance a redox chemical reaction by combining two half-reactions. In the example above the oxidation and reaction involve different number of electrons. Hence in order to be able to add both redox we need to time each half-reaction by a number so that the number of electrons cancel out. As the first reaction involved three electrons and the second two, we will do:

$$\begin{array}{lr}
2\cdot(\ Al_{(s)} \longrightarrow Al_{(aq)}{}^{+3} + 3\,e^-\ ) & \text{(oxidation)} \\
3\cdot(\ Cu_{(aq)}{}^{+2} + 2\,e^- \longrightarrow Cu_{(s)}) & \text{(reduction)} \\
\hline
2\,Al_{(s)} + 3\,Cu_{(aq)}{}^{+2} + \cancel{6\,e^-} \longrightarrow 2\,Al_{(aq)}{}^{+3} + 3\,Cu_{(s)} + \cancel{6\,e^-} \quad + & \text{(redox)}
\end{array}$$

The overall balanced redox equation is:

Oxidation: $-3e^-$

$$2\,\overset{0}{Al}\,(s) + 3\,\overset{II}{Cu}{}^{2+}\,(aq) \longrightarrow 2\,\overset{III}{Al}{}^{3+}\,(aq) + 3\,\overset{0}{Cu}\,(s)$$

Reduction: $+2e^-$

Balancing redox reactions in acidic medium

Redox reactions happen in either basic or acidic medium. Here we will go over how to balance redox reactions in acidic medium. In order to do this we will first separate the reaction in two half-reactions. In each semi-reaction we will balance all elements but hydrogen and oxygen. When all elements are balanced, we will proceed to balance O by adding H_2O molecules and we will balance H by adding H^+. Finally, we all add electrons to compensate the charge of the reaction. Let us work on an example:

$$MnO_{4\,(aq)}{}^- \quad + \quad Fe_{(aq)}{}^{2+} \quad \longrightarrow \quad Mn_{(aq)}{}^{2+} \quad + \quad Fe_{(aq)}{}^{3+}$$

One of the semi-reactions involve Manganese whereas the other involves Iron. The redox number of Mn in permanganate is VII hence Manganese is being reduced, as its redox number decreases from VII to II, whereas Iron is being oxidized as its redox number increases from II to III. The oxidation half-reaction does not contain hydrogen or oxygen hence we will only have to balance the charger with one electron:

$$Fe_{(aq)}{}^{2+} \quad \longrightarrow \quad Fe_{(aq)}{}^{3+} \quad + \quad e^- \ \text{(oxidation)}$$

The reduction half-reaction contains oxygen. Hence, we will have to add H_2O molecules to balance oxygen and H^+ to balance hydrogen. In particular, we will need two water molecules–as $MnO_4{}^-$ has four oxygens and we will have to add eight protons as we are adding four molecules of water. Finally, we need to add three electrons to equalize the charge:

$$MnO_{4\,(aq)}{}^- \quad + \quad 8\,H_{(aq)}{}^+ \quad + \quad 5\,e^- \quad \longrightarrow \quad Mn_{(aq)}{}^{2+} \quad + \quad 4\,H_2O_{(l)} \ \text{(reduction)}$$

As the oxidation involves one electron and the reduction three, we need to time the oxidation by three:

$$\begin{array}{lr}
5\cdot(\ Fe_{(aq)}{}^{2+} \longrightarrow Fe_{(aq)}{}^{3+} + e^-\ ) & \text{(oxidation)} \\
MnO_{4\,(aq)}{}^- + 8\,H_{(aq)}{}^+ + 5\,e^- \longrightarrow Mn_{(aq)}{}^{2+} + 4\,H_2O_{(l)} & \text{(reduction)} \\
\hline
5\,Fe_{(aq)}{}^{2+} + MnO_{4\,(aq)}{}^- + 8\,H_{(aq)}{}^+ + \cancel{5\,e^-} \longrightarrow 5\,Fe_{(aq)}{}^{3+} + Mn_{(aq)}{}^{2+} + 4\,H_2O_{(l)} + \cancel{5\,e^-} \quad + & \text{(redox)}
\end{array}$$

Oxidation: $-1e^-$

$$8\,H^+\,(aq) + \overset{VII}{Mn}O_4\,(aq) + 3\,\overset{II}{Fe}{}^{2+}\,(aq) \longrightarrow 3\,\overset{III}{Fe}{}^{3+}\,(aq) + \overset{II}{Mn}{}^{2+}\,(aq) + 4\,H_2O\,(l)$$

Reduction: $+5e^-$

Example

Balance the following redox in acidic medium:

$$Cr_2O_7{}^{2-} \quad + \quad SO_3{}^{2-} \quad \longrightarrow \quad Cr^{3+} \quad + \quad SO_4{}^{2-}$$

Answer: We locate the same element in both sides of the reaction with different redox number. We found Chromium in the form of dichromate ($Cr_2O_7{}^{2-}$) with redox number VI and Chromium in the product side with redox III. Therefore Chromium is being reduced, as its redox number decreases. Differently, we found sulphur in the reactants side in the form of sulfite ($SO_3{}^{2-}$) with redox number of IV and in the product side with redox number of VI. Therefore Sulphur is being oxidized. We will first set up the oxidation half-reaction knowing that we have different amounts of Cr in both side and that we will have to add water molecules to balance O and protons to balance H. As we have seven oxygens in $Cr_2O_7{}^{2-}$ we will have to add seven water molecules. Also, as we add seven water molecules we will have to add fourteen protons. We will need six electrons to compensate the charge:

$$Cr_2O_7{}^{2-} \quad + \quad 14\,H^+ \quad + \quad 6\,e^- \quad \longrightarrow \quad 2\,Cr^{3+} \quad + \quad 7\,H_2O \text{ (reduction)}$$

For the reduction half-reaction, we have a difference of one oxygen atoms and hence we will need one water molecule and two protons; we will need two electrons to compensate the charge:

$$SO_3{}^{2-} \quad + \quad H_2O \quad \longrightarrow \quad SO_4{}^{2-} \quad + \quad 2\,H^+ \quad + \quad 2\,e^- \text{ (oxidation)}$$

In order to add both half-reactions as the reduction involves two electrons and the oxidation six, we will have to multiply the reduction half-reaction by three:

$$1\cdot\left(\, Cr_2O_7{}^{2-} + 14\,H^+ + 6\,e^- \longrightarrow 2\,Cr^{3+} + 7\,H_2O \,\right) \qquad \text{(reduction)}$$
$$3\cdot\left(SO_3{}^{2-} + H_2O \longrightarrow SO_4{}^{2-} + 2\,H^+ + 2\,e^-\right) \qquad \text{(oxidation)}$$

$$+$$

$$Cr_2O_7{}^{2-} + 14\,H^+ + 3\,SO_3{}^{2-} + 3\,H_2O + \cancel{6\,e^-} \longrightarrow 2\,Cr^{3+} + 7\,H_2O + 3\,SO_4{}^{2-} + 6\,H^+ + \cancel{6\,e^-} \qquad \text{(redox)}$$

After we cancel the electrons and eliminate protons and water molecules, the balanced reaction is:

Oxidation:−2e⁻

$$\overset{\text{VI}}{Cr_2}O_7^{2-}\,(aq) + 8\,H^+\,(aq) + 3\,\overset{\text{IV}}{S}O_3^{2-}\,(aq) \longrightarrow 3\,\overset{\text{VI}}{S}O_4^{2-}\,(aq) + 2\,\overset{\text{III}}{Cr^{3+}}\,(aq) + 7\,H_2O\,(l)$$

Reduction:+3e⁻

Balancing redox reactions in basic medium

In order to balance a redox in basic medium we need first to balance the reaction in acidic medium. After, we will compensate all H^+ with OH^- in both sides of the reaction. Mind that when combining H^+ with OH^- we obtain H_2O. For example, in order to balance the following reaction in basic medium:

$$3\,Fe_{(aq)}{}^{2+} \quad + \quad MnO_{4\,(aq)}{}^{-} \quad + \quad 4\,H_{(aq)}{}^{+} \quad \longrightarrow \quad 3\,Fe_{(aq)}{}^{3+} \quad + \quad Mn_{(aq)}{}^{2+} \quad + \quad 2\,H_2O_{(l)}$$

we will add four OH^- in both sides:

$$3\,Fe_{(aq)}{}^{2+} \quad + \quad MnO_{4\,(aq)}{}^{-} \quad + \quad 4\,H_{(aq)}{}^{+} \quad + \quad 4\,OH_{(aq)}{}^{-} \quad \longrightarrow \quad 3\,Fe_{(aq)}{}^{3+} \quad + \quad Mn_{(aq)}{}^{2+} \quad + \quad 2\,H_2O_{(l)} \quad + \quad 4\,OH_{(aq)}{}^{-}$$

And after cancelling the four protons with the four hydroxyls, we have:

$$3\,Fe_{(aq)}{}^{2+} \quad + \quad MnO_{4\,(aq)}{}^{-} \quad + \quad 4\,H_2O_{(l)} \quad \longrightarrow \quad 3\,Fe_{(aq)}{}^{3+} \quad + \quad Mn_{(aq)}{}^{2+} \quad + \quad 2\,H_2O_{(l)} \quad + \quad 4\,OH_{(aq)}{}^{-}$$

Now we have four water molecules in the left and two in the right. We will cancel two molecules from each side in order not to report water molecules twice:

$$3\,Fe_{(aq)}{}^{2+} \quad + \quad MnO_{4\,(aq)}{}^{-} \quad + \quad 2\,H_2O_{(l)} \quad \longrightarrow \quad 3\,Fe_{(aq)}{}^{3+} \quad + \quad Mn_{(aq)}{}^{2+} \quad + \quad 4\,OH_{(aq)}{}^{-}$$

STUDENT INFO

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Worksheet

Redox balancing

1. Balance the following redox reaction indicating the oxidation semi reaction, the reduction semis reaction and the number of electrons exchanged:

Redox reaction		
$Fe_{(s)} + O_{2\,(g)} \longrightarrow Fe^{3+}_{(aq)} + O^{2-}_{(aq)}$		
Oxidation semi-reaction		$n^{R}_{e^-} =$
Reduction semi-reaction		$n^{O}_{e^-} =$
Total reaction		$n^{Tot}_{e^-} =$

Redox reaction		
$Zn_{(s)} + NO_3{}^-{}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + NO_{2\,(g)}$ (acidic medium)		
Oxidation semi-reaction		$n^{R}_{e^-} =$
Reduction semi-reaction		$n^{O}_{e^-} =$
Total reaction		$n^{Tot}_{e^-} =$

Redox reaction		
$MnO_4{}^-{}_{(aq)} + SO_3{}^{-2}{}_{(aq)} \longrightarrow MnO_{2\,(s)} + SO_4{}^{-2}{}_{(aq)}$ (basic medium)		
Oxidation semi-reaction		$n^{R}_{e^-} =$
Reduction semi-reaction		$n^{O}_{e^-} =$
Total reaction		$n^{Tot}_{e^-} =$

EXPERIMENT

Types and color of radiation

Goal

The goal of this experiment is to learn how to classify radiation in terms of its frequency and color.

Types and color of radiation

Depending on its frequency–or on its wavelength–radiation can be classified as gamma rays, x rays, ultraviolet (UV), visible, infrared (IR), microwaves or radio waves. For example, radiation with wavelength of 10^{-2} nm belongs to gamma rays radiation, whereas radiation with wavelength of 10^{4} nm belongs to the Infrared. Gamma rays are the most energetic type of radiation, whereas radio waves are the less energetic waves. At the same time, radio aves have the largest wavelength. Light not always have color. Only a small range of wavelengths belong to visible radiation and the visible spectrum correspond to the set of visible frequencies. This means you will not be able to see for example, IR radiation or gamma rays. The color of the radiation is also dependent on the wavelength–of the frequency as both are related–and for example $\lambda = 450$ nm will be blue light. Ultraviolet radiation is the most energetic visible radiation whereas infra red waves are the less energetic waves of the visible spectrum.

Table 5 Types and color of radiation

Type of radiation	ν (Hz)	Color of radiation	λ (nm)
Gamma	$>3 \times 10^{19}$	Violet	380-450
X-rays	$3 \times 10^{19} - 3 \times 10^{16}$	Blue	450-485
UV	$3 \times 10^{16} - 8 \times 10^{14}$	Cyan	485-500
IR	$4 \times 10^{14} - 4 \times 10^{11}$	Green	500-565
MicroW	$3 \times 10^{11} - 3 \times 10^{8}$	Yellow	565-590
RadioW	$3 \times 10^{8} - 3 \times 10^{3}$	Orange	590-625
		Red	625-740

Remember that the energy, frequency and wavelength of a radiation are related by means of the following equations:

$$E = h\nu \qquad c = \nu \cdot \lambda \qquad E = \frac{hc}{\lambda}$$

where c is the speed of light in the vacuum, 3×10^{8} m/s and $h = 6.6 \times 10^{-34}$Js is the Plank's constant. E is the energy in Joules, ν is the frequency in Hz and λ is the wavelength in m. Using wavelength values in nm, we can use the following set of formulas:

$$E = 6.6 \times 10^{-34}\nu \qquad 3 \times 10^{17} = \nu \cdot \lambda \qquad E = \frac{1.98 \times 10^{-16}}{\lambda}$$

STUDENT INFO

Name: ______________________ Date: __________

Worksheet

Types and color of radiation

1. Use the relationship between the different properties of light in order to classify the following list of radiations in terms of its type as: gamma rays, x-rays, UV, IR, Microwaves, or radiowaves.

Type of radiation			
E(J)	λ (nm)	ν (Hz)	Type of radiation
		5×10^{10}	
		2×10^{12}	
		8×10^{15}	
	1×10^{5}		
	5×10^{3}		
	6×10^{-4}		
3×10^{-28}			
3×10^{-23}			

2. Use the relationship between the different properties of light in order to classify the following list of radiations in terms of its color as: violet, blue, cyan, green, yellow, orange, red. Radiation outside the visible spectra is colorless.

Color of radiation			
E(J)	λ (nm)	ν (Hz)	Radiation color
		5×10^{14}	
		7×10^{14}	
		4×10^{14}	
	700		
	550		
	400		
4×10^{-19}			
3×10^{-19}			

EXPERIMENT

The photoelectric effect

Goal

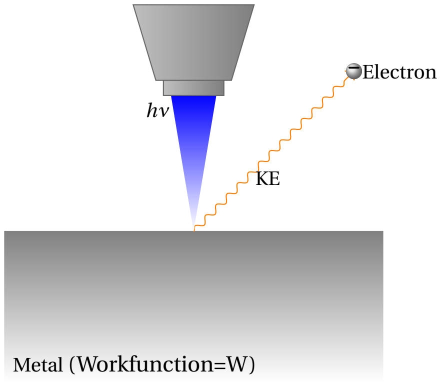
The goal of this experiment is to familiarize with the calculations involved in the photoelectric effect.

The photoelectric effect

The energy values for the different energy levels in J is a very small number (e.g. $-2.178\times10^{-18}J$). Sometimes, it is convenient to use another energy unit, called electron-volt, that makes these values have more reasonable values.

$$\boxed{1eV = 1.60218\times10^{-19}J} \quad \text{or} \quad \boxed{\frac{1eV}{1.60218\times10^{-19}J}}$$

For example, the energy in eV of the first level is $E_1 = -13.6eV$, whereas the energy of the third level is $E_3 = -1.5J$. The photoelectric effect was a mysterious phenomena discovered early in the twentieth century. Scientist found that if you expose a metal to light, under certain conditions, it emits electrons–it produced electricity. They found that the intensity–the brightness–of the radiation was not a key component of this phenomena, and not by increasing the intensity you were able to produce electrons.



They key was the frequency of the radiation. For frequencies above a specific threshold radiation was produce. If the frequency was above that threshold–called threshold radiation, ν_c–then the larger the intensity of the radiation the more electrons were produced. On that time, the current theory of light, associating intensity of light with energy, was unable to explain this phenomena. Albert Einstein used Plank's theory of the blackbody radiation to solve this mystery. He assumed that light is made of a steam of particles called photons, each with a given energy, $h\nu$. The electrons of a metal are held by attractive forces. A property called work function W–or biding energy–tells how strongly the electrons of a metal are held together. Metals such as Fe have high workfunctions in comparison to metals such as Ca, that means it takes more energy to remove an electron from the metal. Therefore, if the energy of the radiation was enough to overcome this forces, the electron emission will happen. In other words, if $h\nu$ is larger than W the electron emission will happen and the electrons emitted will have a kinetic energy (KE) equal to:

$$\boxed{KE = 4.13\times10^{-15}\nu - W} \quad \text{Photoelectric effect}$$

where:

KE is the kinetic energy of the emitted electrons in eV

4.13×10^{-15} is called Plank's constant in eV· s, h

v is frequency in hertz (Hz)

is the workfunction of the metal in eV

What is the threshold frequency, v_c? When the energy of the radiation is the same as the work-function of the metal, then electrons are ejected. This way, the threshold frequency is just the work-function of the metal converted in units of frequency:

$$\boxed{v_c = \frac{W}{4.13 \times 10^{-15}}} \quad \text{threshold frequency}$$

4.13×10^{-15} is called Plank's constant in eV· s, h, v_c is the threshold frequency in hertz (Hz) and W is the workfunction of the metal in eV. Einsteins theory of the photoelectric effect shocked the scientific community. Before this theory, light was considered a wave. Based on Einstein's theory, wave poses properties of both a particle and a wave, and depending on the experiment one experience light as a wave or as a particle.

Example

A metal with workfunction of 5eV is exposed to a radiation source with frequency of 2×10^{15} *Hz. Indicate whether electrons will be ejected and if so, indicate the kinetic energy of these.*

Answer: Using the photoelectric effect equation we have that a radiation of 10^{15}Hz frequency has an energy of

$$4.13 \times 10^{-15} \cdot 2 \times 10^15 = 8.27eV$$

As this values is larger than the workfunction of the metal (W=5eV), therefore electrons will be ejected with a kinetic energy of:

$$KE = 8.27 - 5 = 3.27eV$$

STUDENT INFO

Name: ______________________ Date: __________

Worksheet

The photoelectric effect

1. Obtain the threshold frequency for the given metals to produce the photoelectric effect.

Threshold frequency					
Metal	W(eV)	ν_c (Hz)	Metal	W(eV)	ν_c (Hz)
Ag	4.64		Mn	4.10	
Ba	2.52		Fe	4.67	
Al	4.20		Ca	2.87	

2. For the following different metals, indicate whether the radiation would induce electron ejection. If so, indicate the kinetic energy of the ejected electron.

Threshold frequency				
Metal	W(eV)	ν (Hz)	Photoelectric Effect? yes/no	KE(eV)
Ag	4.64	1×10^{16}		
Ag	4.64	2×10^{14}		
Al	4.20	1×10^{16}		
Al	4.20	1×10^{14}		
Ca	2.87	7×10^{14}		
Ca	2.87	7×10^{15}		

EXPERIMENT

Bohr's model of hydrogen

Goal

The goal of this experiment is to get familiar with the calculations involved Bohr's model, in particular, the calculation of energy levels and energy transitions.

The Bohr model

The Bohr model explains the electronic structure of hydrogen, in particular the spectrum of hydrogen and the position of the different energy lines. This model is based on the idea that the electron of hydrogen moves around the nucleus only in certain allowed circular orbits. Each orbit is called energy level, being characterized by an energy E_n and an integer number n. The following formula gives you the energy value for each level:

$$\boxed{E_n = -2.178 \times 10^{-18} \frac{1}{n^2}} \quad \text{Bohr formula in J}$$

where:

E_n is the energy of the level n in joules

n is the number of the level

$-2.178 \times 10^{-18} J = R_H$ is called the Rydberg constant $(-13.59 eV)$

For example, the energy of the first level is $E_1 = -2.178 \times 10^{-18} J$, whereas the energy of the third level is $E_3 = -5.44 \times 10^{-19} J$; the higher the n the larger–more positive–is the energy of the level.

Energy levels of hydrogen

Let's gain deeper insight into the idea of an energy level. These are just numbers that represent the location–in energy units– of the different states that an electron can occupy in an hydrogen atom. The first level is E_1 and is the most negative energy value, being also the most stable level. In another words, the electrons in this level are tightly bonded to the nucleus. This energy levels is called the fundamental energy level. The following levels $(E_2, E_3, \cdots, E_n)$ have a more positive energy. For example, comparing E_2 and E_4, we have that an electron on the level number four $(E_4 = -0.85 eV)$ is less stable than on level two $(E_2 = -3.40 eV)$. Hence it would be easier to remove an electron from level number four than from level two. For small n values the levels are spread from each other. However, when n increases, the energy levels are more and more closer to each other. Electrons occupying energy levels with n higher than one are called excited states. Finally, there are infinite number of levels and the highest energy (E_∞) level has an energy of 0J. The electron transitions between the different energy levels is what generates the emission spectrum of hydrogen.

Transition energies

Bohr's models is able to explain the atomic spectrum of hydrogen. Each line in the spectrum represents a transition between two levels of energy. For example the line at 102nm represents the transition of an electron between the level three and the level one, we call this $3 \rightarrow 1$. The atomic spectrum of hydrogen is obtained by means of exciting hydrogen atoms with energy, so that the electron jumps from a lower level into a higher level. When the atom relaxes, it emits light as the electrons move back from high every levels into lower–more stable–levels. This is called an emission. The different possible emissions have names based on the person who discovered then. For example, the set of emissions that end up in the fundamental level (n=1) are called the Lyman series. Similarly, the set of emissions that end up in the level (n=2) are called the Balmer series. For example: 3 → 2, 4 → 2 or 5 → 2. Only the Balmer series correspond to emissions on the visible spectra. The Lyman series belong to the ultraviolet and the Paschen and Bracket series belong to the infrared. Bohr's formula gives you the values of the energy levels. If we subtract the energy values for two energy levels, we obtain the energy for a transition. The energy for an electron transition between two energy levels, from n_1 to n_1 is given by:

$$\boxed{\Delta E_{n_2\to n_1} = -2.178\times 10^{-18}\left(\frac{1}{n_2^2}-\frac{1}{n_1^2}\right)} \quad \text{Energy transition formula}$$

where:

$\Delta E_{n_2\to n_1}$ is the energy in joules for the transition, the line in the spectra

n_2 is the number of the final level

n_1 is the number of the initial level

$-2.178\times 10^{-18}J = R_H$ is called the Rydberg ($-13.59eV = R_H$)

Example

Calculate the following transition energies:

(a) $\Delta E_{4\to 3}$ *in J*

(b) $\Delta E_{4\to 3}$ *in eV*

(c) Calculate the final energy level for a transition with energy 1.34eV knowing the first energy level involved in the transition is $n = 3$

Answer: (a) We will use the energy transition formula to calculate the energy needed to move one electron from $n_1 = 4$ to $n_2 = 3$:

$$\Delta E_{n_2\to n_1} = -13.59\left(\frac{1}{n_2^2}-\frac{1}{n_1^2}\right) = -13.59\left(\frac{1}{3^2}-\frac{1}{4^2}\right)$$
$$= -13.59\Big(0.111-0.0625\Big) = -0.66eV$$

The transition energy is negative, this means the atom releases energy when transitioning between these two levels.
(b) We will use the energy transition formula this time in eV:

$$\Delta E_{n_2\to n_1} = -2.178\times 10^{-18}\left(\frac{1}{n_2^2}-\frac{1}{n_1^2}\right) = -2.178\times 10^{-18}\left(\frac{1}{3^2}-\frac{1}{4^2}\right)$$
$$= -2.178\times 10^{-18}\Big(0.111-0.0625\Big) = -1.058\times 10^{-19}J$$

(c) In this case, we know $\Delta E_{n_2\to 3}$ and we know the initial level is $n_1 = 3$. We can certainly solve for n_2:

$$1.34 = -13.59\left(\frac{1}{n_2^2}-\frac{1}{3^2}\right) = -13.59\left(\frac{1}{n_2^2}-\frac{1}{9}\right)$$

Solving for n_2 we have: $n_2 = 9$. Mind you need to square root n_2^2 to get the final value of n_2.

STUDENT INFO

Name: ______________________ Date: __________

Worksheet

Bohr's model of hydrogen

1. Calculate the energy levels for the following n numbers:

Energy levels		
n	E_n, J	E_n, eV
1		
2		
3		
4		
7		
10		
20		
∞		

2. Calculate the energy transition for the following pairs of n values:

Energy levels			
n_i	n_f	$\Delta E_{n_i \to n_f}$, J	$\Delta E_{n_i \to n_f}$, eV
1	2		
1	3		
1	4		
4	1		
2	3		
2	4		
2	5		
9	10		

EXPERIMENT

Quantum numbers

Goal

The goal of this experiment is to get familiar with the rules that relate the four different quantum numbers.

Quantum mechanics and electronic structure

Bohr's model was simplistic but still, it was able to correctly predict the fine structure of hydrogen–the atomic spectrum of hydrogen–and the energy transitions. The downside of this model resulted from considering that the electron moves in different orbits and that it did not provide a complete description of the behavior of the electrons in atoms. A correct assumption of the model was that the energy levels of the atom were quantized–electrons can only exist in specific energy levels characterized by the number n and not in a continuum of energy. This section will cover a more realistic theory that describe the structure of the atom: quantum mechanics. The outcome of this section is the existence of orbitals and quantum numbers.

Quantized energy and continuum energy

Quantum mechanics is a theory shared in physics and chemistry that describe the nature of matter. It is used to do modeling in chemistry–to model molecules and atoms. Think of an engineer designing a plane. Before building and selling the plane engineers carry computer-based modeling to ensure the plane will work properly. In a similar way, chemists carry modeling to describe the properties of chemicals. The theory behind engineering modeling is classical mechanics, based in Newton's law.The theory behind chemistry modeling is quantum mechanics, based in Schrödinger equation. Classical mechanics is based on the idea that the energy of a system, a plane or a car, is continuum, that is, the car or a plane can have any possible energy, starting from zero to any number you can think of. Differently, quantum mechanics is based in the idea that the energy of a system, an atom or a molecule, is quantized, that is it can only be certain specific values.

The Schrödinger equation

The Schrödinger equation is the fundamental equation in quantum mechanics. It was formulated by Erwin Schrödinger in 1926, an Austrian physicist, and it was based on complex mathematical techniques. This equation incorporates the electron behavior in terms of particle through its mass and in terms of wave through a wave-like function, Ψ.

$$\hat{\mathrm{H}}\Psi = \mathrm{E}\Psi$$

$\hat{H}$ is called hamiltonian operator and contains a kinetic energy and potential energy component. Ψ, the wave function depends on the location of the space of each electron in the system. E are the energy levels. The solution of this equation required advanced calculus and will not be covered here. In contrast to simple algebraic equation in which the answer is a simple number, the result of solving the Schrödinger equation is a set of functions and a number for each of these functions. The functions are called wavefunctions ($\{\Psi_1, \Psi_2, ..., \Psi_n\}$) and a for each wavefunction there is a energy value, leading to a set of energy levels ($\{E_1, E_2, ..., E_n\}$). The introduction of The Schrödinger's equation mark the beginning of a new field in physics and chemistry called quantum mechanics or wave mechanics. Quantum mechanics do not allow us to specify the location of an electron in an atom, however it allows to identify the are in the space where the electron is most likely to be located at a given time. Electron density refers to the probability that an electron is found in a particular region of an atom.

The wave function: orbitals

In quantum mechanics the wave function Ψ of an atom or a molecule (a system) is a complex function–it has an imaginary component–that contains all information of the system. By means of this function, we can simulate the behavior of the system and extract its properties. You want to think of Ψ as a box that contains information, in particular all information of the system you want to simulate. When expressed in spherical coordinates, Ψ depends on the position (r) and two angles (θ and φ) of each electron of the system. This function Ψ *per se* has no real meaning or interpretation. Differently, Ψ^2 its square value has a real physical interpretation, representing the probability of finding an electron near a particular point in space. An orbital is a single-electron wave function. In another words is a wave function that contains information of a single electron. The square value of an orbital represents the probability of finding an electron at a specific location. Electrons are very different than larger objects such as for example a tennis ball. Larger objects are localized, that is they are located at a specific point in space. Differently, electrons are delocalized, that means they are not located at a single point in space and time, and therefore we can only guess the probability of finding the electron an a specific point. Orbitals are mathematical functions with a radian and an angular component.

$$\Psi(r,\theta,\varphi) = \psi_n(r) \cdot Y_{\ell}^{m_\ell}(\theta,\varphi)$$

The radial part $\psi_n(r)$(r) describes how the orbital changes with the distance (r) along the space, whereas the angular part–these are called spherical harmonics, $Y_{\ell}^{m_\ell}$–describes the symmetry of the orbital. An a note, the term orbital is different than the term orbit, which is used in classical physics to describe the trajectory of an object (e.g. a planet) and was used in Bohr's model to describe the different states of movement of the electrons in the hydrogen atom. When we say that an electron is at a given orbital, we mean that the electron density distribution is described by the square of the wavefunction associated with that orbital.

Orbitals are described by three quantum numbers

In Bohr's model of the hydrogen atom, just a single number n was necessary to describe the electron state. Differently, in quantum mechanics, each orbitals Ψ_{n,ℓ,m_ℓ} is characterized by a three quantum numbers. These numbers differentiate each orbital from the rest. They are called: the principal quantum number n, the angular quantum number ℓ and the magnetic quantum number m_ℓ. These quantum numbers are interrelated and not all combinations are allowed. We will describe this in the following, but just as an example a possible combination would be $\Psi_{1,0,0}$ and $\Psi_{1,1,0}$ would be an impossible combination. In the following, we will describe each quantum number separately. The principal quantum number n is related to the size of the orbital. The larger this value, the larger the orbital and hence an electron in the orbital would have a greater average distance from the nucleus. n can only be integral values such as 1, 2 or 4 and it cannot be zero. The second quantum number ℓ is called *angular quantum number* and it describes the shape of the orbital. The values of ℓ are dictated by the value of n. In particular, ℓ goes from 0 until $n-1$. For example if $n = 3$, therefore l can be: $0, 1$or 2. As such, n and ℓ can never be the same value, that is the reason why $\Psi_{1,1,0}$ does not represent a good orbital. The third quantum number m_ℓ is called *magnetic quantum number* and is vaguely related to the orientation of the orbital. The values of m_ℓ depend on the value of ℓ. In general m_ℓ varies in a range of numbers indulging zero, $-l$, $-l-1$, ..., 0, ..., $l+1$, l. For example, if $\ell = 3$, m_ℓ can be any of these values: -3, -2, -1, 0, 1, 2, or 3. Three quantum numbers are necessary to describe each orbital. Still, a fourth quantum number, the spin number m_s was necessary to understand why magnetic fields affect the emission spectra lines of hydrogen and sodium. m_s is called *spin quantum number* and can only be either $+^1/_2$ or $-^1/_2$.

Example

Indicate if the following combination of quantum numbers is allowed: (a) n=1, ℓ=1, m_ℓ=0, and m_s=$+^1/_2$; (b) n=2, ℓ=0, m_ℓ=0, and m_s=$+^1/_2$; (c) n=3, ℓ=3, m_ℓ=-1, and m_s=$-^1/_2$

Answer: The four quantum numbers are not independent. The quantum number n is related to the quantum number l and the number l is related to m_ℓ. The only quantum number that is independent is the spin, m_s which can be $+^1/_2$ or $-^1/_2$. The first combination is not possible, as if $n = 1$, ℓ can only be $n-1$, that is zero. The second combination is correct, as if $n = 2$, the number ℓ can be: 0 or 1. At the same time if $\ell = 0$, then m_ℓ can also be zero. And finally, the spin value of $+^1/_2$ is allowed. The last combination is not allowed, as n and ℓ cannot be the same number.

STUDENT INFO

Name: ______________________ Date: __________

Worksheet

Quantum numbers

1. Indicate if the following quantum number combinations are allowed.

n	ℓ	m_ℓ	m_s	Allowed?
0	0	0	0	
1	0	0	0	
1	1	0	0	
1	-1	0	0	
1	1	1	0	
1	1	-1	0	
1	1	0	$+^1/_2$	
1	0	0	$+^1/_2$	
2	0	0	$-^1/_2$	
2	1	-1	$-^1/_2$	
2	1	1	$-^1/_2$	
2	2	0	$-^1/_2$	

EXPERIMENT

Electron configuration

Goal

The goal of this experiment is to learn how to write down electron configurations of atoms.

Electronic configuration of an atom

Atoms have in general many electrons. These electrons are arranged in the atom in a very specific way creating what we know as electronic structure. You want to think about the electronic configuration of an atom as a code that tells you the orbital location of each electron in the atom. In fact, there are two ways to present electronic configurations. One is called full electronic configuration (for example $1s^22s^1$) and the other one is called condensed electronic configuration (for example $[He]2s^1$). The full configuration display all orbitals in an atom, whereas the abbreviated only display the valence electrons–these electrons are less-tied to the nucleus–and the nobel gas core. At the same time, every orbital is characterized by a set of numbers–these are called quantum numbers. These numbers are not independent from each other and there are constrictions that relate the possible values of the quantum numbers. This section will show you how to construct electron configurations and how to extract quantum numbers from it.

Electron energy levels and sublevels

The electrons in an atom are arranged in different energy levels. Some levels have lower energy and the electron in these levels are close to the nucleus being also strongly attached to it, whereas other levels have higher energy and the electrons in these levels are less attached to the nucleus. Still, each electron in an energy level have the same energy. The energy levels are labeled with a number n that equals to a single number such as 1, 2, 3 and so on. The first energy level is $n = 1$ and never $n = 0$–think of this as an apartment in a building, the first floor is flour one. For example all electrons in level one $n = 1$ have the same energy. There is a limit to the number of electrons in an energy level and we call this occupancy. Only a few electrons can occupy the lower energy levels, while more electrons can be accommodated in higher energy levels. Level one can only fit two electrons, whereas level two can fit a total number of eight electrons. The maximum number of electrons allowed in any energy level is calculated using the formula

$$2n^2 \tag{8}$$

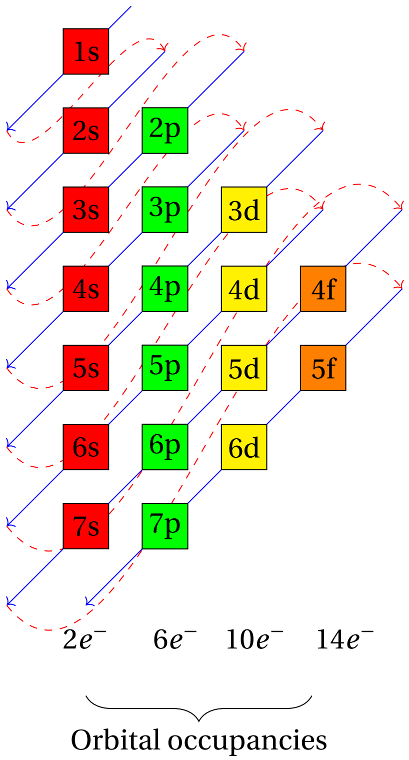
in which n is the energy level. You can see by using this formula that for example, the third level can accommodate 27 electrons. Each energy level consists of one or more sublevels, which contain electrons with identical energy. The number of sublevels in each level corresponds to n. For example, in the first energy level ($n = 1$) we have only one sublevel, whereas in the third energy level ($n = 3$) we have three sublevels.

The Pauli exclusion principle

The Pauli exclusion principle states that no two electrons in an atom can have the same values of the four quantum numbers. It tells that each electron is unique and it has to be differentiated with an unique combination of quantum numbers. Because of this, each orbital can only accommodate two different electrons, one with $-1/2$ spin and one with $+1/2$. Atoms in general contain numerous orbitals and each orbital should be filled with electrons. In every orbital you can fill only a maximum number of electrons. For example, in a s orbital you can place a maximum of two electrons. That is why you will find s^1 orbitals and s^2, with the latest being completely filled with electrons. In a p orbital you can place a maximum of six electrons and in a d orbital a maximum of ten. Finally, in a f orbital you can place fourteen or less electrons. For example, the orbital notation p^2 is correct as in p orbitals you can place six or less electrons. In this case, this orbital still have space to accept more electrons. Differently, the notation d^{12} is incorrect, as in d orbitals you can fit ten or less electrons and never twelve. The aufbau (build up) principle states that the electronic configuration of an atom can be obtained by adding one by one all electrons in the element. In order to fill the orbitals you should follow the Figure below. You start from the top of the table and follow the arrows that indicates the orbitals ordering. For examples the first orbital to be filled will be $1s$. After that you should fill $2s$ and $2p$. After that you should fill $3s$, $3p$, $4s$, $3d$, and $4p$. There is a maximum number of electrons that can occupy each orbital. An s orbital holds a maximum of 2 electrons. A p orbital takes up to 6 electrons, a d orbital can hold up to 10 electrons, and an f orbital holds a maximum of 14 electrons. An orbital can be completely filled with electrons, partially filled or empty. For example a $3s^1$ is half-filled with one electrons and $2p^6$ is completely filled. Another example, a $3d$ orbital is empty and can accommodate a maximum of 10 electrons.

Full electron Configuration

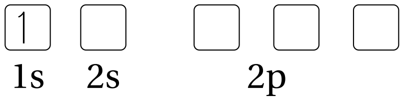
The full electron configuration of an atom is obtained by placing the total number of electrons of the atom in different orbitals with increasing energy. For example, the electron configuration for helium is written as $1s^2$ and the one for Li is $1s^2 2s^1$. First, how do we know the total number of electrons in an atom? That is the same as the atomic number and is indicated in the periodic table. Look for the element and the atomic number is on the top left side of the element. For example, the atomic number of hydrogen is one, and the number of electrons in He is two. Similarly, nitrogen has seven electrons. Second, how do we know what orbitals to fill? The Figure bellow shows the orbital order. You need to start from the top of the image, from orbital $1s$ and proceed to next arrow, starting from the end of the arrow. This way, after $1s$ goes $2s$ and then $2p$, $3s$, $3p$, $4s$, and $3d$. Mind that every s orbital can only fit two electrons, and p orbitals can fit six electrons, and so on. The following example will help you construct the electron configuration for a given atom.



Hund's rule

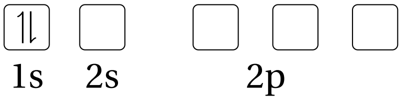
Let us build up the electron configuration of a series of atoms. Starting with hydrogen, with one electron, we have that only the $1s$ orbitals will be filled:

H $1s^1 2s 2p$



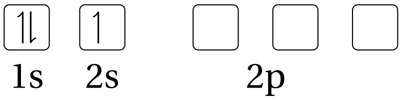
For the case of helium, with two electrons, we have that based on Pauli's principle both electron have different quantum numbers and in order to differentiate this we will represent the pair of electron with arrows in opposite direction. We say both electrons are paired:

He $1s^2 2s 2p$

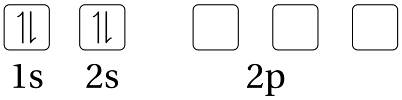


Now, for the case of Lithium and Beryllium we have a similar situation. Lithium has a single unpaired electron and Beryllium has a set of paired electrons in the $2s$ orbital:

Li $1s^2 2s^1 2p$

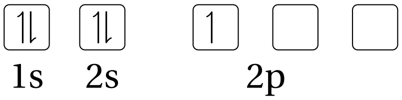


Be $1s^2 2s^2 2p$



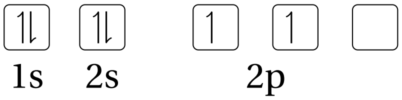
The next element, Boron, has one electron in a p orbitals. Ass all p orbitals are degenerate–they have the same energy–we can place that single electron in any of the p orbitals. Normally, we use the one on the left:

B $1s^2 2s^2 2p^2$



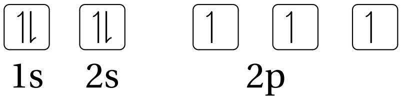
For the case of Carbon, as we need to place two electrons in the p orbitals, Hund's rule states that we need to place the electron maximizing the number of unpaired electrons. In another words, the second electron goes into a separate p orbital, just like below:

C $1s^2 2s^2 2p^2$



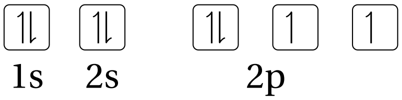
Similarly, for the case of Nitrogen we have:

N $1s^2 2s^2 2p^3$



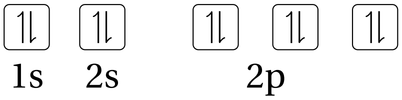
Now, for the case of oxygen, as we cannot place that extra electron in a separate p orbital we have to start pairing electrons:

O $1s^2 2s^2 2p^4$



Finally, for the case of Neon we have the whole second shell filled with electrons:

Ne $1s^2 2s^2 2p^6$



Example

Obtain the electronic configuration of C.

Answer: The atomic number of C is Z=6 and that means C has 6 electrons. The orbital order is: $1s$,$2s$, $2p$, $3s$, etc. Each s orbital can fit two electrons, whereas the occupancy of the p orbitals is six electrons. Hence the electronic configuration of C is: $1s^2 2s^2 2p^2$. The s orbitals are all filled, whereas the p orbital is only occupied with two electrons.

Abbreviated electron configuration

Let us compare the electronic configuration of Ni and Ar. We have:

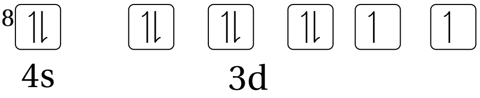
$$\text{Ni } 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$$
$$\text{Ar } 1s^2 2s^2 2p^6 3s^2 3p^6$$

We called these configuration the *full electronic configuration*. If you look carefully, you will see that both configurations resemble and in particular the configuration of Ni is the same as the configuration of Ar with two extra orbitals added. We say Ni has the core of Ar and 10 electrons on its valence. We can represent this fact by indicating Ar with brackets:

$$\text{Ni } 1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8 \qquad \text{(full electron configuration)}$$
$$\text{Ni } [Ar] 4s^2 3d^8 \qquad \text{(abbreviated electron configuration)}$$

The orbital diagrams are boxed diagrams indicating the valence electrons such as:

Ni [Ar] $4s^2 3d^8$



We call this last configuration as the *abbreviated electronic configuration*. You can figure out faster the abbreviated electronic configuration by looking for the noble gas on the table on the row above the element, and the period (row on the table) of the element. Ni is located in the period number four and the noble gas above this period is Ar. At the same time Ar has 18 electrons. That will give you the core $[Ar]$ with 18 electrons, and the remaining 10 electrons (Ni has 28 electrons) starting by the orbital $4s$, according to the period four. The electrons in the noble gas core are called *core electrons*, whereas the rest of the electrons are known as *valence electrons*. For the case of Ni, it has 18 core electrons in the Ar core and 10 electrons in the valence. Let us work another example: Cd. It has 48 electrons, and is located in group 5. The novel gas on the group above is Kr with 36 electrons. The core will be Kr–the novel gas on the period above–and we start right away filling $5s$ electrons–Cd belong to period five. In the valence electrons, we will place 12 electrons. Hence, the abbreviated configuration will be: $[Kr]5s^2 4d^{10}$.

Example

Obtain the full and abbreviated electronic configuration of Silver (Ag, Z=47) using the fact that the 4d and 5s orbitals switch order, being the d orbital filled first.

Answer: The atomic number of Ag is Z=47 and that means Ag has 47 electrons. The orbital order is: $1s$,$2s$, $2p$, $3s$, etc. Each s orbital can fit two electrons, whereas the occupancy of the p orbitals is six electrons, and d orbitals can fit 10 electrons. We will use the fact that the 4d orbital fills before the 5s. Hence the full electronic configuration of Ag is: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6 4d^{10} 5s^1$. As Kr is the noble gas on top with 36 electrons, the abbreviated electronic configuration of Ag is: $[Kr]4d^{10}5s^1$. Ag has 36 core electrons and 11 valence electrons.

rbitals is six electrons. Hence the electronic configuration of C is: $1s^2 2s^2 2p^2$. The s orbitals are all filled, whereas the p orbital is only occupied with two electrons.

STUDENT INFO

Name: ____________________ Date: __________

Worksheet

Electron configuration

1. Indicate the total number of electrons and the full electron configuration of the following atoms:

Atom	# Electrons	Full electron configuration
K		
Ga		
Si		
P		
Se		
Br		
Ar		

2. Indicate the total period, the number of valence of electrons and the abbreviated electron configuration of the following atoms:

Atom	Period	# Valence Electrons	Abbreviated electron configuration
Sc			
Nb			
Rh			
Hg			

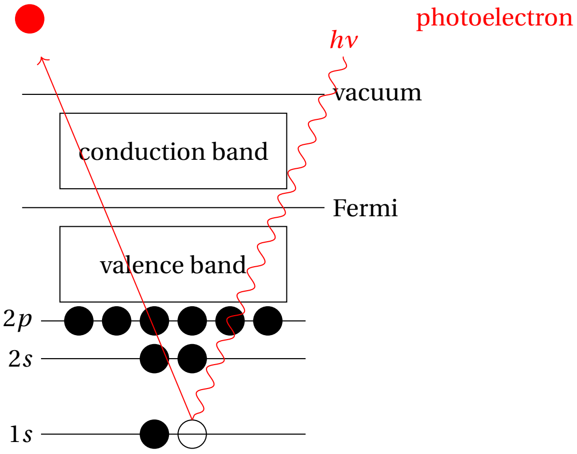
EXPERIMENT

Photoelectron spectra

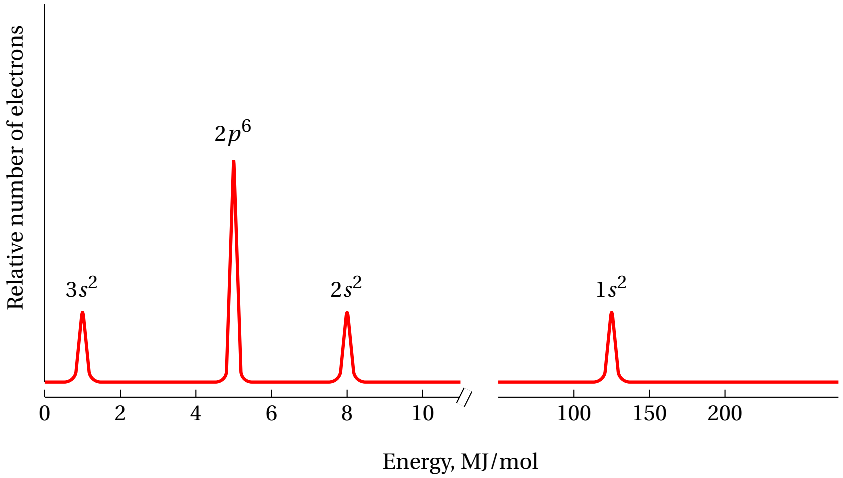
Goal

The goal of this experiment is to learn how to interpret photoelectron spectra.

Photoelectron spectroscopy of atoms



Photoelectron spectroscopy is a technique used to measure the relative binding of the electrons of an atom. It is useful to experimentally obtain the energy levels of an atom and properties such as the ionization energy. The principle behind the technique is similar to the photoelectric effect. By shinning light with enough energy one can remove electrons of an atom. By measuring the kinetic energy of the ejected electron and taking into account the energy of the radiation we can estimate the binding energy of the ejected electron–called photoelectron–in the atom. The name, photoelectron spectroscopy, refers to the fact that electrons are ejected with lights, by using photons. This technique is used to identify the presence of specific atoms as the photoelectron picks are element-specific, that is, each element present a series of characteristics peaks. However, there are many factors that affect the photoelectron picks and for example an 1s peak of H can appear at different energies depending on the nature of the atom bonded to hydrogen. In general, in the vertical axis of a PES plot you will find the relative intensity of the electrons and on the horizontal axis the energy of the photoelectron. The larger the energy the more strongly attached will be the electron in the atom and for example a 1s peck will appear at higher energy than a 2s peak. The higher the intensity of the pick the larger the number of electrons in the level resulting for the photoelectron. For example, the intensity of a $1s^2$ peak should be twice the intensity of a $1s$ peak. By comparing the intensity of the peaks and the relative location one can identify the atom represented on the spectra.



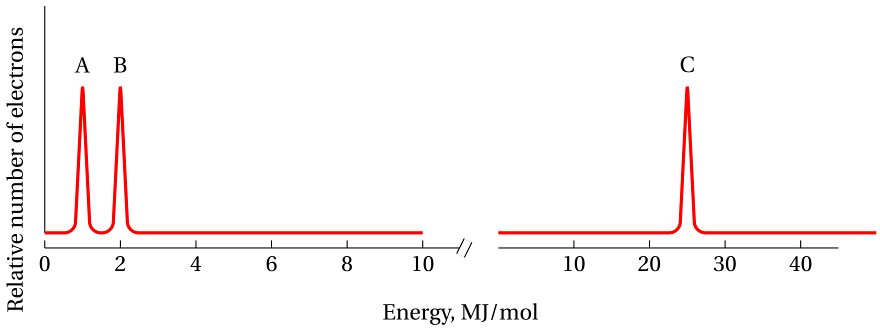
STUDENT INFO

Name: ______________________ Date: __________

Worksheet

Photoelectron spectra

1. For the following idealized photoelectron spectra of an atom, identify the atom represented and the nature of the different peaks



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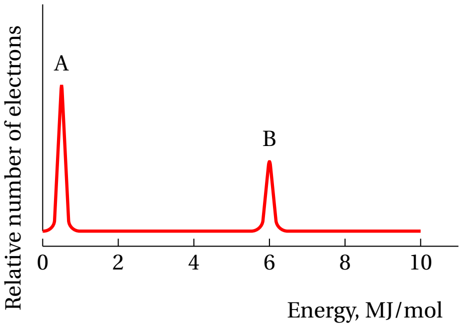
2. For the following idealized photoelectron spectra of an atom, identify the atom represented and the nature of the different peaks



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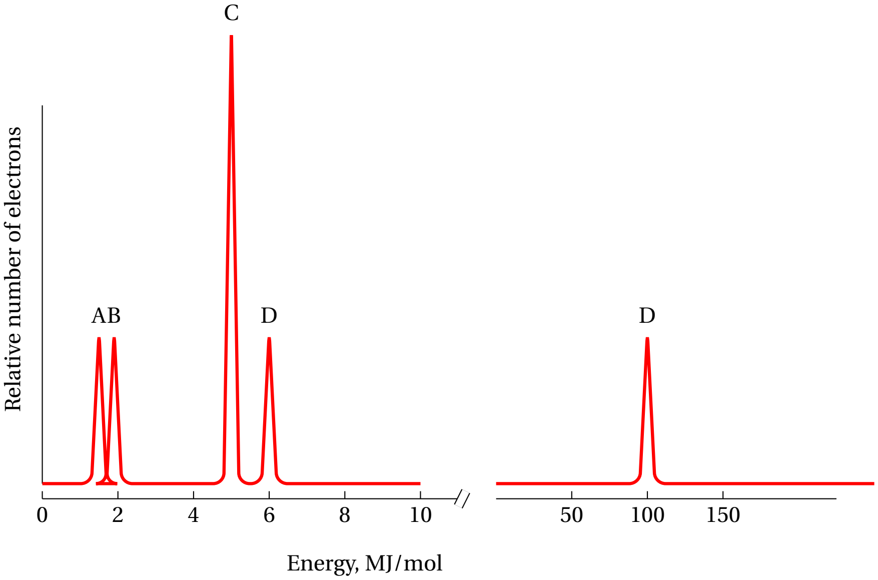
3. For the following idealized photoelectron spectra of an atom, identify the atom represented and the nature of the different peaks



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EXPERIMENT

Lewis Structures

Goal

This experiment will gain insight into lewis structures. On one hand, you will learn how to build up these structures by counting and pairing the valence electrons of the molecule. On the other hand, you will learn how to obtain the number of bonds and lone pairs from a generated lewis structure.

Materials

This is a theory-based experiment. You will not carry any lab work during this experiment and you only need a molecular models kit if available.

Background

Atoms are made of protons, neutrons, and electrons. Electrons–in particular valence electrons–are responsible for the main chemical properties of an atom. These electrons are tied weakly to the nucleus in comparison with the core electrons and hence they can be exchanged easily with other atoms. Atoms in a molecule, with a few exceptions such as the hydrogen case, will tend to be surrounded by eight electrons so that its electron configuration resembles a noble gas. This is known as the octet rule (the duet rule in the case of hydrogen), and this rule is responsible for the common negative charge of F and the positive charge of Na: F ($[He]2s^2 2p^5$) can easily receive an extra electron producing ionic F^- ($[He]2s^2 2p^6$=[Ne]) and atomic Na ($[Ne]3s^1$) can lose an electron producing ionic Na^+ ($[He]2s^2 2p^6 = [Ne]$). The electron-dot structure of an atom or a molecule is a visual representation of the electronic arrangement in an atom or a molecule.

Valence Electrons

The electrons of an atom can be divided in core electrons and valence electrons. The valence electrons of an atom are involved in chemical bonds as they are less bonded to the nucleus. The number of valence electrons in an atoms is the same as the group number. As an example, hydrogen H belong to the group IA and hence has one valence electrons. Similarly, oxygen O belongs to the group VIA and therefore it has six valence electrons.

Example

Indicate the number of valence electrons for the following atoms: N, O, C and S.

Answer: Nitrogen is in group VA and hence it has five valence electrons ($5e^-$). Oxygen belongs to the group VIA and C belong to IVA, hence they have

The Octet rule

Atoms gain or loose electrons when they combine to form molecules. The octet rule says that each atom in a molecule is surrounded by eight electrons. There are two important exceptions to this rule: hydrogen (H) will only surrounded by two electrons, and boron (B) by six.

Lewis dot symbol of an atom

In order to write the electron-dot structure of an atom, you just need to write down the symbol of the atom surrounded by the number of valence electrons located in the four directions of the space: top, bottom, right and left. In order to place the electrons, you start in any of the directions and fill one electron at a time. For example, for the case of three electrons we would have: $\cdot\dot{\mathrm{B}}\cdot$. After the four direction has been filled, you need to start pairing the electrons. For example, for the case of five electrons we would have: $\cdot\underset{\cdot\cdot}{\dot{\mathrm{P}}}:$. Another example, oxygen has six valence electrons and hence, the electron-dot structure would be $\cdot\underset{\cdot}{\ddot{\mathrm{O}}}:$ Similarly, for the case of fluorine the the electron-dot structure would be $:\underset{\cdot}{\ddot{\mathrm{F}}}:$. In the case of an ion, you need to add (if its an anion) or subtract (if its a cation) electrons, and for example the electron-dot structure of O^{2-} is $:\underset{\cdot\cdot}{\ddot{\mathrm{O}}}:^{2-}$

Lewis structure of diatomic molecules

This section will address how to build up electron-dot structures of diatomic molecules. These would be the most simple electron-dot structures of molecules that you will see. The first step is (a) to set up the atoms in the molecule in the form of a line. After that, (b) you need to count the total number of valence electrons in the molecule by adding the valence electrons of each atom (remember the number of valence electrons corresponds to the group number in the A notation, O is 6 as its group is VIA). Then, (c) compute the pairs of electrons–the total number of valence electrons divided by two; pairs of electrons are represented by lines. Finally, (d) you need to start distributing the pairs in the molecule in a very specific way: first connecting the atoms among themselves, and after placing the remaining pairs surrounding the atoms. Overall, each atom should be surrounded by four pairs with the exception of H and B.

Example

Construct the electron-dot structure of HCl.

Answer: We first arrange the atoms in the molecule as H Cl. Now we count the number of valence electrons: H(1) and Cl(7) that gives a total of eight electrons. Let us count the pairs of electrons; we have eight electrons and that is four pairs. Now we distribute the pair on each atoms knowing that each atom has to have four pairs with the exception of hydrogen that can only be surrounded by one pair. $\mathrm{H}:\underset{\cdot\cdot}{\ddot{\mathrm{Cl}}}:$, using lines instead of pairs (this is not necessary but makes the electron-dot structure look better) we obtain

$$\mathrm{H}-\overline{\underline{\mathrm{Cl}}}|$$

Lewis structures of polyatomic molecules

Now we will address how to build up electron-dot structures of more complex molecules. The first step is (a) to arrange the atoms in the molecule, in the form of a central atom and the remaining atoms around it; the central atom is the one with a lower index in the molecule (e.g. in H_2O is O or in NH_3 is N). After that, (b) you need to count the total number of valence electrons in the molecule, by adding the valence electrons of each atoms. Then you (c) calculate the pairs of electrons–the total number of valence electrons divided by two; pairs of electrons are represented by lines. In the following (d) you need to start distributing the pairs in the molecule in a very specific way (this is the key to building good electron-dot structures): first connecting the surrounding atoms to the central atom, after placing pairs on top of the surrounding atoms and finally by placing the remaining pairs in the central atom. Each atom should be surrounded by four pairs (this is the octet rule) with the exception of H and B as they do not follow the octet rule. When you have the final electron-dot structure, the pairs of electrons (or lines) that connect two atoms are called *bonds*, whereas the pairs not involved in an connection are called *lone pairs*. A very important note is that, at this point, is not that important the atom arrangement (if the molecule looks like a line, a triangle or so) as long as the connectivity (which atom goes in the center and in the surroundings) is correct.

Example

Construct the electron-dot structure of H_2O indicating the number of bonds and lone pairs.

Answer: we first arrange the atoms in the molecule as H O H. The central atom is O, as oxygen has the lower index in the H_2O molecule–the index for O is one and the index for H is two. Now we count the total number of valence electrons, including all atoms: 2xH($1e^-$) and O($6e^-$) that gives a total of eight electrons. Now we distribute the pair on each atoms knowing that each atom has to have 8 electrons with the exception of hydrogen that can only be surrounded by two.

$$\mathrm{H}:\underset{\cdot\cdot}{\ddot{\mathrm{O}}}:\mathrm{H}$$

and using lines instead of pairs (this is not necessary but makes the electron-dot structure look better) we obtain

$$\mathrm{H}-\underset{\cdot\cdot}{\ddot{\mathrm{O}}}-\mathrm{H}$$

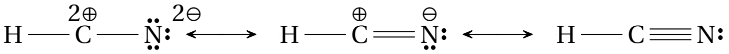
The molecule has two bonds, each one connecting a H to the oxygen atom, and two lone pairs located on the oxygen atom.

Atomic charges in a molecule and polyatomic ions

Often times you are going to encounter electron-dot structures like the ones below

$$:\mathrm{N}\equiv\mathrm{N}: \text{ and } :\mathrm{O}=\mathrm{O}:$$

in which the atoms are connected by means of multiple bonds, double or triple bond. Multiple bonds are formed while constructing electron-dot structures in order to minimize large atomic charges while preserving the octet rule. Look for example the lewis structure for the HCN molecule. The Lewis structure on the left has a positive center and a negative center. On this structure, carbon do not follow the octet rule and neither does nitrogen. We can minimize the large atomic charges while enforcing the octet rule by moving lone pairs from the atoms into the bond. For example, if we move one of the lone pairs of nitrogen into the C-N bond we have the Lewis structure in the center. In this structure above, even when the charges are now low, still carbon and nitrogen do not follow the octet rule as they are not surrounded by four pairs of electrons. We can solve this problems by moving one more time a lone pair from N to the C-N bond forming the Lewis structure on the right, the most stable structure as the atomic charges are small and the octet rule is preserved.



$$\mathrm{H}-\overset{2\oplus}{\mathrm{C}}-\underset{\cdot\cdot}{\ddot{\mathrm{N}}}:^{2\ominus} \longleftrightarrow \mathrm{H}-\overset{\oplus}{\mathrm{C}}=\underset{\cdot\cdot}{\overset{\ominus}{\mathrm{N}}}: \longleftrightarrow \mathrm{H}-\mathrm{C}\equiv\mathrm{N}:$$

As you can see in this example, multiple bonds arise in Lewis structures to minimize large atomic charges. All the structures above are a set of different Lewis structures of the same molecule, each with a different stability. These structure are referred to as resonant structures.

Calculations

① Calculate the number of valence electrons of the molecule. The number of valence electrons of an atom corresponds to its group number. For example, carbon has four valence electrons as it belongs to group 4A. Add all valence electrons of the atoms that compose the molecule.

② Calculate the number of pairs of electrons dividing the number of valence electrons of the molecule by two.

③ Draw the lewis structure of the molecule.

STUDENT INFO

Name: ____________________ Date: __________

Pre-lab Questions

Molecular Geometry

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1. Calculate the number of valence electrons for the following atoms and molecules:

N ____________________

O ____________________

P ____________________

Br ____________________

CO_2^{2-} ____________________

SO_4^{2-} ____________________

ICl_4^{-} ____________________

CH_2Cl_2 ____________________

2. Draw the lewis structure of the following compounds: H_2O, NO, HCl.

3. Draw the lewis structure of the following compounds and indicate their polarities: CH_4, CH_2Cl_2, $CHCl_3$.

STUDENT INFO

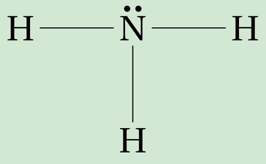
Name:

Date:

Results

EXPERIMENT

Molecular Geometry

Formula	# e^-	# pairs of e^-	Lewis Structure
	①	②	③
$\underline{N}H_3$	8	4	
$H_2\underline{S}$			
$\underline{Be}Cl_2$			

Formula	**# e^-**	**# pairs of e^-**	**Lewis Structure**
$\underline{P}Cl_3$			
$H\underline{C}O_2{}^-$			
$H_2\underline{S}$			
$\underline{Si}H_4$			

Formula	# e^-	# pairs of e^-	Lewis Structure
$O\underline{P}Cl_3$ violates octet rule			
$\underline{P}Cl_5$ violates octet rule			
$\underline{Al}Cl_6{}^{-3}$ violates octet rule			
$\underline{S}O_4{}^{-2}$ violates octet rule			

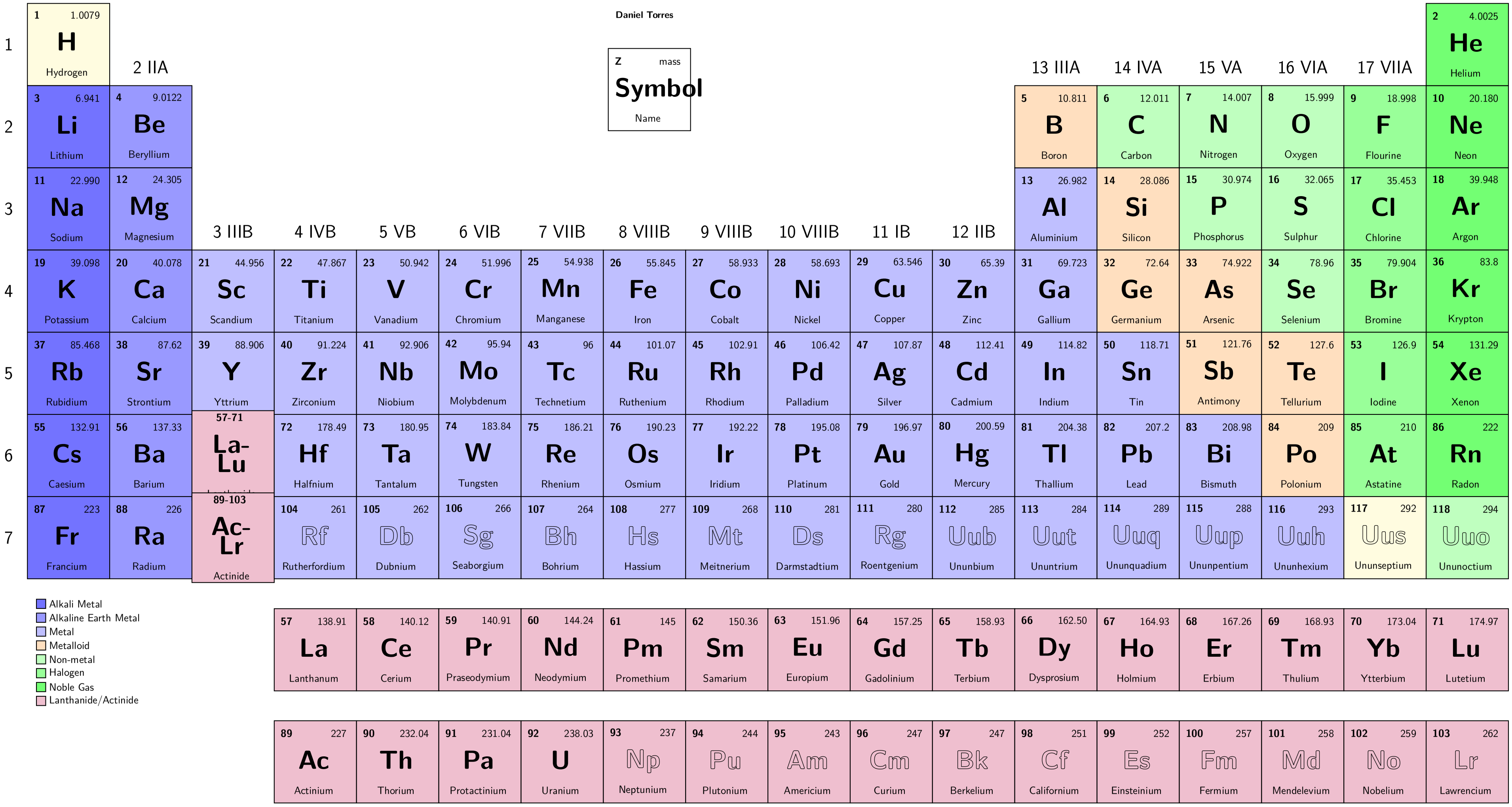
Formula	**# e^-**	**# pairs of e^-**	**Lewis Structure**
$\underline{\mathrm{Xe}}\mathrm{F}_2$ violates octet rule			
$\underline{\mathrm{S}}\mathrm{F}_6$ violates octet rule			
$\underline{\mathrm{Br}}\mathrm{F}_3$ violates octet rule			
$\underline{\mathrm{Se}}\mathrm{F}_4$ violates octet rule			

Formula	# e^-	# pairs of e^-	Lewis Structure
$\underline{N}O_2{}^-$			
$\underline{O}O_2$			
$\underline{S}O_3$			
$\underline{C}Cl_2O$			

Formula	# e^-	# pairs of e^-	Lewis Structure
$\underline{C}O_2$			
$\underline{S}Cl_2O_2$			
$\underline{S}H_2O$			
$\underline{C}N^-$			

Periodic Table of Chemical Elements

Daniel Torres



Z mass
Symbol Name

	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** 4.0025 **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** Flourine	**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** 65.39 **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** 83.8 **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** 95.94 **Mo** Molybdenum	**43** 96 **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** 127.6 **Te** Tellurium	**53** 126.9 **I** Iodine	**54** 131.29 **Xe** Xenon
6	**55** 132.91 **Cs** Caesium	**56** 137.33 **Ba** Barium	**57-71** **La-Lu**	**72** 178.49 **Hf** Halfnium	**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** 261 Rf Rutherfordium	**105** 262 Db Dubnium	**106** 266 Sg Seaborgium	**107** 264 Bh Bohrium	**108** 277 Hs Hassium	**109** 268 Mt Meitnerium	**110** 281 Ds Darmstadtium	**111** 280 Rg Roentgenium	**112** 285 Uub Ununbium	**113** 284 Uut Ununtrium	**114** 289 Uuq Ununquadium	**115** 288 Uup Ununpentium	**116** 293 Uuh Ununhexium	**117** 292 Uus Ununseptium	**118** 294 Uuo Ununoctium
				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** 173.04 **Yb** Ytterbium	**71** 174.97 **Lu** Lutetium
				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

- ■ Alkali Metal
- ■ Alkaline Earth Metal
- ■ Metal
- ■ Metalloid
- ■ Non-metal
- ■ Halogen
- ■ Noble Gas
- ■ Lanthanide/Actinide