24 Lessons That Rocked the World

Ian Guch

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By Ian Guch

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Introduction

Welcome to the second edition of this book. Just between you and me, this edition is way better than the first edition. Not only is my writing better, but the supplemental material has been beefed up to make it even better. You've chosen wisely, friend!

An explanation of how each chapter is laid out:

The following elements can be found in each chapter. There may be one or two things that are added or taken away in a particular chapter, but you can usually find each of the following:

A) Teacher resources that contain the following sections:

- Overview: A brief description of the topic being covered in the chapter.
- Teaching about [name of topic]: Hints about how you might best teach the topic of the chapter to your students. This section may contain anything from illustrative analogies to flow charts.
- Doing the [name of lab]: This section contains several subsections:
  - Equipment: This tells you everything you'll need to perform the lab.
  - Safety: These are the safety precautions you should follow while performing the lab.
  - Room destruction factor: This section explains the potential of each lab to cause a mess in your classroom. Though these labs generally have a low room destruction factor, some can be messier than others.
  - How the lab works: A quick discussion of how you should run the lab. It give suggestions for presenting the labs to the class, as well as addressing some of the areas that students sometimes have problems with.
  - What can go wrong: This section describes common sources of student error in the lab, and how to fix them.
  - Clean up: A complete description of what you need to do to properly dispose of the chemicals in each lab. Regulations may vary in your area. Always follow whatever disposal guidelines are mandated by your school district.
- Solutions to the lab: The solutions for the lab, with short descriptions.
- Solutions to the worksheet: The solutions for the worksheet, with short descriptions.

B) A handout which describes the topic for the students. Whether you decide to give your students these handouts or not, they do at least provide strategies you might want to try in presenting the material.

C) A large lab or activity: These activities generally take about an hour and are good ways of reviewing major chemical topics. Before doing each lab, review the safety issues in the teacher resource section. It's probably also a good idea to try each lab on your own to make sure your equipment is up to the job.

D) A worksheet to review the material: These worksheets have varied questions and review activities relevant to the chapter material. All are good in-class or homework activities.

Some words about copyright:

Please don't photocopy a lot of copies of this book and give them to your friends. Let's be honest – if you do, I'll probably never find out and send the copyright police after you. However, I will point out that I'm just a regular classroom teacher like you, except that I've put a tremendous amount of time and money into putting this book together. If you photocopy this book for someone else (or from someone else), you're not taking money from a big greedy company – you're taking it from an underpaid teacher like yourself.

To new teachers: The wisdom of Mr. Guch:

I've found that many of the teachers that visit my site have just started in the profession and have lots of questions about being a chemistry teacher. For those of you that may need some advice, here's some high-powered wisdom:

- Parents can be very irritating when they call on the phone and yell at you. Although it's annoying, it's also a sign that they support the educational system and their children. Try to remember that when they haul you in front of the school board for giving their kid a B+ instead of an A-.

- Never laugh when disciplining a student for doing something silly. It ruins the effect.

- No matter what happens, always give your students the appearance that you know what you're doing. Oddly enough, they will always believe it even if there's plenty of evidence to the contrary.

- Be nice to the secretaries and custodians at your school. In case you didn't already know, they're the people that really run the school.

- Act nice to your school security staff, because they spend their entire days dealing with the troublemakers at your school. They could use a smile.

- Drink a lot of coffee in the morning. That way, if you wake up in a bad mood you can explain it by saying, “Not enough coffee this morning.”

- If you’re sick, stay home. There’s nothing worse than having to lecture with a fever. On the other hand, teaching while sick would allow you to get the students back for making you sick in the first place.

- Quit smoking, because you’ll die of nicotine deprivation if you don’t. If your administrator asks you why you’re so grouchy (and the coffee explanation stops working), tell him or her that you’re quitting as a good example to the kids. You’ll be a hero!

- Keep your school and home email accounts separate. Otherwise, your brother will get email telling that his child needs to stay after school for cursing and a parent will get email about your recent trip to the proctologist.

- If you’re at a faculty meeting and an administrator presents anything said to be timesaving, vote against it. Trust me on this one.

The final word:

I’ve ranted long enough. I hope this book helps you with your lesson planning and teaching. I’ve also got free labs, worksheets, and activities online at my website (www.chemfiesta.com) in the “For Teachers” section, so stop by and get some free goodies. If you find mistakes in this book, feel free to email me (misterguch@chemfiesta.com) with them so I can fix them when the third edition comes out.

Ian Guch – September 17, 2004

Acknowledgements

Nobody can write a book without the help and support of those around them. I've been very lucky in that people have always jumped out of the woodwork to help out with this project, even though I don't have any money to pay them.

The biggest thanks go to my wife Ingrid. When I was writing the first edition of this book, I spent an entire summer wondering if this project would work, and when it did, I drove her crazy wondering if it would sell. In all that time, she stood behind the project and convinced me that it was worthwhile. For the second edition, she suffered another entire summer of rewriting and the same frustrations from the first edition. Not only didn't she divorce me, but she seems to like me just as much as when the project was originally started. She's great!

My parents have also helped a huge amount with both editions of this book. They edited this book until it was completely deconstructed and put back together in the right order. They also made me into the kind of person who has enough self-confidence to write a giant book, even though he knows it won't be a *New York Times* bestseller. They're swell!

My friend Nancy Levinger helped a lot with the first edition of this book, even though she was in the middle of getting tenure at Colorado State University. Nancy has been the biggest fan of my writing and teaching since I started with this whole ugly business, and I really appreciate her vote of confidence.

I'd also like to thank the whole crew at the Belle View, Virginia post office for answering my many questions about shipping and not getting mad when I mailed out 25 packages in the middle of the Christmas holiday. Here's a big shout out to Carl, Debbie, Bing, and Sonja. On a related note, I'd like to thank Gene the UPS guy for being such a good sport about constantly delivering very heavy shipments of books and packing materials to the house.

Thanks to LaTomya Glass at the U.S. Department of Energy for her help in securing the picture for the book cover. Not only was she friendly, but she sent me the photo twice when it was bent by a mail sorting machine.

Finally, I want to thank any musician, writer, or artist who's worked without the backing of a corporate sponsor. I probably wouldn't have even tried writing a book without the DIY example set by thousands of artists all over the world. Thanks, guys.

Chapter 1 – The Scientific Method

Overview

The scientific method is a series of six steps that scientists follow each time they solve a problem. These six steps include a **purpose** for investigating something, a **hypothesis** about how the phenomenon being investigated will behave when changes are made, a complete list of **materials** needed in the investigation, a complete **procedure** for the experiment, experimental **results**, and a **conclusion** in which it is determined whether the experiment worked. In this chapter, students will be using the scientific method to figure out the best way of making a tower out from toothpicks and sponges.

Teaching the Scientific Method

The scientific method is usually the first topic taught in a chemistry course. The reason is simple: Most students don't have a good feel for scientific inquiry and think that a scientist is an old bald guy with a lab coat who plays with steaming test tubes and flasks (and the odd monster or two). Unless students are taught early on that scientific inquiry is just a methodical variation of how they already solve problems, they'll cling to the idea that science is unapproachable and difficult to comprehend.

The best way to teach the scientific method is to do this lab without giving your students any prior instruction on the scientific method. It's only after the whole class has finished with the lab that you introduce the steps and details of the scientific method. When talking about the scientific method, I let the kids lead the discussion. Inevitably, they use non-chemical examples, making them more comfortable in a class that their friends and parents may have told them is really hard. Expect some difficulty starting the discussion, as students are sometimes reluctant to speak out the first week of school. You may want to use a canned starter topic, using as an example "Problem: My car broke down on the way to work" or something similar.

Another good reason to start your course with a scientific method lesson is that it teaches chemistry without ever mentioning chemicals. Some students are intimidated and confused if you start using complex chemical terminology the first day, so by starting with a lesson that uses familiar objects you'll make the kids more comfortable with the chemical world. That's when you hit 'em with the science.

In most classes there will be at least one student who gives examples of accidental scientific discoveries, pointing out that the scientific method isn't always the way people learn new things. The best known example is probably the discovery of penicillin. This question can be addressed by pointing out that while random luck contributes to a large number of scientific discoveries, a strong and consistent methodology allows you to reproduce and explore them.

Doing the Scientific Method Lab

Equipment:

Part 1:

- A large trash bag full of sponges cut into 1 x 1 x 1 centimeter cubes. Sponges can usually be purchased in bulk from "Dollar Stores", and are reusable for many years. To cut the cubes, use a paper cutter, as scissors take a very long time.
- Six boxes of round cocktail toothpicks
- Six meter sticks or measuring tapes

Part 2:

- Several large trash bags full of sponges cut into 1 x 1 x 1 centimeter cubes. The sponges in these bags should differ from the sponges used in Part 1. For example, if you used natural fiber sponges in Part 1, you should have a bag of plastic sponges, a different brand of natural fiber sponges, and a bag of sponges with scouring pads on one side.
- Six boxes of flat cocktail toothpicks
- Four packages of bamboo fondue skewers

Equipment note:
This lab is designed to be done using sponges, but other materials can also be used to good effect. Marshmallows make interesting towers and are inexpensive but tend to make an unbelievable mess. Gumdrops work well, but tend to make shorter towers than other materials. I've found that it's generally a good idea to stay away from food items, as they have a very short lifespan.

Safety:

There are no potential safety hazards in this lab and goggles are not required.

Room destruction factor:

Problems with this lab are minor. Natural fiber sponges sometimes give off small crumbs which find their way into the sink, but this is easily cleaned up with a wet paper towel. Make sure to stopper your sinks before starting this lab to ensure that the sponges don't lodge themselves in the pipes.

How the lab works:

In the first section of the lab, students will be building a structure out of sponges by sticking them together with toothpicks in whatever fashion seems most natural to them. In the second section, students will alter their construction method with the goal of making an even taller tower.

Students have no problems with the first part of this lab because they have an intuitive feel for how the tower should be built. Students typically start off with methodical building designs which degenerate into chaotic patterns as the tower grows. The frenzied pace of tower formation provides a great deal of fun as the students try to keep the towers from falling over.

The second section of the lab is a little more confusing for students, because they may not understand what you mean when you tell them to change one thing about their procedure. Make it clear to the students that if they change the type of sponge they use, they need to keep the other variables (type of toothpick and building design) the same. After all, if more than one variable is changed at a time, it becomes unclear which changed variable was responsible for the altered outcome. Once the students fully understand this, you'll find that tower construction will build back up to its former pace.

Make sure that you give an identical period of construction time for both parts of the lab. This ensures that any difference in tower height is a result of the variable they've changed and not a difference in allotted building time.

What can go wrong:

- The students make a tower that's 1.5 meters tall, but it collapses right before you measure it. You can eliminate any complaints about this by making it clear to your students that you will only measure the heights of the towers at the end of the allotted building time. This is a fair warning to them that if they have an impressively tall tower they should leave it alone until the end of the lab.
- The towers are smaller in the second trial than in the first. Point out to your students that when we make hypotheses, sometimes they don't work out like we hope. Regardless of outcome, experiments always yield useful information. If nothing else, your students have learned what not to do the next time they build a tower.
- Students sometimes impale themselves on toothpicks while trying to stick them into sponges. Keep a supply of bandages on hand.

Other things to keep in mind:

Make sure your students keep very precise notes about what they've done in this experiment. Impress upon them that a casual reader should be able to exactly duplicate their experiment simply by viewing their procedure. If necessary, do a prelab activity in which your students write the procedure for making a ham sandwich. Reading these procedures aloud should show the kids the shortfalls in the casual descriptions they're used to making.

Because this is the first lab, make sure you emphasize lab protocols. In this lab, it's particularly important that the students return all of the toothpicks and sponges back to the proper containers – when the toothpicks get jumbled together, it's impossible to resort them back into their original groupings without individually going through each one. If your students are sloppy about clean up, have them do it again to show that you won't tolerate carelessness in the lab. Similarly, it's probably a good idea to grade the first lab more harshly than any other. This will serve to train your students to be complete when writing up their labs, and will ultimately save them trouble in the long run.

Tips for doing this lab on a tight budget:

One way to save money when doing this lab is to limit the choices of things the students can change during part two of the lab. For example, instead of allowing them to use different types of sponges, have them alter either the type of toothpick used or the construction style. If you're really on a tight budget, only provide one type of toothpick and one type of sponge – the only variable available for the students to change will be the construction style.

Solutions for the Scientific Method Lab

Part 1:

1) The students should measure the height in centimeters.
2) Students should give a brief description of the tower, as well as a fairly accurate drawing of their structure. Points are assigned based on how complete and precise their descriptions are.

Part 2:

The six steps of the scientific method should be filled out using their own guesses and information. Points are assigned based on how complete and precise these explanations are. A good rule of thumb: If you can't understand exactly what something means, your students haven't done a good job of expressing themselves.

Post-construction:

1) The height should be given in centimeters.
2) If the tower was taller, it's because their hypothesis was correct and [whatever they did] is superior to the previous method of construction. If the first tower was taller, their hypothesis was flawed and your students should explain why.
3) Having a shorter second tower does not make the experiment worthless. Finding out what doesn't work is nearly as valuable as determining what does because it tells us how not to do something, and the experiment may give ideas about how it may be improved in the future.
4) Some students will answer this by saying that the scientific method makes experiments easier because it allows one to arrange their thoughts logically and with maximum precision. Other students may claim that the scientific method makes experiments more difficult; common explanations are that the scientific method hinders the creative process and that it takes a long time to record all of the necessary documentation. Though these arguments have some validity to them, explain to your students that the very methodical nature of the scientific method is what makes it useful and allows scientists to make the most of the data from each experiment.
5) Students should give a description of what they would do differently, basing their changes on the experimental results from parts one and two of this lab. For example, if a lab group finds that using two toothpicks instead of one is a useful construction method, perhaps they would try using three for their next experiment.

Solutions for the Scientific Method Worksheet

There are many possible correct responses to the questions on this worksheet. Questions 1-6 ask the student to solve a problem requiring that they stay warm in a powerless house during the winter. As long as students give responses to this scenario that follow the scientific method, full credit should be given to their answers.

Just to warn you, some students will take this opportunity to give you funny, odd, or disturbing answers to these questions. For example, there's always somebody who claims that the best way to keep warm during the winter is to set the house on fire with a can of gasoline. You shouldn't be too concerned about this, but may want to keep an eye on these students in the future.

On problem 7, students are asked to come up with an example of when they've used the scientific method in their own lives. Again, some students will give you disturbing examples to see your reaction. It may not be a bad idea to speak with these students about curtailing their use of inappropriate subject matter in their chemistry assignments. Truly disturbing examples (suicide, physical or sexual abuse, etc.) should be referred to the school administration.

The Scientific Method

The scientific method is our way of investigating the world around us. In every investigation that scientists undertake, the following six steps are followed:

1) **Purpose**: What do you want to find out? What is the goal of the experiment?

2) **Hypothesis**: What do you think will happen? What is your explanation or guess about what is taking place? A good hypothesis is written as an "If ___, then ___" statement.

3) **Materials**: This is a complete list of all things needed to test your hypothesis.

4) **Procedure**: What will you do to test your hypothesis? This should be a very complete list of specific actions which should be taken.

5) **Results**: When you did the experiment, what happened? What did you see, hear, and smell? Both quantitative (numerical) and qualitative (non-numerical) data should be collected for most experiments.

6) **Conclusion**: What do the results mean? Was your hypothesis correct?

Scientific Method Lab

In this lab, you will be doing an investigation using the scientific method. As you know, the scientific method consists of six steps: Purpose, hypothesis, materials, procedure, results, and conclusion. Each of these steps should be followed when pursuing a scientific investigation.

Your goal in this lab is to make as tall a tower as you can using only the equipment provided to you by your teacher. You'll do two experiments: In the first experiment you'll use the materials available to make a toothpick tower. In the second experiment you'll change one of the variables in your experiment with the goal of making a taller tower. Your success in building a taller tower will allow you to determine if your hypothesis was correct.

Experiment #1: Making a tower from "standard" materials

Using the materials provided to you by your teacher, your lab group should make as tall a tower as you can. At the end of the time allowed for building your tower, the height will be measured from the tower's base to its tallest point. Be careful – your tower may collapse unpredictably as it grows taller!

Post-construction:

1) How tall was your tower? (Don't forget units!) ______________________

2) Briefly describe what your tower looked like and sketch a diagram below:

Experiment #2: Building the better tower

Now that you've built one tower, you should be construction experts. Using your newly-developed tower-building skills you will build a second tower, identical to the first except that one thing will be changed. You may substitute one building material for another or change the design of your tower to make it more stable. However, make sure you only change one thing, because if you make more than one change it will be impossible to tell which change resulted in the height difference.

Record your second experiment using the scientific method. Remember to be complete!

Purpose:

Hypothesis:

Materials:

Procedure:

Results:

Conclusion:

Post-construction:

1) How tall was your second tower? ___________________________

2) Was your second tower taller or shorter than the first? If it was taller, explain why you believe your changes were so successful. If it was shorter, explain why you believe your hypothesis was wrong.

3) Some people in your class will doubtlessly find that their second tower was shorter than the first. Does this make their experiment worthless? Explain.

4) From your experience, does the scientific method make the process of experimentation easier or harder than random guessing? Explain.

5) If you had to build a third tower, what would you change to make it taller than your first two towers? Explain.

Scientific Method Worksheet

For problems 1-6, use the following scenario:

It's winter. The power has gone out in your house and the temperature inside is dropping rapidly. Fortunately, you're taking chemistry so you know how to use the scientific method to solve your problems.

1) What is your purpose?

2) What hypothesis do you have for solving your problem? Remember, you hypothesis should be written in the form "If ___, then ___".

3) What materials will you need to test your hypothesis?

4) What will your procedure be for testing your hypothesis?

5) What results will likely happen as a result of your experiment? (Use your imagination to answer this question).

6) What is your conclusion? (Remember that this should be a restatement of your hypothesis as being either true or false).

7) Think back to some time in your life when you had a problem that needed to be solved. Using the six steps of the scientific method, describe what you did to solve the problem:

a) Purpose:

b) Hypothesis:

c) Materials:

d) Procedure:

e) Results:

f) Conclusion:

Chapter 2 – Mixtures

Overview

Mixtures occur when several different elements or compounds are placed in the same container without undergoing chemical reaction. There are two main types of mixtures: **Homogeneous mixtures** occur when the composition of the mixture is uniform and **heterogeneous mixtures** occur when it is not. In this chapter, students will devise a way to separate a heterogeneous mixture of salt, black pepper, and sand using the equipment available to them.

Teaching About Mixtures

Students usually have an intuitive idea of what a mixture is before they come into a chemistry class. After all, many everyday processes involve mixing things together. An excellent way to tap into this prior knowledge is to have your students generate all of the examples of mixtures you describe in your lectures. For example, when you talk about a heterogeneous mixture, have them come up with an example from their lives. For homogeneous mixtures, this is a little more difficult, as students tend to give examples of chemical compounds (salt, water) rather than examples of mixtures. When this happens, refer back to the definition of a compound and explain that compounds differ from mixtures in that there's only one kind of molecule present in a compound.

Students frequently find themselves stumped by the task of separating a mixture. They know how to make mixtures, but quickly realize that separating mixtures back into their component parts is a much harder proposition. Emphasize to them that this difficulty is natural. After all, putting all of your dirty laundry into big piles in the corner of your bedroom is easy, but separating it into whites, colors, and delicates takes more time and effort.

Students realize very quickly that it's easy to identify and pull apart the components of a heterogeneous mixture. Homogeneous mixtures are another story: If everything in the mixture looks the same, how can you pull them apart? How can you even tell that there's a mixture in the first place? To explain how this is done, tell your students that they'll be learning more advanced separation techniques later in the year.

Doing the Separation of Mixtures Lab

Equipment:

- 1 kilogram (approximately 2 lbs) of granulated table salt
- 250 grams (approximately 0.5 lbs) of ground black pepper
- 500 grams (approximately 1 lb) of sugar
- 500 grams of sand, available with the cement in your local hardware store
- 12-50 mL beakers
- 6 funnels
- 6 filter papers that fit the funnels
- 6 forceps

Mix 500 grams of salt, all of the black pepper, and all of the sand in a large bowl. This will be the heterogeneous mixture for the main part of the lab. For the extra credit portion of the lab, mix the remaining salt and all of the sugar in another bowl. Make sure to label each of the mixtures clearly as “heterogeneous mixture” and “homogeneous mixture” to ensure that your students use the correct mixture for each part.

Safety:

There are no safety issues associated with this lab, though you may want your students to wear goggles so they get into the habit of doing so.

Room destruction factor:

When doing this lab, you can almost count on finding puddles of mixture slurry on the countertops. This mixture is easily wiped off with a wet paper towel and shouldn’t cause any permanent damage to your room.

How the lab works:

The best way to separate the heterogeneous mixture is to pour it into a beaker full of water. The sand will settle to the bottom, the salt will dissolve, and the pepper will float to the top. From here, it’s a simple matter to skim the pepper from the top and filter the solution to remove the sand from the salt water.

Some students will be convinced that the best way to separate a heterogeneous mixture is by picking it apart with forceps. These students will be convinced that they can tell the difference between grains of salt and sand, despite any evidence to the contrary. When these students have made their “separation”, have them pour their pile of “salt” into a beaker of water. Most likely, they’ll find that some of the salt falls to the bottom and won’t dissolve.

Some students finish this lab more quickly than others, necessitating the need for an extra credit section. In this lab, the extra credit involves the separation of a homogeneous mixture, namely the mixture of sugar and salt. Needless to say, this separation is difficult to do, particularly because the grains are visually indistinguishable and have similar solubilities in water. As far as I know, there's not an easy way to make this separation, though a sort of separation can be done by burning the sugar from the mixture. Technically, this does remove the sugar from the mixture, though it destroys one of the components. I leave it up to you to judge whether or not this is a valid separation.

What can go wrong:

Students may find that one component in the mixture floats on the surface when placed in water, while another sinks. These components are the same – sometimes sand "floats" on the surface of the water due to surface tension. If this happens, suggest to the students that they gently swirl the mixture to see if it has completely mixed. Their misconception should be cleared up when they see the particles sink to the bottom.

Solutions for the Mixture Separation Lab

Materials used:

As this lab is typically done at the beginning of the year, students are frequently careless about listing all of the materials used. By grading this section harshly, your students will learn to make a more complete materials list.

Procedure:

Again, students can be careless about describing exactly what they did in their lab procedure. Grade this section carefully to ensure they get into the habit of completeness.

Questions:

1) The mixture was heterogeneous. You can tell because some parts of it appear different than others. For example, the pepper is black, while the sand is tan and the salt is white.
2) You can't use forceps to separate this mixture because the particles are very small and have similar appearances. It would also take a very, very long time.
3) Different methods need to be used to separate different mixtures. While the use of water to separate this mixture was successful, it would be less successful when trying to separate two things that don't dissolve in water.

Extra credit:

Give some points for student effort on this section. While the separation will likely be unsuccessful, some students will get some genuinely clever ideas about how it may be done and should deserve some reward.

Solutions for the Mixtures Worksheet

1a) Pudding is a homogeneous mixture.
1b) Apples are heterogeneous mixtures.
1c) Honey is a homogeneous mixture.
1d) Cats are heterogenous mixtures.
2) When there is only one type of atom present, a material is referred to as an element. When there is only one type of molecule present, it is a compound. Mixtures arise when there are two or more types of molecules put together, or two or more elements that are not chemically bonded.
3) Homogeneous mixtures are usually more difficult to separate than heterogeneous mixtures, because in a heterogeneous mixture it's usually easy to see where each of the individual components is. For example, it's simple to distinguish the lima beans from the corn in succotash, but difficult to tell where the sugar in fruit punch is.

<u>Mixtures Handout</u>

There are two main types of mixtures:

1) **<u>Heterogeneous mixture:</u>** Any mixture that does not have uniform composition throughout the sample. If one part of the mixture looks different than another part, the mixture is heterogeneous.

   Examples:

   Pulpy orange juice: The pulp and juice are clearly different.
   Beans and rice: The beans are easy to distinguish from the rice.
   Dirt: You can tell the rocks apart from the grains of sand.

2) **<u>Homogeneous mixture:</u>** Any mixture that has completely uniform composition throughout the sample. If you were to take a sample from one part of the mixture, it would have exactly the same composition as a sample from any other part of the mixture. Homogeneous mixtures are also referred to as <u>solutions</u>.

   <u>*Examples*</u>:

   Fruit punch: It's a uniformly colored liquid.
   Air: It's a perfectly uniform mixture of many gases.
   Steel: It's a solid that contains iron and many other elements.

Separation of a Mixture Lab

In this lab, we will be separating a mixture that contains salt, pepper, and sand. By the end of this lab, all three components should be physically separate from the others. To do this, you may use any equipment that's available in the lab. My one and only hint to you is to *think creatively* when solving this problem.

Materials used:

Please list any materials you used in this lab:

Procedure:

Please list the steps you used to separate the three components of this mixture. This procedure should be complete enough that I could reproduce it exactly using your directions!

Have your teacher initial here when your separation is complete: ______________

Post-lab questions:

Use complete sentences to answer the following questions:

1) Was the mixture that you were given an homogeneous or heterogeneous mixture? How could you tell?

2) Since all three things in the mixture were solids, why couldn't you just use forceps to separate them from each other? Explain.

3) Could you use the methods you used in this lab to separate all mixtures? Why or why not?

Extra Credit:

Obtain a mixture of salt and sugar from the teacher and separate them. On a separate sheet of paper, write a complete description of the materials you used and the procedure required to make your separation.

<u>Mixtures Worksheet</u>

Use your knowledge of mixtures to answer the following questions:

1) Classify the following mixtures as being either heterogenous or homogeneous by circling the appropriate word for each example:

a)	Pudding	**heterogeneous**	**homogeneous**
b)	Apple	**heterogeneous**	**homogeneous**
c)	Honey	**heterogeneous**	**homogeneous**
d)	Cat	**heterogeneous**	**homogeneous**

2) Explain the difference between an element, a compound, and a mixture.

3) Is there any way to separate a homogeneous mixture? As an example, describe whether it's possible to separate salt water. If it is possible, explain how. If it's not possible, explain why not.

4) Why is it easier to separate a heterogeneous mixture than a homogeneous mixture? Explain.

Chapter 3 – Significant Figures

Overview

“Significant figures” is a term that refers to the number of digits in an experimentially derived number that give useful information about the data quality. Data with many significant figures is considered to be precise, and usually implies greater accuracy. In this chapter, students will learn the rules for writing and manipulating significant digits. They will use this knowledge to give the correct number of significant digits for data collected in this lab.

Teaching About Significant Figures

Everybody has problems with significant figures! It’s difficult to figure out why we need significant figures, what the rules are for finding out how many significant figures a number has, and how to do mathematical operations with significant figures.

The single most important thing to teach about significant figures is that we need them because we live in the real world. Although we can imagine finding a measurement to perfect accuracy with some hypothetical instrument, we never actually do because real instruments aren’t infinitely accurate. Because our instruments aren’t perfect, it’s important that we somehow indicate how good our instruments are to anybody looking at our data. We do this by limiting the number of digits we write in a measured number to the significant figures. An example, if I were to tell you that I weighed 80.6388 kilograms, you’d probably assume that I gave you four digits past the decimal because I weighed myself on a special scale that can measure things to that degree of precision. You wouldn’t assume I just used my bathroom scale because the number of significant figures is too high.

Another problem everybody has is with the difference between precision and accuracy. Precision is a measure of how reproducably you can take a measurement. For example, if you measure the weight of a paper clip to be 1.0025 grams, 1.0026 grams, and 1.0025 grams when you weigh it three times, you have a very precise measurement. High precision (or high reproducability) is denoted by a large number of significant figures.

Accuracy, on the other hand, is a measure of how close a measured value is to the actual value. If the paper clip I weighed in the example above actually weighed 1.9871 grams, my measurements wouldn’t have been very accurate because they don’t reflect the true mass of the paper clip.

As you might imagine, very precise measurements are usually also very accurate. This runs counter to what chemistry textbooks say, which is that precision and accuracy are independent of one another. Imagine this: You're designing an instrument that's supposed to reproducably measure the weight of very small objects to a very high degree of precision. When you're done, the balance can weigh objects accurately to the nearest millionth of a gram. If you went to all the time and trouble to make such an instrument, wouldn't you also spend the time and trouble to make that instrument accurate?

Typically, very high quality instruments measure things with high precision and accuracy. High quality instruments that are out of calibration are able to measure things with high precision but low accuracy. Inexpensive instruments usually have low precision and accuracy. That's why a laboratory balance costs $300 USD while a bathroom scale costs $8.

When you're teaching your students the rules for significant figures [p. X], make sure you give your students several examples of each. There's only so far you can go with teaching the definitions – real comprehension comes with repetition. The worksheets included with this lab are good introductions to significant figure problems, but you'll probably want more examples to do as homework or in-class assignments. The more that students practice significant figure problems, the better they'll be at doing them.

Doing the Significant Figure Lab

Equipment:

Station 1:

- One 50 mL graduated cylinder
- Three disposable pipets
- Three pen caps (large marker caps work best)
- 100 mL tap water in a beaker

Station 2:

- One 10 mL graduated cylinder (must have 0.1 mL gradations)
- Three disposable pipets (should be identical to station 1)
- Three pen caps (should be identical to station 1)
- 100 mL tap water in a beaker

Station 3:

- Five pennies
- Five large paper clips
- One digital balance

Station 4:

- Five pennies
- Five large paper clips (identical to station 3)
- One triple-beam balance

Stations 5 and 6:

- One pad of Post-it™ notes at each station
- Five large paper clips at each station

The stations in this lab are paired: Stations 1 and 2 are paired, as are 3 to 4 and 5 to 6. All equipment in the paired stations should be identical. For example, the paper clips in stations 5 and 6 should be from the same box.

<u>Safety:</u>

There are no safety issues in this lab and goggles are not required.

<u>Room destruction factor:</u>

Because the equipment used doesn't generate a mess, neither will this lab.

<u>How the lab works:</u>

Each pair of stations is measuring the same thing: Stations 1 and 2 measure volume, stations 3 and 4 measure mass, and stations 5 and 6 measure length. Ideally, the measurements should be the same, varying only in the number of significant figures that are appropriate for each measuring tool. As a result, most of the grading for this lab is not in the absolute value of the measurements but in the number of decimal places used to write the measurements. Make sure your students are clear on how to properly write the correct number of significant figures for a measured value before setting them loose.

This lab goes quickly. Three to five minutes for each lab station should be sufficient time to complete the questions before rotating to the next station.

<u>What can go wrong:</u>

- It may not be a bad idea to review the use of a triple-beam balance before starting this lab. Some students will need the tutorial before starting station 4.
- The objects being measured in this lab frequently get lost. The equipment list takes this into account, but be ready to produce more replacements if you have a lot of classes doing this lab at once.

Solutions for the Significant Figure Lab

Stations:

Station 1: Because 50 mL graduated cylinders measure volume to the nearest milliliter, answers should be given to the nearest 0.1 mL.
Station 2: Because the 10 mL graduated cylinder will measure volume to the nearest 0.1 mL, answers should be given to the nearest 0.01 mL.
Station 3: The number of significant figures given in each response should be the same as the number of significant figures on the electronic balance readout. Make sure students are not estimating an extra digit, as they do in the other stations.
Station 4: Student answers should have one more decimal place than the smallest gradations on the triple beam balance.
Station 5: Student answers should be to the nearest tenth of a centimeter, as the ruler given to them accurately measures centimeters.
Station 6: Student answers should be to the nearest hundredth of a centimeter, as their ruler measures to the nearest millimeter.

Post-lab questions:

1) We cannot use as many significant figures as we want because significant figures are a reflection of the precision of the instrument used to take a measurement. It doesn't make sense to measure down to the nearest millionth of a centimeter with an ordinary ruler because the ruler simply isn't that sensitive – any digits past the first estimated one are random guesses.
2) If we had an accurate enough instrument, there would be no problem with writing fifteen decimal places. However, for this to be valid, the measurement tool would have to have gradations accurate to 14 decimal places so we could estimate the fifteenth digit.
3) Precision is a measure of how reproducable a measurement is and is denoted by the number of significant figures written in a number. Accuracy is a measure of how close a measured value is to the actual value. Though precision and accuracy aren't necessarily related, precise measurements are frequently also very accurate.

Solutions for the "Significant Figures Worksheet"

1) three
2) three
3) three
4) two
5) four
6) one
7) four
8) two
9) three
10) four
11) three
12) three
13) two
14) three

(continued next page)

15) A precise measurement would not be accurate if a very precise piece of uncalibrated equipment was used. In such a case, there would be the same systematic error in each measurement.
16) Lack of precision implies that you can't reliably reproduce a measurement. If a measurement can't be reproduced, than many of the measurements around it will be inaccurate.
17) The ruler that measured "7.50 centimeters" was telling me that I am able to accurately measure the length of my thumb to the nearest hundredth of a centimeter. The ruler that measured "7.5 centimeters" only gives reliable information to the nearest tenth of a centimeter. The number is the same for both measurements, but the meaning of the number is not.

Solutions for "Using Significant Figures in Calculations"

1) 6.84 → **6.8**
2) -0.5 → **-1**
3) 9.411 → **9.41**
4) 7.888 → **7.9**
5) 3.378261 → **3.4**
6) 0.321488 → **0.321**
7) 28080 → **28,000** or **2.8×10^4**
8) 0.024 → **0.02**
9) 78.512 → **78.5**
10) (30**.**) x (11.3) = 339 → **340**
11) 7.451613 → **7.5**
12) 65.0023 → **65**
13) 3,610,349 → **3,600,000** or **3.6×10^6**
14) (3.81) + 2.45001 – (920) = 913.74 → **910**
15) It's important to use the correct number of significant figures when solving a problem because the number of significant figures in the answer tells the reader about how trustworthy the answer is. If you write too many significant figures, the data appears more precise than it really is. If you write too few significant figures, the data gives less useful information than it should.

<u>Significant Figures</u>

Significant figures are the digits in a measured number that indicate the measuring equipments degree of precision. Generally, when writing down a measurement, you should write all of the digits that you obtained directly with the measuring device and add a final digit that you've estimated. For example, if you have a ruler that can measure length in millimeters, you should write the lengths of objects you've measured to tenths of millimeters.

<u>Rules for Writing Significant Figures:</u>

1) **All nonzero digits are significant.** For example, "3.4 grams" has two significant figures.

2) **Zeros that are between nonzero digits are significant.** For example, "3.04 grams" has three significant figures.

3) **Zeros written to the left of all nonzero digits are <u>not</u> significant**. For example, "0.0034 grams" has two significant figures.

4) **Zeros written to the right of all nonzero digits are only significant if a decimal point is written in the number.** For example, "1000 grams" has one significant figure, while "1000.0 grams" has five. The zeros in the second number indicate that a value can be measured accurately to the nearest tenth of a gram, while writing simply "1000 grams" indicates that the measurement has been rounded to the nearest thousand grams. While both mean the same thing to your calculator, they don't mean the same thing to a reader.

5) **Numbers in scientific notation have the same number of significant figures as the portion of the number that's before the "x 10^n" part of the number.** For example, "4.30 x 10^5 grams" has three significant figures.

<u>Rules for Using Significant Figures in Calculations:</u>

1) **When adding or subtracting, the answer should have the same number of figures to the right of the decimal as the value with the fewest decimal places.** For example, 3.4 + 5.023 = 8.423 → Round this to 8.4, because 3.4 has only one digit to the right of the decimal.

2) **When multiplying or dividing, the answer should have the same number of significant figures as the value with the fewest significant figures.** For example, 1.220 x 3.4870 = 4.25414 → Round this answer to 4.254, because 1.220 has only four significant figures.

Significant Figures Lab

In chemistry, we try to get the most information we can out of every measurement. Whenever we write down a number we've measured, the number of digit the number has reflects the precision of the instrument we used to get it.

This is important in science because when we read somebody else's data we like to know how precise their data really is. If we use the wrong number of digits in our answers, we might fool people into believing that imprecise data is really precise, or vice versa.

When taking measurements, you should always write all values so they show the smallest marking on the instrument, plus an extra digit that you estimate. For example, if you use a ruler that has lines for millimeters, you should write your answers to the nearest tenth of a millimeter because you can estimate the last digit. The exception to this rule is digital equipment, such as an electronic balance. Because you can't estimate the last digit on a digital balance, simply write down the answer on the readout.

In this lab, you will be measuring length, volume, and mass using common laboratory instruments. For each of these tools, you must write down your answer with the correct number of significant digits! Remember, the number of digits you write depends on the instrument you used to take the measurement.

Station 1: Measuring volume with a 50 mL graduated cylinder

Use the 50 milliliter graduated cylinder at this station to find the following volumes. Be sure to use the proper number of significant figures in your answer!

1) What is the maximum volume of the pipet? _______________

2) What is the maximum volume of the pen cap? _______________

Station 2: Measuring volume with a 10 mL graduated cylinder

Use the 10 milliliter graduated cylinder at this station to find the following volumes. Be sure to use the proper number of significant figures in your answer!

1) What is the maximum volume of the pipet? _______________

2) What is the maximum volume of the pen cap? _______________

Station 3: Measuring mass with an electronic balance

Use the electronic balance at this station to find the following weights. Be sure to use the proper number of significant figures in your answer!

1) What is the mass of the penny? _______________

2) What is the mass of the paper clip? _______________

Station 4: Measuring mass with a triple beam balance

Use the triple beam balance at this station to find the following weights. Be sure to use the proper number of significant figures in your answer!

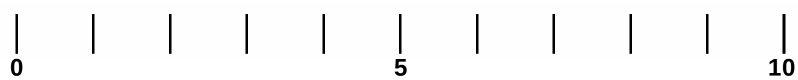
1) What is the mass of the penny? _______________

2) What is the mass of the paper clip? _______________

Station 5: Measuring distance with a ruler

Using the ruler printed below, find the following lengths. Be sure to use the proper number of significant figures in your answer!

Ruler:



centimeters

1) What is the length of the post-it note? _______________

2) What is the length of the paper clip? _______________

Station 6: Measuring distance with a ruler

Using the ruler printed below, find the following lengths. Be sure to use the proper number of significant figures in your answer!

Ruler:

centimeters

1) What is the length of the post-it note? _______________

2) What is the length of the paper clip? _______________

Post-lab questions:

1) Why can't we write numbers with as many significant figures as we want? For example, if we measure something with an ordinary ruler, why is it wrong to write our measurement as "0.928772662 centimeters"? Explain.

2) If we had an accurate enough instrument, is there any reason we couldn't write down a value to 15 decimal places (as in the number 0.123456789012345)? Explain.

3) In your own words, what's the difference between precision and accuracy?

Significant Figures Worksheet

For problems #1-14, write down how many significant figures each number has:

1) 4.53 __________

2) 2.30 __________

3) 1.02 __________

4) 4500 __________

5) 3200**.** __________

6) 0.002 __________

7) 1.000 __________

8) 0.070 __________

9) 0.707 __________

10) 7.070 __________

11) 7070 __________

12) 0.700 __________

13) 3.4×10^4 __________

14) 1.02×10^2 __________

Please answer the following questions:

15) Under what circumstances might a very precise measurement not be accurate? Explain.

16) Are there any circumstances under which an accurate measurement may not be precise? Explain.

17) I measured the length of my thumb and found that it is 7.50 centimeters long. When I used another ruler, I found that the length was 7.5 centimeters. Explain the difference between these two measurements.

<u>Using Significant Figures in Calculations</u>

Solve each of the following math problems and write their answers with the correct number of significant figures:

1) 4.5 + 2.34 = ____________________________

2) 4.5 – 5 = ____________________________

3) 6.00 + 3.411 = ____________________________

4) 3.4 x 2.32 = ____________________________

5) 7.77 / 2.3 = ____________________________

6) 3.890 / 121 = ____________________________

7) 1200 x 23.4 = ____________________________

8) 120 x 0.0002 = ____________________________

9) 78.5 + 0.0021 + 0.0099 = ____________________________

10) (3.4 x 8.90) x (2.3 + 9.002) = ____________________________

11) (2.31×10^3) / (3.1×10^2) = ____________________________

12) 0.0023 + 65 = ____________________________

Ad*vanced problems:*

13) (3.4×10^6) + 210,349 = ____________________________

14) 1.09 x 3.498 + 2.45001 – 2.123 / 0.0023 = ____________________________

15) Why is it important to always use the correct number of significant figures when solving a problem?

Chapter 4 – The Flame Test Lab

Overview

The flame test is used to identify the elements in an unknown sample. During this test, the sample is heated over a flame, causing light to be emitted from the element. The color of the light given off corresponds to the energy difference between the ground and excited states of the element. Because each element has a unique energy difference between the ground and excited states, each element emits a unique color when heated. In this chapter, students will learn about the mechanism by which light is given off. In the lab, they will use the flame test to identify several unknown compounds.

Teaching About Atomic Structure

Teaching the structure of the atom isn't difficult. Protons and neutrons exist in the nucleus and electrons exist as waves in a diffuse cloud around it. The ideas are fairly straightforward and students generally don't have a conceptual problem with them.

The difficult part of teaching atomic structure is not in the structure but in trying to teach students why it's relevant. Textbooks mention something called the "flame test", but it's rarely made entirely clear how this differs from other spectroscopic methods, or if it's a spectroscopic method at all. Students frequently get through the chapter on atomic structure with only a vague idea of what the flame test is and how it is performed.

Making the flame test relevant should be the main focus of this lesson. Why is the flame test handy? It's handy because we can use the colors emitted by atoms to tell different elements apart. Can we tell every element apart from this test? Not with our flame test, but the idea of spectroscopy allows us to do so with a good detector and knowledge of spectral lines. The very simplicity of the flame test may give some students the idea that it's a watered down version of spectroscopy. It's important that you make it clear that there's no real difference between the flame test and more complex spectroscopies, except in the sensitivity of the equipment used.

It's not a bad idea to give students specific examples of what spectroscopy is used for. Water samples are tested for heavy metals and other contaminants using spectroscopy. Elemental compositions of stars and planetary atmospheres can be determined by the emission lines given off. Hair, blood, and urine samples can be tested to see if you're being poisoned. The list of things that spectroscopy is used for is huge – this importance to everyday life should be emphasized to your students.

Doing the Flame Test Lab

Equipment:

- 5 Bunsen burners
- 5 flint strikers
- One box of flat toothpicks
- 100 mL each of five 0.5 molar solutions: $CaCl_2$, KCl, LiCl, NaCl, and $SrCl_2$. Small quantities (~10 mL) of each solution should be transferred to medium-sized vials for use in the lab.
- 10 medium-sized vials for holding these solutions. Five vials should be labeled with the formulas of the compounds while another five should contain the same five solutions but with labels saying "Unknown 1" through "Unknown 5". To prevent fungal growth in these vials, add 1 drop of concentrated HCl to each one.
- 5 sets of forceps or crucible tongs

Safety:

In this lab the students will be holding flaming toothpicks with forceps or crucible tongs. Because Bunsen burners are being used, it's particularly important that you pay close attention to ensure your students are staying on task and mindful of the safety rules.

Some of the chemicals in this lab are dangerous and care should be taken to ensure that your students wash their hands immediately after concluding the lab. Particularly dangerous are lithium chloride and strontium chloride.

Do NOT let your students wear rubber gloves while doing this lab! Rubber gloves, while giving good protection against chemical contamination, can easily melt to the skin if heated above their melting point. For this reason, rubber gloves should NEVER be used around flames.

Most importantly, make sure all students are wearing goggles at all times!

How the lab works:

Each lab station will have fifteen or so toothpicks soaking in each of the solutions mentioned above. To do the flame test, your students will pick the wet toothpicks out of the solutions and hold them over the Bunsen burner. The color of the flame given off immediately after the toothpick is placed in the Bunsen burner is the color that students are trying to find. It's particularly important that the toothpicks be kept in the solutions when not in use because the burning wood in flaming toothpicks appears yellow. By using wet toothpicks, the solution will vaporize before the wood burns, making the colors of the heated metal ions easy to see.

When the students have finished with the toothpicks, they may be placed back into their original solutions. Even though the toothpicks may look scorched (particularly after a few classes have done the lab), the results will be the same. The exception to this is when the toothpicks are dropped on the countertops or on the floor. These toothpicks should be thrown away, as they have probably picked up other contaminants from every surface they've touched.

In the second part of the lab, the solutions should be exactly the same as the ones used in the first part of the lab – the only difference should be that they're labeled as unknowns rather than as "calcium chloride", etc. This will ensure that each element will have the same color flame for both sections of the lab, making identification easier.

<u>What can go wrong:</u>

- All of the samples have exactly the same color flame. This may be caused for a variety of reasons. Most common is that students are holding the toothpick over the Bunsen burner for a very long time, resulting in a yellow flame caused by the burning of the wood in the toothpick. This problem can be solved by emphasizing to students that the color they're interested in is found immediately after placing the wet toothpick in the fire and not after a few seconds. Another reason for uniform flame colors is contamination of the solutions. If you suspect your solution is contaminated, simply throw it out and make a new one.
- The students find that one of the unknown samples doesn't match any of the samples from part 1 of the lab. This is most likely caused by inexact observations. What one student may term "orange-yellow" when recording a color may be interpreted by another as "lemon yellow". In such a case, encourage your students to go back and reexamine the colors of the known elements.
- The Bunsen burners don't work properly. This occurs because the students aren't familiar with the operation of the Bunsen burners. To avoid this problem, review the use of Bunsen burners before starting this lab.

<u>Solutions for the Flame Test Lab</u>

Instead of giving you the colors of each of the elements in this section, I'll let you find them for yourself. Why? Depending on the age and purity of your reagents, your colors may be somewhat different from the pure substance. Another problem is that what I may define as “dark purple” may be different from what you would call “dark purple”. To avoid these problems, determine the colors that your chemicals make before doing this lab – this way, you can ensure the lab works before you give it to your students.

Prelab questions:

1) Different elements give off different colors of light when heated because there is a unique energy gap between the ground and excited states for all elements. The color of light emitted is directly related to this energy gap.
2) Although all elements give off unique colors of light, there are so many different elements that many have similar colors to the naked eye. Only by using equipment that can better see the differences can we tell them apart.

Post-lab questions:

1-5) These answers will depend on what element is in each unknown solution.
6) There are two answers which may receive credit for this question:
   - This method is effective for identifying any unknown element because all elements have a unique energy gap between the ground and excited states.
   - This method would not be effective for identifying any unknown element. Although every element has a unique energy gap, two elements may have a close enough energy gap that the color difference will be indistinguishable to the naked eye.
7) If the toothpicks were contaminated, the colors would be changed.
8) Flame color is determined by the energy difference between ground and excited state. Because this energy difference is unique for all elements, all elements give off unique colors of light when burned.
9) Any example involving the identification of unknown elements is acceptable.

Solutions for the Electrons and Orbitals Worksheet

1) The sketch should include two protons and two neutrons in the nucleus, as well as two electrons circling the nucleus in a single orbital.
2) When an atom absorbs energy in the form of light, heat, or electricity, an electron jumps from the ground state to an excited state. When the electron returns to the ground state, light with a color corresponding to the energy difference between ground and excited state is emitted.
3) Compare the line spectrum of the sun to that of platinum. If lines in the solar spectrum correspond to platinum emission lines, the sun contains platinum.
4) Energy is commonly added to atoms with electricity and light.

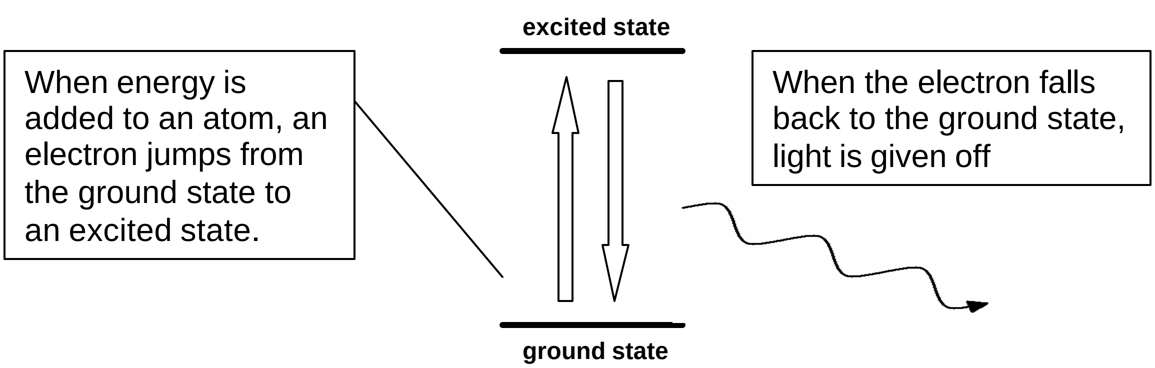
Orbitals Handout

Electrons are found in orbitals that surround the nuclei of atoms. The four types of ground state orbitals include s-orbitals, p-orbitals, d-orbitals, and f-orbitals. In each energy level, s-orbitals have the least energy, followed by p-orbitals, d-orbitals, and f-orbitals.

When energy is added to an atom in the form of light, heat, or electricity, the electrons absorb this energy and use it to jump to higher energy levels called excited states. However, since electrons like to have low energies whenever possible, they eventually return to their original orbitals.

As you've learned, energy can never be created or destroyed. Since this is true, what happened to the energy the electron absorbed when it jumped to the excited state? It has to go *somewhere*!

When electrons return to the ground state, energy is released from the atom as light. The color of the emitted light corresponds to the amount of energy it has. In our case, the amount of energy corresponds exactly to the difference in energy between the ground and excited states. As a result, we can use the color of the light to determine the differences in the orbital energies.



Why is this important? The differences in energy between the ground and excited states are different for every element. As a result, when we see the light that's given off by an atom, we can use the color of this light to tell us what element is present. if we don't know what element is in a sample but have a good idea of the colors given off by all elements in the periodic table, all we need to do to identify our sample is find an element that matches the color of our unknown.

This is the basis for spectroscopy. Though the technology used in modern spectroscopy is far more advanced than what you can do in chemistry class, the idea is exactly the same: By comparing the colors of light given off by an unknown sample to the colors of light given off by known elements, you can determine the identity of the unknown.

Flame Test Lab

As we've discussed in class, atoms give off light after energy has been added to them. This happens because electrons in ground state (low energy) orbitals absorb energy and move to excited state (high energy) orbitals. When these electrons fall back to their original orbitals, the excess energy is given off as light. The color of this light is closely related to the amount of energy released.

This phenomenon can be used to identify unknown elements. Let's say that you have two elements: Element A is known to give off red light when heated, and Element B gives off blue light. If you heat an unknown sample with a flame and the color given off is blue, it seems reasonable to assume that your unknown element is Element B. In this lab, we will be doing the same thing using real elements.

Prelab:

1) In your own words, explain why different elements give off different colors when heated.

2) Even though all elements give off unique colors when heated, it's impossible to identify all elements with the naked eye by doing this test. Explain why.

Section 1: Determining some reference colors

In this section we will be determining the colors that different metallic elements give off when heated. In section 2 we will compare the colors of unknown elements to these reference colors to identify the unknown elements.

There are five stations in this lab. At station 1 you will be determining the color given off by heated calcium by heating a sample of calcium chloride. At station 2, you will determine the color given off by heated potassium by heating potassium chloride. In the same way, you will determine the colors given off by heated lithium in station 3, heated sodium in station 4, and heated strontium in station 5.

Procedure (you may start at any of the five stations):

1) Light your Bunsen burner with a flint striker. Remember to turn the gas off and let it disperse if the burner doesn't ignite in the first three tries.

2) Using forceps, pick up a toothpick that has been soaked in a solution containing one of the five samples (i.e. calcium chloride, potassium chloride, etc.). Make sure you don't touch the toothpick with your bare hands because you may contaminate it.

3) Place the toothpick into the Bunsen burner flame directly over the bright blue cone. Record the color that you observe as the sample is heated. Important: Only record the initial flash of color that you see, NOT the orange-yellow color that comes after the toothpick has been burning for a few seconds. The initial color is the color of the chemical the toothpick has been soaking in, while the color that comes later is simply the color of burning wood. If you do this step correctly, the toothpick should not be badly burned.

4) Return the toothpick to the solution it has been soaking in.

5) When you are told to do so, switch stations and repeat this experiment with the other four stations.

Data:

1) Color of calcium (Ca): ____________________________

2) Color of potassium (K): ____________________________

3) Color of lithium (Li): ____________________________

4) Color of sodium (Na): ____________________________

5) Color of strontium (Sr): ____________________________

Section 2: Identifying the unknowns

In this section you will compare the colors of the unknown solutions to the known colors of the elements in section 1 in order to identify them. For example, if you found that strontium gave off neon green light when heated in section 1 and unknown solution #1 gives off neon green light, you can say with certainty that unknown solution #1 contains strontium.

Procedure:

1) Light your Bunsen burner with a flint striker. Remember to turn the gas off and let it disperse if the burner doesn't ignite in the first three tries.

2) Using forceps, pick up a toothpick that has been soaked in a solution containing one of the five unknowns. Make sure you don't touch the toothpick with your bare hands because you may contaminate it.

3) Place the toothpick into the Bunsen burner flame directly over the bright blue cone. Record the color that you observe as the sample is heated. Important: Only record the initial flash of color that you see, NOT the orange-yellow color that comes after the toothpick has been burning for a few seconds. The initial color is the color of the chemical the toothpick has been soaking in, while the color that comes later is simply the color of burning wood. If you do this step correctly, the toothpick should not be badly burned.

4) Return the toothpick to the solution it has been soaking in.

5) When you are told to do so, switch stations and repeat this experiment with the other four stations.

Data:

1) Color of Unknown 1: ______________________________

2) Color of Unknown 2: ______________________________

3) Color of Unknown 3: ______________________________

4) Color of Unknown 4: ______________________________

5) Color of Unknown 5: ______________________________

Post-lab Questions:

Using the data you collected in parts 1 and 2, write the names of the unknown elements in questions 1-5:

1) What element was in unknown 1? ______________________________

2) What element was in unknown 2? ______________________________

3) What element was in unknown 3? ______________________________

4) What element was in unknown 4? ______________________________

5) What element was in unknown 5? ______________________________

6) Do you think that this method of identifying unknowns would be effective for identifying any unknown element? Explain why or why not.

7) Explain what you think would happen if the toothpicks used in this lab became contaminated.

8) In this lab, you observed that the colors of the flames in each sample are different. Why don't all flames have the same colors?

9) How might this method of identifying unknown elements be important in real life? Explain when it might be handy to identify unknown elements.

Electrons and Orbitals Worksheet

Answer the following questions using your knowledge of atomic structure:

1) Sketch a helium atom, showing and labeling the nucleus, the protons, the neutrons, the electrons, and the orbitals.

2) Explain the process by which light is given off by an atom. In other words, what has to occur before an atom will give off light?

3) I have a theory stating that the sun contains a small amount of platinum. Using what you know about spectroscopy, explain how I might prove or disprove this theory.

4) In the flame test lab, we added energy to atoms using the heat of a Bunsen burner. Give some other ways we can add energy to atoms.

Chapter 5 – The Periodic Table

Overview

The periodic table was designed as a way to arrange all known elements so that their properties and reactivities are easier to understand. Elements in the same column are said to be in the same "group" or "family" and have similar chemical and physical properties. Elements in the same row are said to be in the same "period" and share an energy level. In this chapter, students will learn about the arrangement of elements in the periodic table, and explore the extreme difficulties experienced by the makers of the first periodic table.

Teaching about the Periodic Table

There is good news and bad news when it comes to teaching the periodic table. The good news is that there's nothing inherently difficult about learning the periodic table – students really only need to learn the locations and properties of the most important families. Frequently, this is taught as a coloring exercise – "Color the transition metals orange, the halogens blue," etc.

The bad news is that students find the periodic table to be very, very boring. After all, who wants to color a periodic table in class when there are more interesting things to be done?

How can we teach the periodic table in an exciting way? The bottom line is that one way or another, our students will just have to memorize the table in the traditional ways. However, by giving them equal quantities of memorization and activity, the memorization will be a little easier to take.

Doing the Periodic Table Lab

Equipment:

The equipment for this lab consists of five bags containing identical collections of small and easily obtained items from around the lab. These items should be inexpensive things like paper clips, flints, vials, pens, pencils, nails, thumbtacks, etc. When choosing these items, don't pick a collection of items that will naturally fall into easy categories – this lab works much better if the items are dissimilar in many ways. Along with the familiar items above, include some items your student probably haven't seen before, such as cork borers, solder, parafilm squares, etc.

Safety:

The items chosen for this lab shouldn't be dangerous or toxic (no road flares, blasting caps, sodium hydroxide pellets, and so on). As long as the items are harmless, safety is not an issue with this lab.

Room destruction factor:

This lab is only as messy as the items you place in the bags. By using items that are difficult to break, there should be no clean-up to speak of.

How the lab works:

In this lab, your students will treat each item as an "element" and attempt to arrange them into a "periodic table" based on their properties. This is harder than it might imagine, and takes a good deal of time.

Students typically have no difficulty separating the elements into groups. Popular ways of arranging the elements include by composition (wood, metal, plastic, etc.), by shape (round, pointed, etc.), and by color. The problem with this lab is that it's difficult to group the elements by period. After all, it's not enough that the elements are arranged by increasing size or mass – the elements have to be arranged in a way that elements next to each other have roughly the same size or mass, just as elements in the same period have roughly the same electron energies. This becomes difficult if two elements in the same group have identical masses, as only one element from each group can exist in each period.

As a result, simple groupings can't be used for the elements in this lab. Your students will show great inventiveness and creativity in their efforts to solve this problem, making the lab both interesting to do and to watch.

After the lab ends, you may want to have each group present their periodic tables to the whole class. Students will frequently be surprised at how other groups have arranged their elements, and will undoubtedly have a new appreciation for the difficulty involved in making the first periodic table.

What can go wrong:

- Sometimes students will never come to consensus within their groups. There's no much you can do to resolve these arguments, except to remind each group that this lab is graded, necessitating *some* answer.
- The "elements" get lost or broken. It's a good idea to have at least two spares of each item to keep the lab moving smoothly.
- Students sometimes play with the items instead of classifying them. To avoid this, don't put Silly Putty or Slinky coils in the bags!

Solutions for the Periodic Table Lab

Prelab:

Although fluorine and iodine may be dissimilar, they have some important characteristics in common. For example, both are good oxidizers, are diatomic, and prefer a -1 oxidation state. Though these elements have significant differences, the similiarities still outweigh the differences.

Lab:

The “elements” should be placed into groups by property. Make sure you count the items on this page – students have developed the clever trick of leaving out items that don’t fit into any clear category. Typcially, students place less massive items at the top of each group and heavier ones at the bottom, though different ways of denoting energy level may be favored by different groups. The final product should look (roughly) like a periodic table, though there may be more gaps than the one used to classify real elements.

Postlab questions:

1) One property should have been chosen to classify their items, and there should be some reasonable explanation for why this particular property was used.
2) There should be some continuum from the top to bottom of each column. For example, if they believe that three elements fit into a category, there should be a property that varies slightly as you move down the column. Common properties include mass, size, and density. Elements placed next to one another should also have the same mass, size, density, etc.
3) Students should discuss the differences in how their arrangement differs from those of a neighboring group. Any differences are most likely caused by an emphasis on different properties – their answer should reflect this.
4) Students should explain how their differences were resolved. Sometimes groups agree all the way down the line in how elements should be arranged, but more frequently they vote or decide through painful trial and error.
5) Clearly, emphasis on different properties made it difficult to devise the first periodic table.

Solutions for the Periodic Table Worksheet

1) alkali metals
2) transition metals
3) noble gases
4) actinides

5) Alkaline earth metals are frequently soft, low-density, reactive, and metallic elements that prefer to form ions with a +2 oxidation state.
6) Halogens are extremely reactive diatomic elements that prefer to form ions with a -1 oxidation state. Halogens may be gases (F_2, Cl_2), liquids (Br_2), or solids (I_2). All are volatile under normal conditions. They are difficult to handle and are strong oxidizers.
7) Elements in the same family have similar properties.
8) Elements in the same period have similar orbital energies.

Solutions to the "Make your own periodic table" worksheet

The elements shown correspond to the alkaline earth metals (group 2), group 11, the halogens (group 17), and the noble gases (group 18). Specifically, the identities of the elements should be arranged in the following way:

beryllium (1)	copper (6)	fluorine (2)	neon (8)
calcium (3)	silver (5)	bromine (9)	argon (10)
barium (7)	gold (12)	astatine (11)	xenon (4)

<u>Periodic Table Handout</u>

The periodic table is what chemists have used for over a hundred years to organize the known elements.

Periods correspond to horizontal rows in the periodic tble. Generally, the elements within a period have little in common, except for the energy levels of the electrons.

Families (also called "groups") correspond to the vertical columns in the periodic table. Elements within each family share similar properties. These similarities occur because elements in the same family have the same number of valence electrons. Important families within the periodic table include:

- **Alkali metals (group 1)**: Extremely reactive, soft metals with low density that form ions with a +1 charge.
- **Alkaline earth metals (group 2)**: Slighly less reactive than alkali metals, they are somewhat denser and less soft. They form ions with a +2 charge.
- **Halogens (group 17)**: Highly reactive and electronegative nonmetallic elements that form ions with a -1 charge. They are diatomic, volatile, and very difficult to handle safely.
- **Noble gases (group 18)**: Very stable nonmetallic gases that react poorly with other elements.

Other important sections of the periodic table include:

- **Transition metals (groups 3-12)**: Dense, hard metallic elements that usually form ions with more than one possible positive charge.
- **Lanthanides and actinides** (the two rows at the bottom of the periodic table): The lanthanides are the top row and are reactive, dense metals. The actinides are the bottom row and include mainly radioactive elements that are produced artificially. Uranium is the most important actinide, used for nuclear power and weapons applications.
- **Main group elements:** These elements consist of groups 1, 2, and 13-18. They have very little in common except that they have either s- or p- electrons as their outermost electrons.

Periodic Table Lab

As you've learned, the periodic table is arranged in periods and families (families are also known as groups). Periods correspond to the horizontal rows – for example, lithium, beryllium, boron, carbon, and nitrogen are in the same period. Elements in the same period may have very little in common. For example, lithium reacts readily with water, while nitrogen requires very high temperatures to react with water.

The elements in each family of the periodic table have similar properties. For example, all of the alkali metals (group 1) react readily with water. The reason for this similarity in reactivity is that the electron configurations for every element in a family are similar, containing the same number of valence electrons in the same type of orbital.

We usually take the periodic table for granted, not thinking much about how it was invented. It seems obvious to us now that elements should be grouped by properties and electron configurations. However, back in the 1800's there were many alternative ways to arrange the elements because not all of them had been discovered and nobody really understood atomic structure. The periodic table we use today is the product of many years of revisions by the best scientists of that time.

In this lab, we're going to imagine we're scientists presented with a variety of new elements, and we're going to arrange them in the most logical way, based on their properties.

Prelab:

In the current periodic table, fluorine is shown in the same family as iodine. However, fluorine is a pale yellow gas, while iodine is a violet, shiny solid. Was a mistake made when these were placed in the same group? Explain.

Lab:

Your group will be given a bag containing 20 items. As scientists, it's your job to group these items into families and periods, based on their observable properties. You may make as many families and periods as you like, but there should be some order to how they are arranged. Use the rest of this sheet to draw your periodic table, and make sure you use a ruler!

Postlab questions:

1) What was the main property you used to classify the elements into groups? Explain why you chose this property and not another.

2) Explain how you decided which element should go at the top of each column and which should go at the bottom.

3) Look at the way another lab group arranged their elements. Is it the same as the way you arranged yours? Why or why not?

4) Did you have disagreements within your group about how these elements should be arranged? If so, explain how these disagreements were resolved.

5) Does this exercise give you any insight as to why it may have been difficult to invent the first periodic table? Explain.

Periodic Table Worksheet

For questions 1-4, fill in the blanks with the correct word or phrase:

1) The _________________________ are reactive, light metals that form ions with a charge of +1.

2) The _________________________ are dense, strong metals that have high melting and boiling points.

3) _________________________ are unreactive, nonmetallic gases.

4) _________________________ are mainly radioactive, manmade elements.

Answer the following questions:

5) What are the properties of the alkaline earth metals?

6) What are the properties of the halogens?

7) What similarities to elements in the same family share?

8) What similarities to elements in the same period share?

Make your own Periodic Table Worksheet

You've heard how the periodic table was invented and had a chance to make one in class. Now that you're a pro at classifying elements, you get a chance to make your very own periodic table using real elements. If you do this correctly, your classification scheme should be the same as the actual periodic table.

Unfortunately, you're not going to be given the names of the elements or a complete list of their properties. Using partial information (such as scientists had in the old days), see if you can arrange these real elements into their proper periods and families. One hint: These elements should be arranged into a grid that's three boxes tall by four boxes wide, with no blank spaces.

In no particular order:

Element 1: Solid, metal, does not corrode in air, density = 1.85 g/mL.

Element 2: Yellow gas, highly dangerous to handle, toxic in low doses.

Element 3: White, shiny, metallic solid, reacts slightly in air, density = 1.55 g/mL.

Element 4: Colorless gas, stable in air, forms very few chemical compounds.

Element 5: White, shiny metallic solid, unreactive, good electrical conductor, ductile, density = 10.5 g/mL.

Element 6: Orange-red metallic solid, ductile, density = 8.9 g/mL.

Element 7: White metallic solid, reacts easily in air, density = 3.5 g/mL.

Element 8: Colorless gas, unreactive with any element.

Element 9: Red nonmetallic liquid, irritates skin and lungs.

Element 10: Colorless gas, denser than air, forms no chemical compounds.

Element 11: Radioactive metalloid, very little known about its properties.

Element 12: Yellow metallic solid, extremely malleable, unreactive with most chemicals, density = 19.3 g/mL.

Good luck!

Chapter 6 – Moles, Molecules, and Grams Calculations

Overview

Three of the units most commonly used in chemistry classes are moles, molecules, and grams. Moles and molecules are useful because they allow us to determine how many particles are involved in a chemical reaction. Grams are useful because it is easier to measure mass than molecules or grams in the lab. In this chapter, students will learn how to convert between the units of moles, molecules, and grams, as well as understand why each is important. In the lab, students will determine the numbers of grams, moles, and molecules present in several samples of chemical compounds.

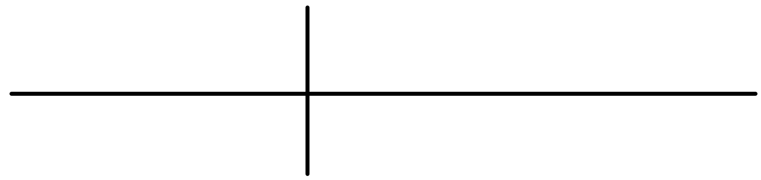
Teaching about Mole Conversions

Students hate learning about mole conversions because they feel as if they need to memorize each type of calculation separately from the others. How do I convert from moles to molecules? Memorize how to solve the problem, then solve them all like this. How do I convert from molecules to moles? Memorize how to set the problem up, and then solve them all like this. No wonder students hate mole calculations! There are six different things to memorize!

Obviously, all this memorization isn't necessary if students learn how to methodically set up these problems. Most of use teach the factor-label method, where the units end up canceling each other in the numerator and denominator at the end of the problem. The difficulty with this method is that students get tripped up when math terms like "numerator" and "denominator" are used. This is a problem all science teachers are familiar with: Information from math classes isn't easily transferred to other parts of the school building.

I teach using the factor-label method, but don't ever use the term "factor-label method" in class. The very term tends to make students believe that unit conversions are difficult, and this is precisely the opposite of what we'd like to have happen. In fact, while teaching this method, I go out of my way to tell students that this method is really easy if you just follow the steps outlined on the next few pages, and that it's almost impossible to get the wrong answer! Child psychology 101: If you tell teenagers something over and over again, it becomes true.

To make this method seem easier, I refer to it as the "t-chart method", for obvious reasons. Here's what the t-charts look like, in case anybody uses a different method and wants to try this out:

A t-chart! 

Now that we're familiar with what the t-chart looks like, we have to explain to the kids how it's used to solve mole conversion problems. As you may have noticed, I don't like using big words in my writing or conversation. When I use big words, I get confused about what I was talking about in the first place. So do people I talk to. Use small words. They mean the same thing, and they're a whole lot easier to understand.

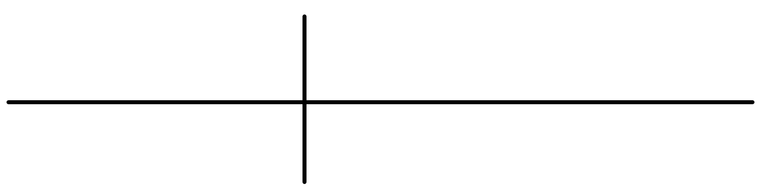
I've also found that students learn easier when presented with a list of things to do. Instead of making unit conversions a single, very large operation, a list of steps breaks the problem into smaller pieces that are more easily understood. The list of operations for unit conversions is as follows:

1) Draw the t!
2) Put the number that the problem gives you in the top left.
3) Put the units of what the problem gives you in the bottom right.
4) Put the units of what you're trying to find in the top right.
5) Stick the unit conversion factor in front of the units on the right.
6) Multiply the numbers on the top of the t-chart together, then divide by the number on the bottom. The unit for this answer is whatever unit is in the top right corner of the t.
7) For multistep problems, use the answer from your t-chart as the starting point for another t-chart.

Let's see how this method is used with an example:

Example: How many molecules of methane are there in 60 grams?

Step 1: Draw the t!



Step 2: Put the number that the problem gives you in the top left.

In this example, the only number provided by the problem is "60 grams". Put this in the top left quadrant of the t, as shown on the next page:

60 grams CH_4	

Step 3: Put the units of what the problem gives you in the bottom right.

Because "grams CH_4" is the unit that the problem gives you, place this in the bottom right quadrant of the t. You'll notice that this has the effect of cancelling the unit in the top left of the t.

60 grams CH_4	
	grams CH_4

Step 4: Put the units of what you're trying to find in the top right.

This step can get trickly. After all, we want to find the number of molecules of methane, but before we can do this, we first need to find the number of moles of methane present in the sample. As a result, the thing we're trying to find in this t-chart is the number of moles of methane, rather than the number of molecules.

60 grams CH_4	moles CH_4
	grams CH_4

Step 5: Stick the unit conversion factor in front of the units on the right.

The unit conversion factors for these sorts of problems is fairly easy. If you've got the unit "moles", **always** put 1 in front of it in a conversion factor (the exception to this is stoichiometry calculations, as we'll discuss later). If you've got the unit "grams", **always** put the molar mass of the compound in front of it in a conversion factor. If you've got the unit "molecules" or "atoms", **always** put 6.02×10^{23} in front of it. As a reminder to students, I've included a "Ten Commandments of Mole Calculations" worksheet which contains ten rules for doing unit conversions.

Once the appropriate conversion factors have been placed in the t-chart, it looks like this:

60 grams CH_4	1 moles CH_4
	16 grams CH_4

Step 6: Multiply the numbers on the top of the t-chart together, then divide by the number on the bottom. The unit for this answer is whatever unit is in the top right corner of the t.

In this case, we end up with an answer of 3.75 moles CH_4.

Step 7: For multistep problems, use the answer from your t-chart as the starting point for another t-chart.

Because this is a multistep problem, we'll need to do another calculation to go from moles to molecules. This problem will be solved in exactly the same way as we saw above, except that "3.75 moles" will be placed in the top left part of the t-chart in step 2. I will leave the t-chart making to you as an exercise. The answer, incidentially, is 2.26×10^{24} molecules CH_4.

Doing the Moles, Molecules, and Grams Lab

Equipment:

- 5 balances, though you can get away with fewer if necessity demands it
- 35 empty film canisters
- 35 grams of ethanol
- 35 grams of sodium chloride
- 35 grams of sand
- 35 mL of water
- 35 grams of sugar
- 7 weighing boats or watch glasses

This equipment should be enough for five lab groups, with two spare sets of chemicals in case of spills. Make sure the film canisters are labeled with an indelible marker or paper label.

Safety:

The ethanol should be kept away from open flames and other heat sources. The other chemicals used in this lab are generally harmless.

Room destruction factor:

The main danger in this lab is to your balances. If you use electronic balances, make sure your students clean up any spilled substances before they leak into the electronics and muck them up.

How the lab works:

In this lab, your students will weigh each sample and do calculations to determine how many moles and molecules are present in each. The trick here is in preweighing the samples to be as close to five grams as possible – your students will be surprised to find that the numbers of moles and molecules vary for each sample, even though the masses are the same.

Many of your students will have problems with these calculations, so be ready to jump in with helpful advice on making t-charts. My handy advice for this lab consists of "Follow the steps we discussed in class". This drives students crazy (expect to hear them complain that they're *trying* to follow the steps), but I've found students always eventually manage to figure out t-charts on their own. Many students are strongly math-phobic, and the best way for them to get over it is to simply force them to figure it out on their own. It's not much fun, but it does work.

Clean up:

When you have finished with the lab, all chemicals can be placed in the trash. The exception is ethanol, which should be disposed of in a marked organic waste container as required by your school system.

What can go wrong:

- After only a few weighings, you'll find that the film canisters no longer contain five grams of each compound. To compensate for the gradual loss of compound caused by spillage, add about half a gram to each film canister between classes to keep the weight near five grams.
- Your students have problems with the calculations. Before doing this lab, make sure you do plenty of unit conversion problems in class. By the time your students get to this lab, they should have a reasonably good idea of how to do these calculations.

Solutions to the Moles, Molecules, and Grams Lab

Prelab section: 0.63 moles / 3.8×10^{23} molecules

Lab section:

The data table should contain the name and formula of ech substance, as well as the molar mass and their measured mass. This information should be arranged in neat, easy-to-read columns and rows.

Calculations section:

All of the answers below assume that there are exactly 5.00 grams in each film canister. As the lab progresses, these values may vary by as much as 20%, depending on how is spilled.

- Moles of ethanol: 0.11 moles
- Molecules of ethanol: 6.54×10^{22} molecules
- Moles of salt: 0.085 moles
- Molecules of salt: 5.15×10^{23} molecules
- Moles of water: 0.28 moles
- Molecules of water: 1.67×10^{23} molecules
- Moles of sand: 0.083 moles
- Molecules of sand: 5.01×10^{22} molecules
- Moles of sugar: 0.015 moles
- Molecules of sugar: 8.80×10^{21} molecules

Post-lab questions:

1) There were more molecules of water than any other substance. Though ideally all of the masses for each compound should be identical, this will likely not be true after spills.
2) It's handy to know the number of moles of a substance when doing chemical reactions – after all, when molecules react, we talk about molecules interacting with each other, not grams.
3) 2.78 moles, 1.67×10^{24} molecules
4) 3.13 moles, 1.88×10^{24} molecules
5) Because the molar masses are different. One mole of a compound with a small molecular mass weighs less than one mole of a compound with a large molecular mass.

Solutions to the Moles, Molecules, and Grams Worksheet

1) 1.13 moles
2) 5.42×10^{23} molecules
3) 7.48 moles
4) 3.93×10^{21} molecules
5) 790. grams
6) 103 grams
7) 2.76×10^{23} molecules

Unit Conversion Ten Commandments

1) Thou shalt remember that 6.02×10^{23} of anything is equal to one mole.

2) Thou shalt not put any units after the number 6.02×10^{23} except "molecules" or "atoms". It is a grave sin to write "grams" after this unit!

3) Thou shalt not put any number other than "1" in front of the word "moles" in conversion factors. This, too, is a great sin.

4) Thou shalt not put a number smaller than 1.0×10^{19} or greater than 1.0×10^{26} in front of the word "atoms" or "molecules", as this is an unreasonable quantity of chemical.

5) Thou shalt not put a number smaller than 0.0001 or greater than 10,000 in front of the word "grams", as this is an unreasonable quantity of chemical.

6) Thou shalt always make sure that all units cancel in a conversion except for the one in thine answer.

7) Thou shalt always put units in thine final answer. Failure to do so will bring the wrath of Teacher upon thee.

8) Thou shalt always show thine work.

9) Thou shalt stay after school when thou dost not understandeth conversions.

10) Thou shalt study unit conversions, using the problems in thine homework.

Moles, Molecules, and Grams Handout

In class, we learned how to convert between moles, molecules, and grams. The big question that most people have is "Why do we need to learn this?"

Measuring the mass of verious objects is easy – all you have to do is place the object on a balance and read the display. Unfortunately, in chemistry, we're not usually concerned with the mass of things, but with the numbers of atoms or molecules present. Mass doesn't directly tell you either of these – instead, you have to convert to moles of molecules, both of which are measurements of how many particles are present.

To make these conversions, use the chart below:

$$\textbf{Grams} \quad \xrightarrow{\hspace{6em}\textit{Molar mass}\hspace{6em}} \quad \textbf{Moles} \quad \xrightarrow{\hspace{6em}\textit{6.02 x 10}^{23}\hspace{6em}} \quad \textbf{Molecules}$$

Think of this chart as a road map, where "Grams", "Moles", and "Molecules" are city names. Think of "Molar Mass" and "6.02 x 10^{23}" as the names of highways between the cities.

Let's say we want to convert from moles to grams. Using the chart, we can see that if we want to go from "Moles" to "Grams", we need to take the road labeled "Molar Mass". This tells you that the conversion factor you use to go from moles to grams is the molar mass of the compound you're working with.

You'd do a similar calculation going the other direction, or going from moles to molecules. If you do a two step calculation, such as "Molecules" to "Grams", you'd need two conversion factors, one for each step. After all, if you're traveling to a faraway city, you may need to travel halfway on one road before switching to another.

The most important thing you can do to make sure you solve these problems correctly is to work slowly and carefully. **Do not plug numbers in at random**, because you'll get them wrong! Use common sense and the chart above, and you'll never get lost in your unit conversions!

Moles, Molecules, and Grams Lab

As we've discussed in class, it's easy to make conversions between moles, molecules, and grams. For example, if we want to go from moles to grams, we use the molar mass to make this conversion. If we want to go from moles to molecules, we use Avogadro's number, or 6.02×10^{23}. Finally, if we want to go from grams to molecules, we use a two step process where we first convert from grams to moles, then from moles to molecules.

In this lab, we will be weighing samples of five different substances, then finding out how many moles and molecules of each one are present.

Prelab:

If you weighed 25 grams of sodium hydroxide in this lab, how many moles of sodium hydroxide would you have? How many molecules?

Lab:

In this lab, your job will be to weigh five substances and determine the number of moles and molecules present in each sample. You may find the following information handy during this lab:

Substance	Chemical Name	Formula
alcohol	ethanol	C_2H_6O
salt	sodium chloride	$NaCl$
sand	silicon dioxide	SiO_2
water	water	H_2O
sugar	sucrose	$C_{12}H_{22}O_{11}$

Data Table:

Make a data table that contains the name and formula of each substance, the molar mass of each substance, and the mass in grams of each substance. It is important that your data table is complete, as you will need this information to find the number of moles and molecules of each substance in the next section!

Calculations: Using your data, find the following values

Number of moles of ethanol: ______________________________

Number of molecules of ethanol: ______________________________

Number of moles of salt: ______________________________

Number of molecules of salt: ______________________________

Number of moles of water: ______________________________

Number of molecules of water: ______________________________

Number of moles of sand: ______________________________

Number of molecules of sand: ______________________________

Number of moles of sugar: ______________________________

Number of molecules of sugar: ______________________________

Post Lab Questions:

1) Which of the molecules had the largest number of molecules? Was it the one with the largest mass? Explain.

2) Can you think of a case where it might be handy to know the number of moles you had of a substance? Explain.

3) Water has a molecular formula of H_2O. If I have 50 grams of water, how many moles of water do I have? How many molecules?

4) Natural gas has a formula of CH_4. If I have 50 grams of natural gas, how many moles of natural gas do I have? How many molecules?

5) I had the same weight of water and natural gas in problems 3 and 4. Why didn't the answers come out the same? Explain.

Moles, Molecules, and Grams Worksheet

Using your knowledge of mole calculations, do the following conversions:

1) 45 grams of NaOH to moles

2) 0.90 moles of AgF to molecules

3) 4.5×10^{24} molecules of CH_4 to moles

4) 1.98 grams of $PbSO_4$ to molecules

5) 1.98 moles of $Pb(SO_4)_2$ to grams

6) 8.4×10^{23} molecules of Li_2CO_3 to grams

7) 39 grams of $NaNO_3$ to molecules

Chapter 7 – Percent Composition Lab

Overview

The term "percent composition" refers to the analysis of a chemical compound to determine how much of each element is present. For example, if we wanted to find the percent composition of hydrogen in water, we'd calculate the percentage of mass in water that is due to the mass of hydrogen atoms. In this chapter, students will learn how to calculate percent compositions of various chemical compounds. In the lab, they will use this knowledge to find the percent composition of water in magnesium sulfate heptahydrate.

Teaching about Percent Composition

Percent composition is an easy concept to teach, but sometimes causes problems when students attempt to do calculations with real data. The problem is not the concept – students have no difficulties with that. The problem is that many students can't do the math.

For this reason, we have to sneak the math into this lesson in an intuitive way. Students have a hard time visualizing the percent composition of chlorine in calcium chloride. They don't, however, have trouble visualizing what percent of their feet have the big toe on the left, what percent of the students in the room are girls, and what percent of people are left-handed. These problems are simple, and students already know how to solve them.

After students solve the mysteries of the above problems, try fitting them into the equation for percent composition:

$$\textbf{Percentage composition of X} = \frac{\textbf{mass of element X}}{\textbf{molar mass of the compound}} \times 100\%$$

By this point, the equation explains itself. If any students still have problems with it, go through the calculation to determine the percent of girls in the classroom. This should clear up questions about how to use the equation, and the jump from "percentage of girls" to "percent composition" is a small one.

Before doing this lab, make sure you spend some time talking about the formulas of hydrates. Remember, the formula of a hydrate is written as the formula of the anhydrous compound, followed by a dot and the number of water molecules that are present. For example, copper (II) sulfate pentahydrate is $CuSO_4 \cdot 5\ H_2O$. To find the number of complexed water molecules in the lab, simply divide the number of moles of water removed (which can be determined by measuring the mass of water removed) by the moles of anhydrate present after heating. This can be a difficult calculation to conceptualize, so make sure you go over it before starting the lab.

Doing the Percent Composition Lab

Equipment:

- 6 crucibles with lids
- 100 grams of magnesium sulfate heptahydrate, more commonly known as Epsom salts.
- 6 Bunsen burners
- 6 ring stands
- 6 clay triangles
- 6 crucible tongs
- A wall clock with a second hand, or 6 stopwatches
- 2 flint strikers
- 6 small disposable pipets (plastic and glass are both fine)
- 2 balances

Safety:

The most significant safety hazard in this lab is the heat generated by the Bunsen burner. During the course of this experiment, the crucible, crucible tongs, crucible lid, clay triangle, ring stand, and Bunsen burner will be extremely hot. The students should be warned that hot glassware looks the same as very cold glassware, and that the best way to test to see whether glassware is hot is to splash a small drop of water on it to see if it sizzles. If students suffer burns during this lab, they tend to be minor because our our natural tendency to pull away from very hot objects. However, if there is any doubt as to the seriousness of a burn, seek medical attention immediately!

Students should not use rubber gloves during this lab as they melt to the skin at high temperatures. A good rule of thumb: If something's too hot to touch with your bare hands, it's too hot to touch with rubber gloves!

As always, students should wear goggles during this lab. Though Epsom salts are generally harmless, the crystals jump around when heated, causing a very slight eye hazard. If students get magnesium sulfate heptahydrate in their eyes, flush them thoroughly with water and seek medical attention.

Room destruction factor:

Crucibles are very low-impact items. As a result, they are broken in great quantity during this lab, making it potentially very expensive. In case you've got many more lids than crucibles, keep in mind that the lids are made of the same material and can also be used to heat small amounts of compound when turned upside down.

How the lab works:

During this lab, students will find the weight of a hydrate before heating and compare it to the mass after heating. The difference in weight will be equal to the weight of water that was removed. Inexplicably, this calculation sometimes gives the students trouble, so make sure you describe what's going on before you start the lab.

In step three, students will sometimes go to great lengths to put exactly 10.00000 grams of magnesium sulfate in the crucible. The weighing process can be speeded hugely by telling them that anything between 9.5 and 10.5 grams is fine.

For clean up, the magnesium sulfate (both heptahydrate and anhydrous) can be put down the sink, as it's fairly harmless.

What can go wrong:

- "My Bunsen burner won't light." Common cause: The gas isn't on, the striker has no flint, or the burner is set with too much oxygen entering the bottom. Most likely, your students just need practice using the striker.
- The crucible breaks. When this occurs, issue the students a new crucible and have them start over. If the crucible is dropped and the magnesium sulfate can be easily recovered, have them sweep it off the lab bench into a watch glass for weighing.
- Students don't know when the heating process is complete. When magnesium sulfate hepahydrate is heated, it appears to foam up as the water is removed. When this foaming has stopped, a white, crusty lump is left in its place. As long as students have reached this stage, extra heating won't have any effect on their results.

Solutions to the Percent Composition Lab

Lab procedure section:

2, 3) Units should be given in grams.
4) Hydrated magnesium sulfate makes medium sized, colorless crystals.
5) The answer should be between 9.5 and 10.5 grams.

7) Units should be given in grams.
8) Between 4.64 grams (if the starting mass was 9.5 grams) and 5.13 grams (if the starting mass was 10.5 grams).
9) The anhydrous magnesium sulfate looks like a compressed white lump of powder.

Post-lab questions:

1) The mass of water removed is equal to the mass from step 5 minus the mass from step 8. This number should be between 4.86 and 5.37 grams, though it may be slightly lower if the hydrate was heated for too short a period of time, a common source of laboratory error.
2) The percentage of water is equal to the mass of removed water divided by the mass of the hydrate before heating, times 100. Ideally, this number should be 51.1%, though will likely be smaller due to the incomplete removal of water from the hydrate. Answers between 45-50% are reasonable.
3) The number of moles of water removed is equal to the mass of water removed (answer from #1) divided by the molar mass of water (18 grams). This value should be between 0.27-.30 moles. The number of moles of magnesium sulfate anhydride is equal to the mass of magnesium sulfate anhydride after heating (step 8) divided by the molar mass of magnesium sulfate anhydride ($MgSO_4$, 120.4 g/mol). The answer should be between 0.0385 – 0.0426 moles. The empirical formula of magnesium sulfate hydrate will be $MgSO_4 \cdot x\ H_2O$, where x is the number of moles of water divided by the number of moles of anhydrate. Ideally, this should be 7, as the correct formula is $MgSO_4 \cdot 7\ H_2O$.
4) In this lab, students should have noticed that hydrates and anhydrates have different appearances. These differences in appearance are frequently used to detect moisture in a room, as the appearance of the anhydrate will change to reflect the presence of atmospheric water.

Solutions for the Percent Composition Worksheet

1) 50%
2) 75%
3) 67%
4) 2%
5) 35%

Percent Composition Handout

When we calculate percent composition, we're really trying to determine how much of the compound's weight is due to each of the elements present. For example, if I were to ask you to find the percent composition of oxygen in water, I'm really asking you to find the percent of the total weight of water that's due to oxygen atoms.

To solve this problem, use the following procedure:

1) **Assume you have one mole of the substance you're analyzing.**

   This assumption makes the problem easier to solve. In our case, we'll just assume we have one mole of water.

2) **For each element, multiply the atomic weight of that element by the number of atoms of that element in each molecule. Keep these numbers separate from each other for now.**

   For hydrogen, multiply the atomic weight of hydrogen (1 g/mol) by the number of atoms of hydrogen per water molecule (2) to find 2 g/mol.

   For oxygen, multiply the atomic weight of oxygen (16 g/mol) by the number of atoms of oxygen per water molecule (1) to find 16 g/mol.

3) **Add these numbers together. This gives you the molar mass of the compound.**

   $$2 \text{ g/mol} + 16 \text{ g/mol} = 18 \text{ g/mole}$$

4) **Divide the mass of the element you're looking for by the molar mass of the compound. Multiply this number by 100 to find the percent composition.**

   The mass of oxygen from step 2 is 16 grams/mole, and the molar mass of water from step 3 was 18 grams/mole. Dividing these numbers, we get:

   $$(16 \text{ g/mol}) / (18 \text{ g/mol}) = 0.89$$

   When we multiply 0.89 by 100, we find that the percent composition of oxygen in water is **89%**.

Percent Composition Lab

Introduction:

You've probably noticed that some consumer goods contain a small packet labeled "Silica gel: Do not eat". What's that packet for, anyway?

As you know, many fragile goods can be easily damaged by moisture. The silica gel in each packet is used to soak up water from the atmosphere. This minimizes moisture that causes damage during shipping.

Many ionic compounds can be used to soak up water. Before they absorb water, they're referred to as "anhydrous", which means "without water". After they've soaked up the maximum amount of water, they're called "hydrates", making that water molecules are stuck to them. If you heat hydrates to very high temperatures, they "dehydrate", meaning that the water is lost. Once all of the water is lost, these compounds are again referred to as "anhydrous".

In this lab, we will be dehydrating the hydrate of magnesium sulfate. Using the data from this lab, you will determine the percent composition of water and empirical formula of the hydrate.

Lab Procedure:

1) Obtain a clean crucible, crucible lid, crucible tongs, and Bunsen burner.

2) Find the weight of the crucible and crucible lid:

   Weight of the crucible and lid: ____________________________

3) Using a clean spatula, add about 10 grams of hydrated magnesium sulfate to the crucible. Put the cover on the crucible, and weigh the crucible, cover, and magnesium sulfate hydrate.

   Weight of the crucible, lid, and hydrate: ____________________________

4) Describe the appearance of the hydrated magnesium sulfate crystals below:

5) Using the data from steps 2 and 3, calculate the mass of hydrated magnesium sulfate in the crucible:

Mass of hydrated magnesium sulfate: ____________________________

6) Prepare a ring stand with a clay triangle in the ring. Place the crucible in the clay triangle and heat over the Bunsen burner for 10 minutes.

7) Remove the crucible from the clay triangle with the crucible tongs and let it cool on the countertop for five minutes. With the crucible tongs (it may still be hot!), place the crucible and contents on the balance and find its mass.

Mass of crucible and anhydrate after heating: ___________________

8) Using your data from steps 2 and 7, calculate the mass of the anhydrous magnesium sulfate:

Mass of anhydrous magnesium sulfate: ____________________________

9) Remove the lid from the crucible and look at the anhydrous magnesium sulfate. Describe its appearance here:

10) Clean up: The anhydrous magnesium sulfate can be washed down the sink once it has cooled.

Post-lab questions: (Show all work!)

1) Using the data from steps 5 and 8, calculate the mass of water that was removed when you heated the crucible:

2) What was the percentage by mass of water in the original sample?

3) Calculate the number of moles of water you removed from the magnesium sulfate hydrate, as well as the number of moles of magnesium sulfate anhydrate you formed. Use the results of these calculations to determine the empirical formula of the hydrate.

- *Moles of water removed from the magnesium sulfate hydrate:*

- *Moles of magnesium sulfate anhydrate formed:*

- *Empirical formula of the hydrate:*

4) In steps 4 and 9, you saw what both the hydrated magnesium sulfate and anhydrous magnesium sulfate crystals looked like. Do you think it's possible to use the appearance of hydrates to determine if there is moisture in a room? Explain.

Percent Composition Worksheet

1) What is the percent composition of sulfur in sulfur dioxide (SO_2)?

2) What is the percent composition of carbon in methane (CH_4)?

3) What is the percent composition of oxygen in lithium hydroxide (LiOH)?

4) What is the percent composition of hydrogen in sulfuric acid (H_2SO_4)?

5) What is the percent composition of nitrogen in ammonium nitrate (NH_4NO_3)?

Chapter 8 – Ionic and Covalent Compounds

Overview

The two main types of chemical compounds are ionic and covalent compounds. Ionic compounds are formed when atoms donate or take electrons from one another, forming cations and anions that stick to each other. Covalent compounds are formed when two atoms with similar electronegativities share pairs of electrons. The properties of each type of compound are different, and can be used to tell the two types of compounds apart. In this lab, students will use experimental observations to determine whether unknown compounds are ionic or covalent.

Teaching about Ionic and Covalent Compounds

Ionic and covalent compounds are covered extensively in chemistry. After all, most of the chemistry we study involves chemical compounds. However, students sometimes have difficulty telling what type of compound will be formed when two elements react with one another.

The main thing that determines whether a compound will be ionic or covalent is the electronegativity of its substituent elements. Electronegativity, as you're probably already aware, is a measure of how much an element wants to pull electrons away from an element that it has bonded to. Chemical compounds in which the two elements have very dissimilar electronegativities are ionic, because the electronegative element pulls valence electrons completely off the electropositive element. Chemical compounds in which the two elements have similar electronegativities are covalent, because neither element wants to give up any of its electrons.

I use this analogy in class: Consider a barbeque. Let's say that you really like to eat meat, and you're sitting next to a vegetarian. Each of you has been given a hamburger. To make both of you happy, the vegetarian will give you his hamburger. You've gotten what you want because you love meat, and the vegetarian has gotten rid of his meat. This makes you friends. Clearly, this illustrates how an ionic compound is formed. Substitute "electronegative" for "like to eat meat" and "electron" for "hamburger", and this describes what happens during the formation of an ionic compound.

On the other hand, let's consider a case where you're at a barbeque and there's only one hamburger left. You love meat, as does the guy sitting next to you. If you ate the hamburger, the other guy would be unhappy with you. If the other guy eats the hamburger, you'll be unhappy. As a result, the only way that both of you can be happy is if you share the hamburger. This illustrates how a covalent compound is formed, as two electronegative elements have to share electrons to make both of them stable.

Generally, ionic compounds are formed when metals bond with nonmetals, and covalent compounds are formed when two nonmetals bond with one another. (For the most part, metalloids are treated as nonmetals when determining how they'll bond.) The reason for this is the octet rule, which says that "all elements want to gain or lose electrons to be like the nearest noble gas". Because the electron configurations of the nearest noble gases to metals requires that they lose electrons, metals are electropositive and form cations. Because the electron configurations of the nearest noble gases to nonmetals require that they gain electrons, nonmetals are electronegative and form anions. If both elements are nonmetals, they need to share electrons to make both believe they have the same number of electrons as the nearest noble gas.

You may have noticed that I tend to give elements human traits, saying that they "like" to do one thing or another, or that doing something makes them "happy". At the college level, nothing annoys most professors quite as much as hearing students explain that fluorine "likes" to gain electrons – after all, fluorine isn't sentient and can't think anything at all. They prefer to explain that fluorine's energy is minimized by gaining one electron, and that the formation of NaF when sodium and fluorine bond creates an energetically favorable situation.

Of course, they are entirely correct. Atoms don't think about anything and have no feelings whatsoever. However, I've found that most high school chemistry students have a difficult time imagining energetic concepts, as thermodynamic concepts are dealt with primarily in physics classes. By teaching students that atoms "want" to do something, the main idea will become clear in their minds. Once they understand how the various elements react with one another, you can hit them with an explanation of how energy plays a part in chemical reactivity.

Now that your students know what the difference is between an ionic and covalent compound, and why each is formed, the next question is this: How can you tell them apart from each other in the laboratory?

Each type of compound has characteristics which distinguish them from the other. Keep in mind that these characteristics are *general* characteristics, and may not work for all chemical compounds. However, by testing each of them, your students should get a "big picture" idea of what type of compound is present. A list of the general properties of each type of compound is given to you below:

1) Ionic compounds have high melting and boiling points, while covalent compounds have low melting and boiling points.

This is, by no means, set in stone. However, when you examine ionic compounds, they generally have much higher melting points than covalent compounds, typically over 100^0 C, and frequently over 300^0 C. On the other hand, most covalent compounds have a melting point less than 100^0 C, while some are well below freezing. When researching this property, I selected ten ionic and ten covalent compounds at random from the CRC Handbook of Chemistry and Physics and found that the average melting point of the ionic compounds I picked was 392^0 C, with six of them over 100^0 C and two over 1000^0 C. The covalent compounds, on average, had a melting point of 26^0 C, with three under 0^0 C and eight of them under 50^0 C.

I usually tell my students that if a compound melts after less than 30 seconds of heating over the Bunsen burner, this is an indication that it may be a covalent compound. Be sure to tell them that this isn't absolute proof – just one piece of evidence for making a final determination.

2) Ionic compounds conduct electricity when dissolved in water or molten. Covalent compounds never conduct electricity.

The reason that ionic compounds conduct electricity when dissolved in water or molten is that the cations and anions can move around easily, and it's the motion of ions that causes electricity to be conducted through the liquid. Solid salts, though they contain ions, do not conduct electricity as the ions are locked into place. Covalent compounds don't conduct electricity because there are no ions to move the electrons.

Students will frequently point out that water is a covalent compound that conducts electricity well. You can dispel this myth with a simple demonstration. Immerse a conductivity tester (see the equipment list for more information) in distilled water. It won't conduct electricity, as the concentration of ions in distilled water is only 2.0×10^{-7} M (from the autoionization of water itself). Next, have each student in your class swirl their finger around in the water and try it again – you'll find that the water magically conducts electricity. This newly-found conductivity arises from the dissolved salts on your students' hands, not the water itself. Tap water conducts electricity because there are minerals dissolved in it.

It's important that you make it clear to students that ionic compounds do not conduct electricity when they are solids. I've found that students sometimes believe that ionic compounds conduct electricity as well as metals, even though they are excellent insulators. You can demonstrate this with the conductivity tester by checking the conductivity of a nail, followed by the conductivity of a salt crystal.

You should also explain the idea of solubility to your students before doing this lab. Though ionic compounds conduct electricity when dissolved in water, ionic compounds will not cause water to become conductive if they don't dissolve. Demonstrate this by testing the conductivity of a calcium hydroxide solution – it won't conduct electricity because calcium hydroxide is very nearly insoluble in water.

3) Ionic compounds rarely burn, while covalent compounds burn more frequently.

Most covalent compounds are organic, and most organic compounds burn. As a result, most covalent compounds tend to be flammable. On the other hand, most ionic compounds do not contain organic groups, so they don't burn.

Of course, there are exceptions to this rule. Methylene chloride (CH_2Cl_2) doesn't burn and carbon dioxide is used in many fire extinguishers. Some ionic compounds that contain large organic anions burn (soaps, for example) and some ignite spontaneously in air. These examples, however, are fairly uncommon, making flammability a good indicator of compound type.

Many ionic compounds decompose at high temperatures, which is distinctly different than burning. An example of this is calcium carbonate, which decomposes to form calcium oxide and carbon dioxide. This shouldn't cause a problem during this lab, as these compounds are not commonly found in high school chemistry labs.

4) Ionic compounds are hard and brittle, while covalent compounds have a wide variety of textures.

Almost without exception, ionic compounds are hard and brittle. This stems from the strong electrostatic interactions between the ions in ionic crystals. These interactions are strong enough that it is difficult to separate ions from one another (causing them to be hard), and when you do, it causes the entire crystal to fracture (causing them to be brittle). Covalent compounds, on the other hand, are frequently held together by intermolecular forces such as dipole-dipole forces or hydrogen bonds, causing much weaker interactions. As a result, covalent compounds are generally not as hard as ionic compounds, nor does a dislocation cause the entire structure to disintegrate. Similarly, amorphous covalent solids (such as rubber, plastic, etc.) are, for the most part, flexible and shatter-resistant.

Of course, there are many exceptions to this rule for covalent compounds. Covalent solids such as sucrose (table sugar) form both hard and brittle crystals, as do amorphous solids such as glasses. Network atomic solids also tend to have many of the same properties as ionic compounds, particularly those of being hard and brittle. However, because covalent compounds are far more likely to be soft and flexible than ionic compounds, this rule is useful when running across a flexible compound.

5) Ionic compounds form crystals, while covalent compounds have many different appearances.

Ionic compounds form crystals to maximize the interactions between cations and anions. After all, crystals won't be stable if two anions are right next to one another – these interactions are best minimized by creating a crystalline stacking pattern. Covalent compounds, on the other hand, less frequently arrange into crystals, as they are held together by a variety of different means. Sucrose and silicon, for example, form very regular crystals, while rubber and wood do not.

Doing the Ionic and Covalent Compound Lab

Equipment:

- 100 grams of two ionic compounds. Good examples for this lab are sodium chloride, sodium acetate, potassium chloride, copper (II) chloride, and lead iodide. Nitrates should not be used, as they explode at high temperatures.
- 100 grams of two covalent compounds. Good examples for this lab are sucrose (table sugar), water, agar, plastic, and wax.
- 4 small screw-top storage bottles, labeled "Unknown 1", "Unknown 2", etc.
- 6 crucibles
- 2 Bunsen burners
- 1 flint striker
- 1 conductivity tester. A multimeter can be used for this with great success. If you don't have a multimeter, you can easily make a conductivity tester by creating a circuit consisting of a battery, a light bulb or LED, and two bare wires. When the wires are placed in a conducting solution, the bulb will light. If you're unsure of how to make your own conductivity tester, ask the guys at Radio Shack for help. They live for this kind of thing.
- 1 dissecting microscope for determining if very small particles of powder are crystalline.
- 5 liters of distilled water (for dissolving the solids)
- 5 – 50 mL beakers
- 5 spatulas
- 2 crucible tongs
- 2 ring stands
- 2 rings

- 2 clay triangles
- 1 hammer

Safety:

The Bunsen burners in this lab create a fire hazard for several reasons. Clearly, the Bunsen burner and ring stand apparatus will be hot after use, as will the crucible tongs. However, an additional fire hazard may be created from whatever compounds are being used for the lab. Though it is fine to use compounds that burn when exposed to heat, make sure your students use only very small quantities of mildly flammable materials. For this reason, nitrates, volatile liquids, and other potentially explosive compounds should not be used as unknowns in this lab. If you're unsure about the safety of your unknowns, *don't use them!*

Similarly, the compounds used in this lab should not be hazardous for any other reason. Sulfuric acid, sodium hydroxide, and sodium hydride are all bad choices for unknowns because of their extreme reactivity. Sodium cyanide, warfarin, and other strong poisons should not be used either. Again, if you're unsure about the safety of your unknowns, *don't use them!*

Because heat is being used to test the melting and boiling points of these compounds, make sure your students don't wear gloves during this lab, as they may melt to the skin, causing very severe burns.

Students should wear goggles at all times during the setup, performance, and clean-up of this lab.

Room destruction factor:

The potential for room destruction is low, but the potential for irritation is high. For example, I once used an old bottle of yeast as one of the covalent compounds, creating a terrible smell when it was burned. For this reason, it's probably wise to consider the smells these compounds will make when burned before doing the lab.

How the lab works:

Armed with the properties of ionic and covalent compounds, your students will attempt to determine whether each of the four unknowns are ionic or covalent based on whatever tests they can devise. Typical tests include hitting the compound with a hammer (to test hardness and brittleness), heating the compound (to test melting and boiling point, and to see if the compounds are flammable), dissolving the compound (to test the conductivity in water), and examination of the compound under a dissecting microscope (to determine if fine powders are crystalline).

Some students may also come up with more unconventional tests. These tests rarely have any diagnostic value, but are also rarely dangerous. In the case that you see students performing unusual tests, allow them to continue unless the experiment is blatantly dangerous, as it shows that your students are using critical thinking to solve this problem.

When they have finished with this lab, your students will determine whether each compound is ionic or covalent based on the results of their tests. Make sure they know ahead of time that some of the tests may be contradictory due to the fact that the properties discussed in class are general properties rather than absolute law. The results to these tests and their final answers will be written on their lab sheet.

<u>What can go wrong:</u>

- Your students heat the unknown compounds in the 50 mL beakers rather than the crucibles. Make sure you have several beakers on hand, as this invariably causes beakers to crack or break.
- Covalent compounds sometimes conduct electricity when dissolved in water. The reason for this is not that the rule is faulty – it's almost completely foolproof. Covalent compounds sometimes appear to conduct electricity because they contain small quantities of ionic impurities which cause them to conduct. To avoid this problem, test each unknown before starting the lab to make sure that the covalent compounds don't conduct when dissolved in water. Another reason that covalent compounds may conduct electricity is that students use tap water rather than distilled water when they make their solutions.

<u>Clean up</u>:

To make clean up easy and save money on waste disposal, pick unknowns that can be safely put down the sink or thrown away.

Solutions for the Ionic and Covalent Compound Lab

There is no real answer key for this lab, as you will decide for yourself which unknowns will be used. However, there should be a data table showing the results of the tests performed on each unknown, followed by an identification of the unknown itself. Typically, I will give full credit if the answer is correct, and half credit if the answer is incorrect but supported by the results of each test.

Post-lab question:

It is far more common to have conflicting data about whether a compound is ionic or covalent than to have unambiguous data. Though we've described general properties of ionic and covalent compounds, not all ionic and covalent compounds have all of the properties mentioned. As a result, we have to do several tests to make a definitive conclusion, rather than relying on only one.

Solutions for the Ionic and Covalent Compound Worksheet

1) In an ionic bond, one atom donates electrons to the other. In a covalent bond, both atoms share their electrons equally.
2) The greater the difference in electronegativity, the more likely a compound is to be ionic.
3) If a metal bonds with a nonmetal, a ionic compound will be formed. If two nonmetals bond, a covalent compound will be formed.
4) Sodium citrate has the formula $Na_3C_6H_5O_7$. Because both carbon and hydrogen are present, sodium citrate is flammable.
5) The salts in your skin would dissolve when you get in the bathtub, causing the distilled water to begin conducting electricity.

Ionic and Covalent Properties Handout

Ionic and covalent compounds have very different properties. By understanding these differences, you can determine whether a compound is ionic or covalent by performing just a few tests. These are the most important of these properties:

1) **Ionic compounds have high melting and boiling points, while covalent compounds have low melting and boiling points.**

   Because of the type of bonding involved, ionic compounds generally have melting points over 200^0 C, while covalent compounds usually have much lower melting points. Similarly, the boiling points of ionic compounds are also usually much higher than those of covalent compounds.

2) **Ionic compounds conduct electricity when dissolved in water or molten. Covalent compounds never conduct electricity.**

   In order for electricity to move through a solution, ions need to be able to move around. As a result, ionic compounds conduct electricity when dissolved in water or when melted, but *not* as solids. On the other hand, covalent compounds *never* conduct, because no ions are present.

3) **Ionic compounds rarely burn, while covalent compounds burn more frequently.**

   Chemical compounds usually burn because they contain carbon and hydrogen atoms. Ionic compounds rarely contain these elements, while many covalent compounds (called "organic" compounds) do. As a result, covalent compounds are more likely to burn than ionic compounds.

4) **Ionic compounds are hard and brittle, while covalent compounds have a wide variety of textures.**

   Ionic compounds are hard and brittle because the ions are locked tightly into place by their magnetic interactions. As a result, it's difficult to move these ions apart, and when they do move apart, the whole crystal typically breaks. Covalent compounds are less likely to have these strong interactions, making them less hard and brittle than ionic compounds.

5) **Ionic compounds form crystals, while covalent compounds have many different appearances.**

   The ions in ionic compounds line up with one another so that all ions are near oppositely charged ions, because cations and anions have opposite charges. By stacking these ions in a regular pattern, the opposites attract one another and a crystal is formed. In covalent compounds, there's no need to stack the molecules like this because there are no ions to attract or repel one another. As a result, they less frequently form crystals.

Ionic and Covalent Compound Lab

In class, we've discussed how to tell ionic and covalent compounds apart from one another by their chemical and physical properties. In this lab, you'll get a chance to test your knowledge by identifying four real unknown compounds.

Your job is simple: Determine whether each unknown is ionic or covalent. To do this, you may do any tests you'd like, as long as they don't violate any lab safety rules. In each section below, you should make a data table which describes the tests you performed and their results. At the end, you will use your data to determine whether each unknown is ionic or covalent.

Good luck!

Unknown 1:

Data Table:

Unknown 1 is a ______________________ compound.

Unknown 2:

Data Table:

Unknown 2 is a ______________________ compound.

Unknown 3:

Data Table:

Unknown 3 is a ______________________ compound.

Unknown 4:

Data Table:

Unknown 4 is a ______________________ compound.

Post-lab question:

In this lab, you probably found that you had conflicting data about whether a compound was ionic or covalent. Do you believe that this sort of conflicting data is common for most chemicals, or do you think that the compounds you used are unusual? Explain.

Ionic and Covalent Compound Worksheet

1) What is the difference between an ionic and covalent bond?

2) What is the relationship between electronegativity and bond type? Put another way, why could we predict whether a compound would be ionic or covalent if given a list of the elements' electronegativities?

3) How can you tell whether a compound is ionic or covalent simply by looking at its formula?

4) Most ionic compounds don't burn. However, sodium citrate does. Give a possible explanation for sodium citrate's flammability.

5) Water is a covalent compound and doesn't conduct electricity. With this in mind, explain why you would electrocute yourself if you took a bath with your hair dryer, even if you took a bath is perfectly pure distilled water.

Chapter 9 – Six Types of Chemical Reaction

Overview

All chemical reactions can be placed into six basic types: synthesis, decomposition , single displacement, double displacement, combustion, and acid-base. In this chapter, students will be learning the six types of reaction and how to classify all chemical reactions into one of these categories. In the lab, students will explore one example of each type of reaction.

Teaching about the six types of reaction

Chemical reactions tend to fall naturally into six categories: syntheis, decomposition, single displacement (replacement), double displacement (replacement), combustion, and acid-base reactions. Depending on the textbook you use, there may be other categories of chemical reaction discussed as well. Redox reactions are sometimes treated as a seventh sort of reaction, though most can also be classified as one of the six types above. Nuclear reactions are also sometimes treated as a type of chemical reaction – though they are undoubtedly reactions, they aren't chemical reactions as the changes that take place are in the nucleus of an atom, rather than between two atoms or chemical compounds.

Students sometimes dislike learning about the six types of chemical reaction because it can be a bit boring. The best ways to teach students the six types of reaction are to give them interesting examples of each type of reaction, and follow up with a step-by-step process they can use to determine what type of reaction is taking place.

I have included for you some interesting examples of each type of chemical reaction below, with some background information to make them more interesting:

Synthesis reaction: The explosion of the Hindenburg is a famous example of the synthesis reaction between hydrogen gas and oxygen gas. The equation for this reaction is:

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)}$$

This is an extremely exothermic reaction that is commonly demonstrated in high school classrooms by filling a very small balloon with hydrogen gas and holding a candle under it. When the balloon pops, the hydrogen mixes with atmospheric oxygen, and the flame causes the reaction to take place. Though it's exciting in person, I recommend showing a video of this demonstration rather than actually performing it because of the extreme risk of explosion.

Decomposition reactions: An interesting real-world use of decompositions takes place when the sodium azide in airbags is electrically stimulated to decompose into nitrogen gas and sodium nitride:

$$2\ NaN_{3(g)} \rightarrow Na_3N_{(s)} + 4\ N_{2(g)}$$

Many decompositions involve the very fast release of large quantities of energy, making them difficult to demonstrate in a high school chemistry course. However, dehydration reactions such as the one performed in the percent composition lab (Chapter 7) are also classified as decomposition reactions.

Single displacement reactions: A very common (and very interesting) example of single displacement reactions occurs when copper wire is placed in dissolved silver nitrate to form crystals of silver and dissolved copper (I) nitrate:

$$Cu_{(s)} + AgNO_{3(aq)} \rightarrow Ag_{(s)} + CuNO_{3(aq)}$$

This demonstration is not only interesting because solid silver is formed, but because the clear silver nitrate solution gradually turns blue as copper (I) nitrate is formed.

Double displacement reactions: The classic method for determining the amount of chloride ion in a compound is to react the chloride with silver nitrate to precipitate silver chloride. The silver chloride is then weighed, allowing the original amount of chloride to be determined. This demonstration can easily be demonstrated by combining silver nitrate with sodium chloride in water:

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

Acid-base reactions: An everyday example of acid-base reactions is when Rolaids™ are used to relieve heartburn:

$$Mg(OH)_{2(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + 2\ H_2O_{(l)}$$

Combustion reactions: Anytime something organic burns, a combustion reaction is taking place. An everyday example is the combustion of the propane in a barbeque grill:

$$C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$$

Doing the Types of Reactions Lab

Because this lab has six stations, each will be treated explained in separate sections below. This lab is meant to be run with six different lab groups rotating at intervals of 10-15 minutes. Each station is a self-contained mini lab, so it doesn't matter where each group starts. It goes without saying that goggles must be worn during all stations of this lab.

Station 1: Synthesis reaction

$$2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$$

Equipment:

- Bunsen burner
- Crucible tongs
- 1 meter of magnesium ribbon, cut into 4 centimeter lengths

Safety:

Magnesium burns at a very high temperature, necessitating a higher level of fire safety than is usual. Make sure that the only students near this lab station are the ones actually performing it, and that they're well aware of proper fire safety procedures.

In the very unlikely event that a larger magnesium fire starts, *do not put attempt to put it out with water* as it will only make the fire much worse. Evacuate the room and let trained professionals deal with the problem, as burning magnesium is extremely difficult to extinguish.

Room destruction factor:

A large magnesium fire will destroy your classroom. For this reason, make sure your students have only enough magnesium ribbon as is necessary to perform this lab.

How the lab works:

Students heat magnesium in the presence of atmospheric oxygen, causing white magnesium oxide powder to be formed. Because two elements have combined to form a chemical compound, this qualifies as a synthesis reaction.

Clean up:

Once the MgO powder has cooled, it can be thrown in the trash.

Station 2: Decomposition reaction

$$CuSO_4 \cdot 5\ H_2O_{(s)} \rightarrow CuSO_{4(s)} + 5\ H_2O_{(g)}$$

Equipment:

- 50 grams of copper (II) sulfate pentahydrate
- Bunsen burner
- flint striker
- crucible
- crucible tongs
- ring stand
- ring
- clay triangle
- forceps
- labeled waste beaker

Safety:

The main danger with this station is the heat given off by the Bunsen burner while heating the copper (II) sulfate pentahydrate. As always, make sure your students are aware of the dangers in working with Bunsen burners, particularly that they should not wear gloves.

Copper (II) sulfate is poisonous and care should be taken to ensure your students wash their hands after performing this station, even if they haven't actually touched any of it with their bare hands.

Room destruction factor:

It's unlikely that this station will damage your classroom in any way, though care should be taken to ensure that flammable objects are kept away from this station.

How the lab works:

This station is basically the same thing as the percent composition lab, except that copper (II) sulfate pentahydrate is used in place of magnesium sulfate heptahydrate. Students should notice that the copper (II) sulfate hydrate turns from a deep blue color to sky blue as it dehydrates.

Before doing this station, ensure your students recall how to name hydrates. Otherwise, they may have trouble generating the equation for this process.

Clean up:

Copper (II) sulfate anhydrate may be rehydrated by dissolving it in water. When the water evaporates, you're left with brand new copper (II) sulfate pentahydrate crystals, which may be reused.

Station 3: Single displacement reactions

$$Zn_{(s)} + 2\ HNO_{3(aq)} \rightarrow Zn(NO_3)_{2(aq)} + H_{2(g)}$$

Equipment:

- 50 grams of mossy zinc
- 50 mL of 5 M nitric acid
- 1 box of rubber gloves
- plastic disposable pipet
- 10 mL beaker or small test tube
- 250 mL beaker
- 100 grams of sodium bicarbonate
- fume hood

Safety:

Nitric acid is extremely corrosive and causes serious, slow-healing burns upon contact with bare skin. For this reason, students should always wear rubber gloves when performing this station. Students should also be cautioned to wipe down the countertop to ensure that any spilled nitric acid is removed for the next lab group. Only very small quantities of nitric acid should be given to students at any one time to lessen the severity of potential spills.

If you prefer, this station can be done using magnesium and acetic acid, forming magnesium acetate and hydrogen gas. Though nowhere near as dramatic, it may be a better choice for younger or particularly accident-prone students.

This station should always be performed in a fume hood because very toxic nitrogen dioxide gas is formed. This gas has an unmistakable red color, making it very easy to see. Students should be informed that nitrogen dioxide is a by-product of this reaction and not one of the products they're studying.

Room destruction factor:

Nitric acid is extremely corrosive and causes rapid damage to wood and metals. As a result, it's a good idea to wipe down the inside of your hood with a paper towel soaked in sodium bicarbonate solution after the lab is finished to ensure that it is completely neutralized.

How the lab works:

Students place a chunk of zinc into the nitric acid and watch as hydrogen bubbles evolve and the zinc is converted to zinc nitrate. The reaction is very rapid, usually causing students to step back a few paces in surprise.

Clean-up:

When the reaction has stopped, students should pour the nitric acid solution into a 250 mL beaker containing 100 grams of sodium bicarbonate. This instantaneously neutralizes the nitric acid – when all classes have finished this lab, the sodium bicarbonate/nitric acid waste may be safely washed down the sink. Additionally, the inside of the hood should be wiped down with a saturated sodium bicarbonate solution to ensure that any stray nitric acid drops have been neutralized.

Station 4: Double displacement reactions

$$2\ KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow 2\ KNO_{3(aq)} + PbI_{2(s)}$$

Equipment:

- 100 mL of 0.25 M KI solution
- 100 mL of 0.125 M $Pb(NO_3)_2$ solution (make with distilled water)
- 2 watch glasses
- 2 plastic disposable pipets, labeled "KI" and "$Pb(NO_3)_2$"
- 100 mL beaker labeled "Station 3 Waste"

Safety:

Lead-containing compounds are toxic, but are safe to handle with bare skin as long as they are cleaned off quickly with soap and water. When inhaled or ingested, lead compounds cause a wide variety of health problems. Additionally, solid lead nitrate is explosive when heated, so make sure that students only have access to the solution.

Room destruction factor:

Lead (II) iodide is a very bright yellow powder that can be hard to clean from some surfaces. Make sure your students remove all traces of the yellow precipitate before moving to the next station and there should be no problem.

How the lab works:

Students combine lead (II) nitrate and potassium iodide solutions, precipitating lead (II) iodide. The presence of this bright yellow precipitate alerts students that a reaction has taken place.

The main thing that causes problem in this lab is that students use the wrong pipet for dispensing the chemicals, causing lead (II) iodide to form in the pipet instead of the watch glass. Labeling each pipet and tethering them to the correct bottle should keep this from happening.

Clean up:

All solutions should be poured into the waste container. Make sure your students wash their hands after this station to remove any stray lead salts. When the lab is finished, lead waste should be disposed of professionally.

Station 5: Acid-base reactions

$$NaOH_{(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$$

Equipment:

- 50 mL phenolphthalein solution
- 100 mL 0.1 M HCl solution
- 100 mL 0.1 M NaOH solution
- 3 disposable plastic pipets, labeled “indicator”, “HCl”, and “NaOH”
- 2 watch glasses

Safety:

HCl and NaOH are both extremely caustic and can cause skin and eye damage. At these concentrations, the danger of skin burns is very low and can be taken care of with postlab hand-washing. As always, goggles should be worn during this station, as both HCl and NaOH are strong eye hazards.

Phenolphthalein is a very effective laxative and seems to stick well to the skin. Make sure your students wash their hands extremely well with soap after this station. When making the phenolphthalein indicator for this lab, it is important that you wear gloves, as a large spill on your hands could cause the phenolphthalein to be absorbed through your skin. Generally, it’s a good idea to give the students only a small quantity of indicator at a time. It’s also a good idea to withhold the laxative properties of phenolphthalein from your students, as some may imagine the unpleasant pranks which may be caused with it.

How the lab works:

Students place a few drops of NaOH solution into a watch glass and add phenolphthalein indicator, causing the solution to turn pink. As hydrochloric acid is added, the solution will become colorless as the base is neutralized. If your countertops are black, your students may find it easier to see the color changes if they place a piece of white paper under the watch glass.

It is common that students either don't see the initial pink color or don't see a color change as the hydrochloric acid is added to it. These problems are caused if students don't follow instructions and add the proper chemicals at the right times. Should they occur, simply tell your students to try the procedure over again and to be very careful about reading the instructions.

Clean up:

All solutions can be placed down the drain at the conclusion of the lab.

Station 6: Combustion reactions

$$2\ C_2H_6O_{(l)} + 6\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 6\ H_2O_{(g)}$$

Equipment:

- 100 mL ethanol
- 2 books of matches
- 1 watch glass
- 1 disposable plastic pipet

Safety:

Ethanol is extremely flammable and should be handled carefully. The risk of serious fire can be minimized by providing the students with only a very small bottle of ethanol, rather than a large reagent bottle. Reagent-grade ethanol is also extremely toxic and students should be warned against spending too much time around its fumes.

Room destruction factor:

As long as fire safety rules are followed, your room should be OK.

How this lab works:

Students place a small amount of ethanol into a watch glass and light it with a kitchen match. The resulting fire is evidence of a combustion reaction.

Clean up:

The only waste produced by this station is the burned match. Once it has cooled, it may be safely thrown in the trash.

Solutions to the Types of Reaction Lab

Station 1:

Step 3: The magnesium flame burns with a white flame and white powder is formed as the product.

Question: $2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$

Station 2:

Step 3: As the deep blue crystal heats up, it it changed into a chalky, sky-blue powder.

Question: $CuSO_4 \cdot 5\ H_2O_{(s)} \rightarrow CuSO_{4(s)} + 5\ H_2O_{(g)}$

Station 3:

Step 3: The zinc almost instantly dissolves, and the solution becomes murky as red/brown gas is given off.

1) $Zn_{(s)} + 2\ HNO_{3(l)} \rightarrow Zn(NO_3)_{2(aq)} + H_{2(g)}$
2) The bubbles given off are hydrogen gas.

Station 4:

Step 2: A bright yellow precipitate is formed.

1) School buses and pencils, though different paints are now used due to the toxicity of lead compounds.
2) The atoms in this reaction have recombined to form totally new compounds. When water and alcohol are mixed, no new compounds are formed, so the resulting solution doesn't change color.
3) $Pb(NO_3)_{2(aq)} + 2\ K\,I_{(aq)} \rightarrow PbI_{2(s)} + 2\ KNO_{3(aq)}$

Station 5:

1) Acid-base reactions take place almost instantaneously, as demonstrated by the reaction in this station.
2) It takes amore time for the medicine to be absorbed through the stomach lining and travel through the bloodstream than it does for a base to directly neutralize an acid.

3) $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)}$
Note: Some students may believe that the indicator should be shown in the equation – explain to these students that the indicator only shows what is going on in the reaction but isn't directly involved.

Station 6:

Step 2: A blue/yellow flame forms when ethanol is burned.

1) Students should explain their reasons for believing that ethanol is a good fuel or a poor one.
2) $2\ C_2H_6O_{(l)} + 6\ O_{2(g)} \rightarrow 4\ CO_{2(g)} + 6\ H_2O_{(g)}$

Solutions for the Six Types of Reaction Worksheet

1) combustion
2) single displacement
3) decomposition
4) double displacement
5) synthesis

6)	$\underline{2}\ LiOH + \underline{1}\ H_2SO_4 \rightarrow \underline{1}\ Li_2SO_4 + \underline{2}\ H_2O$	acid-base
7)	$\underline{1}\ C_5H_{12} + \underline{8}\ O_2 \rightarrow \underline{5}\ CO_2 + \underline{6}\ H_2O$	combustion
8)	$\underline{2}\ Fe + \underline{3}\ Ni(OH)_2 \rightarrow \underline{3}\ Ni + \underline{2}\ Fe(OH)_3$	single displacement
9)	$\underline{1}\ NaHCO_3 \rightarrow \underline{1}\ NaOH + \underline{1}\ CO_2$	decomposition
10)	$\underline{1}\ Pb(C_2H_3O_2)_2 + \underline{2}\ NaBr \rightarrow \underline{1}\ PbBr_2 + \underline{2}\ NaC_2H_3O_2$	double displacement

<u>Six Types of Chemical Reaction Handout</u>

All chemical reactions can be grouped into one of six categories:

1) **Combustion**: A combustion reaction occurs when oxygen reacts with another compound to form water and carbon dioxide. These reactions are exothermic, meaning that they produce heat. One example of this is the combustion of napthalene, a common ingredient in mothballs:

$$C_{10}H_{8(s)} + 12\ O_{2(g)} \rightarrow 10\ CO_{2(g)} + 4\ H_2O_{(g)} + \text{heat}$$

2) **Synthesis**: A synthesis reaction occurs when two elements or compounds combine to form a more complex molecule. These reactions can be expressed by the equation A+B → AB. One example of a synthesis reaction is the combination of iron and sulfur to form ion (II) sulfide:

$$8\ Fe_{(s)} + S_{8(s)} \rightarrow 8\ FeS_{(s)}$$

3) **Decomposition**: A decomposition reaction is the opposite of a synthesis reaction, and occurs when a molecule breaks apart to form two elements or less complex molecules. These reactions can be expressed by the equation AB → A + B. One example of a decomposition reaction is the electrolysis of water to form oxygen and hydrogen gas:

$$2\ H_2O_{(l)} \rightarrow 2\ H_{2(g)} + O_{2(g)}$$

4) **Single displacement**: Single displacement reactions occur when a single element reacts with a chemical compound and switches places with one of the elements in the compound. These reactions can be expressed by the equation A + BC → AC + B. An example of a single displacment reaction is when magnesium replaces hydrogen atoms in water to make magnesium hydroxide and hydrogen gas:

$$Mg_{(s)} + 2\ H_2O_{(l)} \rightarrow Mg(OH)_{2(s)} + H_{2(g)}$$

5) **Double displacement**: Double displacement reactions occur when the cations of two ionic compounds switch places. These reactions can be expressed by the equation AB + CD → AD + CB. One example of a double displacement reaction is the reaction of silver acetate with sodium chloride to form silver chloride and sodium acetate:

$$AgC_2H_3O_{2(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaC_2H_3O_{2(aq)}$$

6) **Acid-base**: Acid-base reactions are simply double displacement reaction in which water is produced from the H^+ ion from the acid and the OH^- ion in the base. In acid-base reactions, the products always include water and an ionic compound, as shown by the following equation. An example of an acid-base reaction is the reaction of hydrobromic acid with potassium hydroxide:

$$HBr_{(aq)} + KOH_{(aq)} \rightarrow KBr_{(aq)} + H_2O_{(l)}$$

How to identify types of reactions

Let's say that you've been asked to determine the type of reaction that's taking place in a chemical process from the equation. To do this, read through the questions below in order until you can answer "yes" to any of them. When you answer "yes", don't move on to the next question, because you've found your final answer! If you answer "no", continue to the next question.

1) Does the chemical equation contain oxygen, carbon dioxide, <u>and</u> water? If it does, it's a **combustion** reaction.

2) Do simple molecules combine to form a more complex molecule? If they do, it's a **synthesis** reaction.

3) Does a complicated molecule break apart to form two or more simpler ones? If it does, it's a **decomposition** reaction.

4) Are there any chemicals anywhere in the equation that consist of only one element (Fe, Na, H_2, etc.)? If so, it's a **single displacement** reaction.

5) Is water formed during this reaction? If so, it's an **acid-base** reaction. If not, it's a **double displacement** reaction.

Types of Reactions Lab

In this lab, we will be investigating the six types of chemical reaction we discussed in class. Each lab station will take between 10-15 minutes – if you finish the experiment ahead of time, use the remaining time to answer the questions at the end of each station.

It is important that you wear goggles during every section of this lab, as there are many dangerous chemicals in use. If you're not sure about the proper safety procedures for any section of this lab, ask your teacher. After all, it's better to ask for clarification than to injure yourself!

One more thing: Make sure you read the introductions to each section before starting the lab. There may be some information in there to help you with the questions!

Station 1: Synthesis reaction

Chemists always use extreme care when working with magnesium metal because of its high reactivity. One of magnesium's dangerous reactions is its reaction with atmospheric oxygen to form magnesium oxide. This reaction gives off a great deal of light and heat, and is nearly impossible to stop once it has started. This is the reaction we are going to be doing at this station.

People often mistake magnesium fires for combustion reactions because a great deal of fire and smoke are generated. Actually, it doesn't fit the definition of a combustion reaction, as carbon dioxide and water are not generated. Instead, the magnesium metal combines with oxygen to form magnesium oxide. Because two elements combine to form a more complex molecule, this is classified as a synthesis reaction.

Procedure:

1) Obtain a small piece of magnesium ribbon.

2) Light a Bunsen burner, making sure it has been adjusted so a bright blue cone of flame is visible.

3) Using your crucible tongs, hold the magnesium ribbon directly over the blue cone of the Bunsen burner flame until the reaction starts. Don't worry – you'll know when it happens! Write your observations here:

4) When the reaction is finished, clean up your lab station by sweeping the magnesium oxide residue into the trash.

Station 1 Question: Write the chemical equation for the synthesis of magnesium oxide from magnesium and oxygen:

Station 2: Decomposition reaction

Some of the most violent chemical reactions known are decomposition reactions. The sodium azide that inflates airbags and most high explosives work by decomposing into simpler molecules. In this section, we will decompose copper (II) sulfate pentahydrate into copper (II) sulfate anhydrate and water.

Procedure:

1) Set up a ring stand with a ring, clay triangle, and Bunsen burner.

2) Use forceps to place a crystal of copper (II) sulfate pentahydrate into a crucible. Use your crucible tongs to place the crucible in the clay triangle.

3) Light the Bunsen burner and heat the crucible for 5 minutes. Write down your observations here:

4) When the crystal no longer appears to be changing color, turn off the Bunsen burner and let the crucible cool for 5 minutes. The copper (II) sulfate anhydrate should be placed into the labeled waste beaker for recycling.

Station 2 question: Write the equation for the decomposition of copper (II) sulfate pentahydrate into copper (II) sulfate anhydrate and water.

Station 3: Single displacement reactions

Single displacement reactions are important in the area of electrochemistry. Without a good knowledge of single displacement reactions, it would be impossible to build batteries of any kind.

In the reaction you will be observing, zinc reacts with nitric acid to form hydrogen gas and zinc nitrate. When this reaction takes place, electrons move from zinc to hydrogen. In a battery, these electrons can be used to generate electricity.

Procedure: (This station should be done in a fume hood!)

1) Place a chunk of zinc into the 10 mL beaker.

2) Have each member of your lab group put on a pair of rubber gloves.

3) Fill a pipet with 5 M nitric acid and squirt it into the 10 mL beaker. Write your observations below:

4) When the reaction has stopped, carefully pour the contents of the 10 mL beaker into the 250 mL beaker containing sodium bicarbonate. This will neutralize the remaining nitric acid. Wipe the countertop in the fume hood with a paper towel to ensure that no nitric acid droplets remain.

Station 3 questions:

1) Write the equation for the reaction you observed:

2) What do you think the bubbles were that you saw being formed?

Station 4: Double displacement reactions

Double displacement reactions are extremely common. For example, the shells of many sea animals are made of calcium carbonate, formed from double displacement reactions within the animals.

The reaction you will be studying is the reaction of lead (II) nitrate and potassium iodide to form dissolved potassium nitrate and a lead (II) iodide precipitate. This reaction was once used in the manufacture of paint for household products. You will no doubt be able to guess what extremely common school product was painted with this pigment.

Station 4 procedure:

1) In a watch glass, place 5 drops of 0.25 M potassium iodide solution using the labeled pipet.

2) To this watch glass, add 5 drops of 0.125 $Pb(NO_3)_2$ solution using the labeled pipet. Write your observations here:

3) Clean up: Rinse the precipitate into the labeled waste beaker.

Station 4 questions:

1) The precipitate you formed, lead (II) iodide, was once used as a pigment in paint for a common school product. What do you think used to be covered with this paint?

2) Why do two clear solutions form a colored compound when they are combined? After all, when you mix water with rubbing alcohol, the resulting solution is colorless!

3) Write the equation for the reaction you observed in this station:

Station 5: Acid-base reactions

Many reactions that take place in the body and in the atmosphere are acid-base reactions. One example of acid-base chemistry is acid rain: Acid rain occurs when acidic gases are released by power plants and dropped hundreds of miles from their sources, killing fish and plants.

In this section, you will observe the reaction of hydrochloric acid with sodium hydroxide to form sodium chloride and water.

<u>Station 5 procedure:</u>

1) Place 5 drops of 0.1 M NaOH into a watch glass. Add one drop of indicator solution to it. The color you see indicates the presence of base.

2) Slowly add hydrochloric acid to the watch glass containing sodium hydroxide until the color of the solution changes. The new color indicates that the base in the solution has been neutralized and that the solution is now becoming acidic.

3) Clean up: Rinse the contents of the watch glass into the sink and wash your hands.

<u>Station 5 questions</u>:

1) Antacids such as Tums, Rolaids, and Maalox all cure heartburn by neutralizing stomach acid with a base. Does the experiment above suggest why these antacids work so quickly to cure heartburn?

2) Acid-reducing medications such as Pepcid AC, Tagamet HB, and Axid AR work by a completely different mechanism that does not neutralize stomach acid. Instead, these medications work by making the stomach secrete less acid in the first place. Do you think this explains why these medications work so slowly? Explain.

3) Write the equation for the reaction you observed in station 5:

Station 6: Combustion reactions

Modern life depends on combustion reactions. The energy we use in our homesis generated by power plants that burn coal. The energy that powers our cars is generated by the combustion of gasoline. The heat that cooks food on a gas stove is generated by the combustion of natural gas. Without combustion reactions there wouldn't be enough energy to keep society running.

Ethanol is commonly added to gasoline to improve air quality. In this station, we will be observing the combustion of ethanol (C_2H_6O).

Station 6 procedure:

1) Place about 10 drops of ethanol into a watch glass with a pipet.

2) Light and drop a match into the ethanol, making sure not to put your hand too close to the flame. Write your observations here:

3) Clean up: Once it has cooled, throw the match into the trash.

Station 6 questions:

1) Ethanol fires are far less hot than gasoline fires. This makes them safer to use in automobiles, but less capable of producing large amounts of energy. Keeping this in mind, do you think that ethanol is a good source of fuel for automobiles? Explain.

2) Write the equation for the combustion of ethanol:

<u>Six Types of Reaction Worksheet</u>

Fill in the blanks using the correct type of reaction:

1) A(n) ________________________________ reaction is one in which oxygen bonds with another element to form carbon dioxide, water, and heat.

2) A(n) ________________________________ reaction takes place when an element switches places with one of the elements in a chemical compound.

3) A(n) ________________________________ reaction is one where a complex molecule breaks down to form two or more less complicated molecules.

4) A(n) ________________________________ reaction is one where the cations of two ionic compounds switch places.

5) A(n) ________________________________ reaction is when two or more simple molecules join to form a more complicated molecule.

For the following questions, balance the equation and indicate what type of reaction the equation shows:

6) ____ $LiOH$ + ____ $H_2SO_4 \rightarrow$ ____ Li_2SO_4 + ____ H_2O

Type of reaction: ________________________________

7) ____ C_5H_{12} + ____ $O_2 \rightarrow$ ____ CO_2 + ____ H_2O

Type of reaction: ________________________________

8) ____ Fe + ____ $Ni(OH)_2 \rightarrow$ ____ Ni + ____ $Fe(OH)_3$

Type of reaction: ________________________________

9) ____ $NaHCO_3 \rightarrow$ ____ $NaOH$ + ____ CO_2

Type of reaction: ________________________________

10) ____ $Pb(C_2H_3O_2)_2$ + ____ $NaBr \rightarrow$ ____ $PbBr_2$ + ____ $NaC_2H_3O_2$

Type of reaction: ________________________________

Chapter 10 – Conservation of Mass Lab

Overview

The law of conservation of mass states that the mass of the products of a chemical reaction is the same as the mass of the reagents. This is easy to imagine – after all, the weight of a pizza is the same as the weight of the ingredients in it. In this chapter, students will learn about the law of conservation of mass and verify it using the reaction of acetic acid with sodium carbonate.

Teaching about the law of conservation of mass

Students don't have any trouble with the basic concept that the weight of what is made in a reaction is the same as the weight of the reagents. Unfortunately, many students tend to believe that the law of conservation of mass is like socialism, "eating healthy", or do-it-yourself electrical work: A nice idea that never really works out the way you'd like•.

There's a good reason for this: When students do labs, the weight of what they make is *never* the same as the weight of the reagents. If your students do a good job in the lab, their yield may be 90%. If your students do a terrible job, their yield may be 20%. It's obvious to us that the missing mass has gone somewhere, but students sometimes think the mass just vanished. Even worse is when students get a yield greater than 100%. We may realize that this happens due to the presence of impurities, but in your students minds, they actually created new matter!

The way to combat this is not to simply insist that the law of conservation of mass is true, because students, like anybody else, want more information before accepting an idea. Instead, the best way to teach the law of conservation of mass is to use a real-world example of when the law of conservation of mass is used. I recommend using Velveeta macaroni and cheese.

Pull out a box of the macaroni and cheese where you are provided with a can of premade cheese sauce and a box of noodles. Ask your students how much the cheese weighs, how much the noodles weight, and how much the water used to make the noodles weighs. Next, ask them how much the final macaroni and cheese product weighs. Students won't have any trouble identifying that the ingredients and final product have the same mass.

• If you are a socialist, healthful eater, or home do-it-yourselfer that was offended by this joke, please address all complaints to the author at his home address: 1060 West Addison, Chicago, IL, 60613-4397.

Let's translate macaroni and cheese into chemical terms. Write the recipe for macaroni and cheese on the board, followed by the equation for the synthesis of water:

8 oz. macaroni + 10 oz. cheese + 10 oz. water → 28 oz. mac and cheese

4 grams hydrogen + 32 grams oxygen → ? grams water

Have your students determine how much water will be made in the second equation. Most will get the idea that all you need to do is add the masses of the hydrogen and oxygen together to find the mass of water. Voila! They understand the law of conservation of mass!

Let students know that cooking and chemistry are essentially the same thing. After all, both require a talented person to mix a number of things together using specialized techniques, resulting in some desired product. Because most students will know how to cook *something*, this should give them hope that they can successfully perform chemical reactions.

Of course, we still have a problem: If the law of conservation of mass states that the mass of what you make is equal to the mass of the ingredients, why don't we ever get 100% yield for a chemical reaction? The answer, I tell my students, is experimental error.

What is experimental error? It's the process of screwing stuff up in the lab. In my experience, there are three main ways that experiments are screwed up in the lab:

1) **Human error**: *You* screwed up! Before students are offended that you think they would screw up in the lab, explain that it's impossible not to make mistakes in the laboratory. These mistakes can range from splashing chemicals while pouring them together, leaving residue on the inside of glassware, and any number of procedural errors. It's impossible not to make any errors, making it inevitable that yields will be less than 100%. Most error is human error.

2) **Instrumentation error**: *The black box* screwed up! Modern technology has given us a wide variety of machines that give us data in magical and mysterious ways. Take an electronic balance, for example: Place an object on the pan and the mass magically shows up on the screen. Unfortunately, it's hard to tell when these machines are mistaken, making them a possible source of error. Students may, however, be unhappy to find that most instruments work reliably, making this sort of error unlikely, but not impossible.

3) **Unknown error**: *Something* screwed up! Was atmospheric humidity the culprit? Was there a mini-earthquake that threw off the data? Probably not – most unknown errors are just human errors that we can't even imagine we made in the first place.

No matter what the source, students should be made aware that error is inevitable, no matter how carefully an experiment is performed. This may make them feel better if they do something wrong in an experiment or if their percent yields are lower than they'd like.

Doing the conservation of mass lab

Equipment:

- 250 grams of sodium carbonate
- 2 liters of 2 M acetic acid
- 8-250 mL beakers
- 20 latex gloves
- 40 rubber bands
- 2 balances with 500 gram weighing capacity
- Many paper towels

Safety:

Acetic acid is an eye hazard and may cause some skin irritation, particularly on broken skin. Keep a bottle of saturated sodium bicarbonate solution handy for spills or skin burns. If acetic acid is introduced to the eyes, flush immediately with water and call for medical assistance!

Room destruction factor:

The product of this reaction, sodium acetate, tends to be spilled on the countertops, leaving threadlike crystals behind when the water evaporates. Have students wipe down the countertops several times with damp and dry paper towels to minimize the formation of these crystals.

How the lab works:

Students place acetic acid in a beaker and place sodium carbonate in one of the fingers of a rubber glove. The sodium carbonate is then tied into the finger of the glove to keep it from mixing with the acetic acid, and the glove is stretched over the beaker. The resulting apparatus is weighed to determine the initial mass. When the finger is untied, the sodium carbonate mixes with the acetic acid, forming carbonic acid and sodium acetate. The carbonic acid decomposes to form carbon dioxide (which inflates the glove) and water. When the reaction has finished, students weigh the whole apparatus and find that the mass hasn't changed.

Students enjoy this lab, particularly the inflation of the rubber glove. Despite the fears of some students, the glove will not pop and the contents are not particularly dangerous.

The weight of the apparatus after the reaction is invariably smaller than the initial weight. One reason for this change in mass is that carbon dioxide gas has buoyancy in air, making it impossible to accurately measure. Additionally, carbon dioxide (and gases in general) are difficult to work with because they escape easily from closed systems – CO_2 leakage will also be a source of error for this lab. The combined discrepency in mass from both errors typically ranges between 2-4 grams.

It should be noted that most of the mass in this experiment is due to the beaker, the rubber glove, and the rubber band, none of which are reagents or products in this reaction. As a result, the weight of the products will be roughly the same as the weight of the reagents even if a very large amount of carbon dioxide is released. This is, admittedly, dishonest, but it ensures the lab will work successfully.

What can go wrong:

- The rubber glove comes off of the beaker before students have a chance to weigh it. Students simply have to start over.
- The rubber glove springs a hole. Again, students will need to start over.
- The rubber band breaks. Start over.

Clean up:

Everything can be safely placed down the sink, as the only product of this reaction is sodium acetate.

Solutions for the Conservation of Mass Lab

Prelab: The combined mass of the products is the same as the combined mass of the reagents. Since the reagents weigh a total of 636 grams, the products should also weigh 636 grams.

Lab:

Steps 5 and 7: Weight should be measured in “grams”.

Lab questions:

1) The law of conservation of mass is confirmed, as the products weigh roughly the same as the reagents.
2) There will be some error, which should be documented by students.
3) Possible sources of error include a leaky glove, buoyancy of CO_2, etc.
4) The law of conservation of mass was difficult to devise because it’s hard to get exactly the same weights for products and reagents when doing a reaction due to experimental error.

Solutions for the Conservation of Mass Worksheet

1) 78 grams, because 72 + 6 = 78 grams.
2) The law of conservation of mass tells us that the masses of products and reagents is the same because the number of atoms of each element is the same for both the products and reagents. If we didn’t balance equations, the number of atoms of each element wouldn’t be the same.
3) The mass is released in the form of water and carbon dioxide, both of which are in the car exhaust. The combined weight of the carbon dioxide and water will be 35 kilograms plus the weight of the oxygen it reacted with.
4) This doesn’t violate the law of conservation of mass because other undesirable products are produced that account for the missing weight.

Conservation of Mass Handout

The law of conservationof mass states that the total mass of the products in a chemical reaction is the same as the total mass of the reagents for that reaction.

What does this mean? Let's say you're making chili and the recipe calls for one pound of ground beef, two pounds of beans, two pounds of various vegetables, ½ pound of various spices, and 2 pounds of tomato sauce. What would the final weight of the chili be? It should be clear that the answer is 1 + 2 + 2 + ½ +2 = 7 ½ pounds, because that's the weight of the ingredients you used to make it. This is exactly the same thing that happens in a chemical reaction.

There's nothing mysterious about chemical reactions. Chemical reactions are the same thing as cooking[1] and chemical equations are the same thing as recipes[2]. In most cases, the rules of chemistry and cooking are the same, except that you shouldn't eat the products of a chemical reaction[3].

[1] Except that you use expensive glassware instead of pots and pans.
[2] Except that chemistry books cost far more than cookbooks.
[3] Depending on your skill, you may also be able to go without goggles in the kitchen.

Conservation of Mass Lab

The law of conservation of mass states that whenever we do a chemical reaction, the products we make have the same weight as the reagents we started with. No mass is created or destroyed during this process.

This should make intuitive sense. It's like saying that when you make a pizza, the weight of the pizza is equal to the weights of all the toppings, cheese, tomato sauce, and crust. In chemistry, the same thing is true: When we do a reaction, the weight of what we make is the same as the weight of the ingredients.

In this lab, we're going to test the law of conservation of mass.

Prelab:

The chemical equation for the combustion of benzene is this:

$$\mathbf{2\ C_6H_6 + 15\ O_2 \rightarrow 12\ CO_2 + 6\ H_2O}$$

If I start with 156 grams of C_6H_6 and 480 grams of O_2, what will be the combined weight of the CO_2 and H_2O that are formed?

Lab:

1) Each group should get a 250 mL beaker, a latex glove, and 4 rubber bands.

2) Weigh 13 grams of sodium carbonate and place it in one of the fingers of a rubber glove. Tie the finger shut with rubber bands.

3) Measure 125 mL of 2 M acetic acid with a graduated cylinder and pour it into the beaker.

4) With help from all members of your group, fasten the rubber glove around the mouth of the beaker with the remaining three rubber bands. Make sure the sodium carbonate doesn't fall into the acetic acid!

5) <u>Before dropping the sodium carbonate into the beake</u>r, find the weight of the entire experimental apparatus (the glove and beaker):

   Initial weight of the experimental setup: ________________________

6) Making sure the glove is firmly fastened around the mouth of the beaker, untwist the finger of the rubber glove and let the sodium carbonate fall into the beaker. It is extremely important that no sodium carbonate be left in the glove, because if it is the reaction will not be complete.

7) When the reaction has finished, weigh the resulting solution without taking off the rubber glove. This is your final weight:

Final weight: ________________________

8) Clean up: The glove may be thrown into the garbage and the rubber band may be saved. The solution in the beaker may be poured down the sink.

Questions:

1) Did your experiment confirm the law of conservation of mass? Explain, using your lab data.

2) Do your results show any error for this lab? Explain, using your lab data.

3) What are some possible sources of error for this lab? These should be reasonable!

4) Why do you think it took so long for people to come up with the law of conservation of mass? Explain, using your experience from this lab.

Conservation of Mass Worksheet

Use your knowledge of the law of conservation of mass to answer these questions:

1) The chemical reaction for making benzene from carbon and hydrogen is:

$$6\ C + 3\ H_2 \rightarrow C_6H_6$$

If 72 grams of carbon and 6 grams of hydrogen react completely, how many grams of benzene will be formed? Explain your answer.

2) Do you think the law of conservation of mass has anything to do with the reasons why we need to balance equations? Explain.

3) When the tank of my car is empty, I fill the tank with about eight gallons of gas. This much gas weighs approximately 35 kilograms. However, when it's time to refill my car again, I don't have to remove 35 kilograms of product from my car. What has happened, and does it violate the law of conservation of mass? Explain.

4) Many drug companies make protease-inhibitor drugs that stop the replication of HIV. These drugs cost thousands of dollars because making them is very inefficient – a typical yield is 2%! Does this violate the law of conservation of mass, or is there a better explanation?

Chapter 11 – Writing Complete Equations

Overview

Students have a hard time learning to write complete chemical equations. Though they typically remember either to balance the equation or to write the subscripts and reaction conditions, they rarely do both simultaneously. In this chapter, students will learn the elements of a complete chemical equation and use tthem to write the equations of five reactions they will see you perform.

Teaching about writing complete equations

In the last chapter, I mentioned that chemical equations are nothing more than recipes for doing chemical reactions. Like any recipe, it's not just a list of ingredients. Instead, a complete chemical equation explains how to make the chemical change occur in the lab, as well as other handy information that will help the experimenter. These terms require specialized knowledge of chemical changes, just as recipes for food require that you know terms like "poach" or "broil".

Some of the more common terms that you may encounter in a chemical process are given to you below. Some involve chemical changes the experimenter will need to perform, while others just give information intended to help the experimenter.

Term	What it means
(l)	The chemical is a liquid.
(s)	The chemical is a solid.
(g)	The chemical is a gas.
(aq)	The chemical is aqueous (dissolved in water).
Δ	Written above an arrow, it indicates that heat needs to be added to the reagents for them to react.
any temperature	Written above an arrow, it indicates that the reagents should be brought to this temperature for them to react properly.
any pressure	Written above an arrow, it indicates that the reagents should be brought to this pressure for them to react properly.
ΔH = [a number]	Written after an equation, it indicates the amount of energy change during the reaction. A positive sign indicates an endothermic reaction and a negative sign indicates an exothermic reaction.
anything else	This is a specialized instruction for performing the reaction.

There are many, many symbols which may be used in a chemical reaction. Tell your students only the most common ones – let their organic chemistry professors explain what "refluxing" is.

Doing the writing complete equations lab

During each of these lab stations, you will perform a chemical reaction before the class, explaining what reagents you are using and the products that are formed. Your students will use their knowledge of chemical symbols and terms to write complete, balanced equations for each reaction. Students enjoy this lab because the demonstrations are dramatic and very interesting.

I don't allow the students to perform these reactions for themselves because they involve more danger than I believe high school chemistry students are prepared for. Many of these labs involve fire or strong acids, making this a suitable demonstration lab but not a suitable lab for the students to perform. Introductory chemistry students simply don't have enough experience in the lab to perform these demonstrations safely.

That said, some of these demonstrations may be more dangerous than beginning or inexperienced chemistry teachers will feel comfortable with. **It is very important that if you don't feel 100% confident that you can do these demonstrations without harming yourself or others, you should not perform them!** Performing chemical reactions that you don't feel comfortable with ensures that one day somebody will be badly injured.

As always, make sure that you and your students wear goggles during this lab!

Demonstration #1 – Silver nitrate and hydrochloric acid

$$AgNO_{3(aq)} + HCl_{(aq)} \rightarrow AgCl_{(s)} + HNO_{3(aq)}$$

Equipment:

- 3 grams of silver nitrate
- 3 mL concentrated hydrochloric acid
- Large test tube (capacity of ~15 mL)
- 5 mL distilled water
- Plastic disposable pipet

Safety:

Hydrochloric acid is both an eye and skin hazard. Everybody in the classroom should wear goggles and you should wear rubber gloves. If you dress nicely, you may also want to wear a lab coat.
Silver nitrate, while not terribly dangerous, undergoes a photochemical reaction that forms a dark black silver precipitate. You'll notice that if you get silver nitrate on your hands, over a period of four hours your hands will turn dark black as the silver is reduced from the +1 to the 0 oxidation state. This isn't dangerous, but it is annoying to have black hands for a week.

Room destruction factor:

A large lump of silver chloride is produced, which has the potential of clogging your sink. Personally, I save the silver chloride that's produced by this reaction – students seem to get a kick out of seeing that the same element that is in much of their jewelry can be turned into a white powder.

How the lab works:

In this lab, you'll be adding hydrochloric acid to an aqueous solution of silver nitrate, forming silver chloride. Students find this transformation to be very interesting, as two colorless solutions instantly form a giant white blob in the test tube.

The procedure for this demonstration:

1) Pour 5 mL of distilled water in the test tube.

2) Add the silver nitrate and swirl until it has completely dissolved.

3) Fill the pipet with hydrochloric acid and quickly squirt it into the test tube. Silver chloride precipitate will form as the hydrochloric acid hits the solution.

4) Clean up: The solution can be neutralized with sodium bicarbonate (add $NaHCO_3$ until it stops bubbling). The resulting liquid can be decanted and poured down the sink. The silver chloride may be saved or thrown away.

What can go wrong:

If tap water is used instead of distilled water, it will become cloudy as soon as the silver nitrate is added. This cloudiness is due to the reaction of silver with the dissolved chloride in tap water.

Demonstration #2: Nitric acid and sodium carbonate

$$2\ HNO_{3(aq)} + Na_2CO_{3(s)} \rightarrow 2\ NaNO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Equipment:

- 10 mL of 5 M HNO_3
- 10 grams Na_2CO_3
- 250 mL beaker
- 50 mL tap water
- Rubber gloves

Safety:

Nitric acid is unbelievably corrosive and causes very serious, slow-healing burns. Make sure that everybody in the classroom wears goggles and that you wear gloves at all times during this demonstration! If you wear nice clothes, put on a lab coat. If you have a fume hood, I recommend you perform the lab in the hood for additional safety.

Room destruction factor:

Chemistry classrooms generally aren't harmed much by nitric acid spills, as countertops are extremely unreactive. However, if you are teaching chemistry in a room intended for another course, splashed nitric acid may damage the desktop.

How the lab works:

In this reaction, carbonic acid and sodium nitrate are formed by the addition of nitric acid to sodium carbonate. However, carbonic acid rapidly decomposes to form water and carbon dioxide – the carbon dioxide causes the bubbles observed in this lab.

The procedure for this demonstration:

1) Pour 50 mL water into a 250 mL beaker.

2) Add 10 grams of sodium carbonate and stir until dissolved.

3) After putting on rubber gloves, *very quickly* pour 10 mL of 5 M HNO_3 into the sodium carbonate solution. A large amount of CO_2 will be formed.

4) Clean up: Add sodium bicarbonate to the beaker until all bubbling stops. Once neutralized, the solution can go down the drain. Sodium

bicarbonate should also be sprinkled on the desktop to neutralize any stray nitric acid that may have splattered out of the beaker.

Demonstration #3: Calcium and hydrochloric acid

$$Ca_{(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$$

Equipment:

- 1 small chunk of calcium metal
- 25 mL of 6 M HCl
- 100 mL beaker
- forceps
- rubber gloves

Safety:

Hydrochloric acid is caustic and should be treated with caution. Wear rubber gloves, goggles, and a lab coat when performing this station.

This demonstration gives off a considerable quantity of heat. Unfortunately, this may cause the hydrochloric acid dissolved in the solution to vaporize, forming acidic fumes. For this reason, do this demonstration in the fume hood.

Calcium metal can cause skin burns with prolonged exposure, though I've never observed this. As with any reactive metal, use caution.

Room destruction factor:

This reaction isn't terribly vigorous and no room-damaging compounds are produced. However, neutralize the HCl before pouring it down the sink, as acids damage metal pipes.

How the lab works:

In this lab, calcium metal will react with hydrochloric acid, forming dissolved calcium chloride and bubbles of hydrogen gas. This demonstration bubbles vigorously, making it interesting to watch.

The procedure for this demonstration:

1) Pour 25 mL of 6 M HCl into a beaker.

2) Add a small chunk of calcium. You will observe vigorous bubbling.

3) Clean up: When the bubbling stops, the calcium metal should be completely dissolved. Neutralize any residual hydrochloric acid in the beaker with sodium bicarbonate and pour the mixture down the sink.

Demonstration #4: Isopropanol and oxygen

$$2\ C_3H_8O_{(l)} + 9\ O_{2(g)} \xrightarrow{\Delta} 6\ CO_{2(g)} + 8\ H_2O_{(g)}$$

Equipment:

- 30 mL of 100% isopropanol in a small beaker
- A fireproof countertop (lab counters are fine)
- One paper towel
- One book of matches

Safety:

In this lab, 30 mL of isopropanol are spread over a countertop and lit with a match, forming carbon dioxide and water. The entire countertop catches fire, making this a dramatic demonstration. Make sure that all flammable chemicals and students are well away from the countertop during this demonstration and that the beaker containing the isopropanol has been removed before ignition.

When lighting the isopropanol, don't use the hand that you used to spread the isopropanol on the countertop. Doing so will cause your hand to catch fire. Just in case this should happen, have a nearby sink filled nearly to the top with water, in addition to all other normal laboratory and fire safety apparatus.

Do not wear rubber gloves during this demonstration! The heat generated by the combustion of isopropanol is enough to melt rubber gloves to your skin, causing very serious and hard-to-treat burns.

Room destruction factor:

As long as proper fire precautions are taken, there should be no damage.

How the lab works:

In this lab you will cover the countertop with a thin coat of isopropanol and light it with a match. The fire will burn for about 10 seconds before going out on its own.

Students should be cautioned against trying this demonstration at home. Most students know that isopropanol and rubbing alcohol is the same thing, so if you believe that any of your students pose a fire risk, tell them only the formula of isopropanol instead of mentioning it by name.

The procedure for this demonstration:

1) Pour 30 mL isopropanol on a fireproof countertop.

2) Wipe down the counter with a paper towel to form a thin layer of isopropanol. This lab is particularly impressive if you have a very long counter that can be fully covered with isopropanol.

3) Have a student turn off the lights.

4) Using the hand you didn't use to wipe down the countertop, light one end of the isopropanol. The desktop will rapidly catch fire from this point.

5) Clean up: Soak the paper towel in water and throw it away in the garbage.

<u>What can go wrong:</u>

- The counter doesn't light. This happens when too little isopropanol is used. Use more isopropanol or wipe down a smaller area of counter.

- You set yourself on fire. Put yourself out by dunking the burning part into water. Isopropanol burns at a relatively low temperature, giving you 1-2 seconds before you need to worry about burns.

Demonstration #5: Methane and oxygen

$$CH_{4(g)} + 2\ O_{2(g)} \xrightarrow{\Delta} CO_{2(g)} + 2\ H_2O_{(g)}$$

<u>Equipment:</u>

- methane gas outlet
- 0.75 meters of flexible tubing
- book of matches
- bucket or plugged sink
- 10 grams of dishwashing liquid
- 2 liters of tap water

Safety:

In this demonstration you'll be igniting a pile of soap bubbles containing methane gas, resulting in a fireball ~1 m in diameter. Because there is a large quantity of gas burning at once, make sure all students are moved at least three meters from the demonstration. As always, make sure everybody in the room is wearing goggles.

Make sure that all flammable materials are moved three meters from the demonstration to eliminate the risk of fire. If you have mobiles or other student work hanging from the ceiling, clear it away from the area you're doing the lab, as the fireball easily reaches the ceiling. Make sure you have a fire extinguisher handy in case this demonstration gets out of control.

Room destruction factor:

There is a remote possibility that this demonstration could cause the fire sprinklers in your classroom to go off. Before doing this demonstration within three meters of a fire sprinkler, have a talk with your building manager to ensure that the sprinklers won't go off during your demonstration. This is, as one might imagine, a very interesting conversation. Sprinklers can be temporarily shielded from the heat by covering them with a plastic drinking cup, though you should check to ensure that this isn't a violation of the fire code in your area before trying this.

How the lab works:

For this demo you'll be lighting methane soap bubbles on fire, creating carbon dioxide and water vapor. A large fireball is also created, making this a very dramatic and exciting demonstration. Expect to hear your students clamoring for you to do it again.

The procedure for this demonstration:

1) Place the dishwashing liquid and water into the bucket or plugged-up sink.

2) Run the flexible tubing from the gas outlet into the bucket so the end is immersed in the soapy water.

3) Turn on the gas and let soap bubbles form for about 30 seconds.

4) Turn off the gas and room lights.

5) Light a match and throw it at the pile of bubbles from approximately 1.5 meters away. A large fireball will form and rise to the ceiling. The residual

bubbles in the sink may take up to ten seconds to completely finish burning.

6) There is no clean-up for this demonstration, as all of the products are gaseous.

What can go wrong:

- The fire sprinklers or alarm go off. If this occurs, leave the room and inform your school administration that it's a false alarm.
- Something catches fire. If it's a stack of papers, put them out. If it's a bottle of flammable chemicals, get everybody out and pull the fire alarm!

Solutions for the Writing Complete Equations Lab

Reaction 1: $AgNO_{3(aq)} + HCl_{(aq)} \rightarrow AgCl_{(s)} + HNO_{3(aq)}$

Reaction 2: $2\ HNO_{3(aq)} + Na_2CO_{3(s)} \rightarrow 2\ NaNO_{3(aq)} + H_2O_{(l)} + CO_{2(g)}$

Reaction 3: $Ca_{(s)} + 2\ HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_{2(g)}$

Reaction 4:

$$2\ C_3H_8O_{(l)} + 9\ O_{2(g)} \xrightarrow{\Delta} 6\ CO_{2(g)} + 8\ H_2O_{(g)}$$

Reaction 5:

$$CH_{4(g)} + 2\ O_{2(g)} \xrightarrow{\Delta} CO_{2(g)} + 2\ H_2O_{(g)}$$

Questions:

1) Writing chemical equations is better than writing the reaction in paragraph form because you can convey a lot of information in a brief statement.
2) (aq) means that something has dissolved in water, while (l) means that it is a liquid. For example, dissolved NaCl is (aq), while isopropanol is (l).

Solutions to the Writing Complete Equations Worksheet

1)

$$2\ C_2H_{2(g)} + 3\ O_{2(g)} \xrightarrow{\Delta} 2\ CO_{2(g)} + 2\ H_2O_{(g)} \quad \Delta H \text{ is negative}$$

2)

$$4\ Fe_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Fe_2O_{3(s)}$$

3)

$$H_2SO_{4(aq)} + 2\ NH_{3(g)} \rightarrow (NH_4)_2SO_{4(aq)}$$

4)

$$Ca_{(s)} + 2\ H_2O_{(l)} \xrightarrow{70^0\ C} Ca(OH)_{2(s)} + H_{2(g)} \quad \Delta H \text{ is negative}$$

5)

$$2\ NaHCO_{3(s)} \xrightarrow{1200^0\ C} Na_2CO_{3(s)} + CO_{2(g)} + H_2O_{(g)}$$

Common Symbols Found in Equations

The following symbols are used to indicate the state of a chemical in a reaction:

- **(s)** written below a chemical formula means the chemical is a solid. For example, $Na_{(s)}$ means that sodium is a solid.
- **(l)** written below a chemical formula means the chemical is a liquid. For example, $H_2O_{(l)}$ means that water is a liquid.
- **(g)** written below a chemical formula means the chemical is a gas. For example, $O_{2(g)}$ means that oxygen is a gas.
- **(aq)** written below a chemical formula means the chemical is dissolved in water. For example, salt water would be written as $NaCl_{(aq)}$.

The following symbols are written above or below the arrow in a chemical equation to indicate what reaction conditions are required to make the reaction proceed:

- **“D”** means that you heat the reagents to make the reaction take place. Typically, this indicates extremely high temperatures, as you would find over a Bunsen burner.
- Any temperature value (for example, “100^0 C”) specifies the temperature required to perform the reaction.
- Any pressure value (for example, “760 mm Hg”, “1.5 atm”, or “110 kPa”) indicate the pressure required to perform the reaction.
- Chemical formulas written below the arrow indicate the solvent needed to perform the reaction.
- Any other words either above or below the arrow are explicit instructions for performing the reaction. If the word “boil” is written above the arrow, the reaction mixture should be boiled.

The symbol “D**H**” is sometimes written after a chemical reaction to indicate that heat has been given off or absorbed during a chemical reaction. If this value has a positive sign, the reaction is endothermic, meaning that it absorbs heat and feels cold. If this value has a negative sign, the reaction is exothermic (it releases heat) and feels hot. The larger the number, the more exo- or endothermic the reaction.

Writing Complete Equations Lab

For the following demonstrations, write complete equations that show what happened in each reaction. To receive full credit, each equation:

- Should have all chemical formulas written correctly.
- Should be balanced correctly.
- Should have symbols indicating the states of all products and reagents.
- Should have symbols indicating the necessary reaction conditions.

Lab: Please write the complete, balanced equation for each of the demonstrations shown to you in class.

Reaction 1:

Reaction 2:

Reaction 3:

Reaction 4:

Reaction 5:

Questions:

1) Do you think this method of writing chemical equations is better or worse than writing them out in paragraph form? Explain why or why not.

2) Explain the difference between the (aq) and (l) symbols.

Writing Complete Equations Worksheet

Write complete, balanced equations for the following processes, using all of the symbols you've learned in class. When a compound is described as being "dissolved", assume it is dissolved in water.

1) When acetylene gas (C_2H_2) is heated with oxygen gas, carbon dioxide gas, water vapor, and heat are formed.

2) When iron powder reacts with oxygen in the atmosphere, it forms iron (III) oxide (also known as "rust").

3) When dissolved sulfuric acid (H_2SO_4) reacts with gaseous ammonia, dissolved ammonium sulfate is formed.

4) When calcium metal reacts with liquid water at a temperature of 70^0 C, calcium hydroxide powder, hydrogen gas, and heat are formed.

5) When sodium bicarbonate powder is heated at a temperature of 1200^0 C, it decomposes to form solid sodium carbonate, gaseous carbon dioxide, and water vapor.

Chapter 12 – Simple Stoichiometry

Overview

Simple stoichiometry consists of convertinvg grams of reagent to grams of product for a chemical reaction. In this chapter, students will learn how to perform stoichiometry calculations and use them to determine the percent yield of the reaction between sodium hydroxide and acetic acid.

Teaching Simple Stoichiometry

Stoichiometry is the process of determining the relationship between grams of product and reagent in a chemical reaction. These calculations take the form of a t-chart, as shown in the student handout. Students generally dislike stoichiometry because success hinges on balancing the equation correctly and performing a three step mathematical calculation. As we've mentioned before, most students dislike math.

Fortunately, it's not difficult to make students comfortable with stoichiometric calculations. Some suggestions:

1) Above all else, tell students that stoichiometry is easy! If your students can multiply and divide, it's a simple matter to solve stoichiometry problems. Of course, getting students to the point that they can reliably set up the problems may be more difficult, but if you go through the steps slowly in class, they suddenly look simple.

2) Break stoichiometry into a series of smaller, easier steps! The handout in this chapter gives an idea of how this can be done. By treating what looks like a very difficult, very large problem as a series of small and easily-understood problems, it becomes far easier.

3) Do lots of labs! Students find repeated worksheets tedious, so you may want to try doing several short labs involving stoichiometry. Any reaction will do – all you need to do is give students a stoichiometric quantity of the two reagents and let them calculate how much product should be formed.

4) Be very patient! While some students will learn how to perform stoichiometry calculations perfectly on the first try, others will need to have every step explained to them for a couple of days. Remember, no matter how frustrating it is for you to repeat the same instructions over again, it's doubly so for the students who don't understand what you're talking about.

One aspect of stoichiometry calculations we explore in this lab is that of percent yield. Though stoichiometry calculations allow us to determine how much product should be formed during a chemical reaction, they can't tell us how much product will actually be formed. The reason is experimental error.

I suggest reviewing experimental error before performing this lab. The idea that a yield may be below 100% because some of the product may not have been formed, or because some of the reagents splashed out of the container is reasonable to most students. The idea that a yield may be above 100% because other compounds have been inadvertently added to the reaction mixture may make less sense. To explain this, use an analogy:

> *While cooking a pizza, I used one pound of dough to make the crust, ½ pound of tomato sauce, ½ pound of cheese, and ½ pound of other toppings. When I took the pizza out of the oven, it weighed 3 pounds, giving me a yield of 120%. The mystery was solved when I bit into the pizza – apparently, cockroaches had infiltrated the pizza while I was turning the oven on.*

OK, I admit it – this example is disgusting. However, it also makes the point that it is possible to get more than 100% yield if there are glaring enough experimental errors. It's also amusing, which helps make stoichiometry a little less painful for math-phobic students.

One final thing: You've mentioned t-chart calculations to your students before, mentioning that the only thing you ever write in front of the word "moles" in a t-chart is "1". Stoichiometry calculations provide an exception to this rule because of the mole ratio. The new rule should be: The only time you ever write anything other than "1" in front of moles is in the mole ratio step.

Doing the Percent Yield Lab

Equipment:

- 8-100 mL beakers
- 6 beaker tongs
- 1 liter distilled water
- 6 stirring rods
- 200 mL glacial acetic acid
- 6 Bunsen burners
- 2 balances
- 6 wire gauze sheets
- 300 grams NaOH pellets
- 6 rings
- 10 disposable pipets

- 6 ring stands
- 5 spatulas
- 2 flint strikers
- 6-50 or 100 mL graduated cylinders

Safety:

Sodium hydroxide pellets are extremely corrosive to flesh, making it very important that students wear goggles at all times. If students notice that their hands feel slippery, have them wash their hands with household vinegar to neutralize the sodium hydroxide. If students get NaOH in their eyes, have them rinse their eyes at the eyewash while you contact medical personnel.

The acetic acid used in this lab may cause eye damage and skin irritation. If any gets into the eyes, rinse at the eyewash and contact medical personnel. For skin contact, have students rinse their hands with saturated sodium bicarbonate solution.

Bunsen burners may cause burns. I recommend that students not wear rubber gloves during this lab because of the danger that hot equipment may cause them to melt to the skin. Remind students that hot and cold glassware look the same.

If students boil their sodium acetate solution for too long, their beaker will crack or shatter. If this happens, have your students wait five minutes before collecting the hot glass fragments.

Room destruction factor:

This lab generally doesn't cause damage to the lab. About the worst thing that can happen is that the sodium acetate solution formed in this lab gets on the countertops and recrystallizes. These crystals pose no significant safety hazard, though they can be hard to clean off of the countertop.

How the lab works:

In this lab, students will add sodium hydroxide pellets to an acetic acid solution. The resulting sodium acetate solution will then be boiled until the water is removed, leaving behind the sodium acetate powder. When the students compare the amount of sodium acetate recovered during this lab to the amount they determined using stoichiometry, they should be able to determine their percent yield of sodium acetate.

In step 3, students are asked to weigh 5 grams of acetic acid. Because acetic acid is a liquid at room temperature, some may measure 5 mL into a graduated cylinder, rather than using a balance. Let your students know that 1 mL = 1 gram only for pure water, and that this rule doesn't hold for all substances.

In step 5, students boil the sodium acetate solution until all of the water has been removed. This step is the source of most error in this lab – either students boil the sodium acetate too long and shatter the beaker or they don't boil it long enough, leaving behind residual water (and giving them a 100+% apparent yield of sodium acetate. Caution students about the danger of heating this solution too long, but don't warn them about the dangers of not heating it long enough, as their greater than 100% yield provides an important lesson about experimental error.

What can go wrong:

- The beaker breaks while heating. If the beaker merely cracks, this shouldn't affect the results of this lab. If the beaker breaks apart, the fragments can still usually be weighed to determine percent yield, as long as the sodium acetate didn't spray over too large an area.
- The beaker boils over. The main source of this problem, interestingly enough, is that there is residual soap left over from the last time the beaker was washed. Have your students try the lab again, making sure the beaker is completely rinsed out. A lower heat setting may also solve this problem.
- The solution splatters when boiling. This is normal, but should be commented on by the students as a possible source of error.
- The percent yield is greater than 100%. As mentioned above, this is caused by residual water left over when the beaker is heated.
- The percent yield is less than 100%. This can be because they made large procedural errors (spills, etc.) or because their stoichiometry / percent yield calculations are incorrect.
- Percent yield is exactly 100%. Though it sometimes happens, perfect data sounds a little bit fishy. Check the data to ensure that your students didn't cook their data to make themselves sound perfect.
- The water evaporates during heating, then returns. The second appearance of "water" isn't water at all, but melted sodium acetate. Once the sodium acetate has begun to melt, students can be sure that the water is completely removed. If this melted sodium acetate continues to be heated, eventually it turns brown/black, which indicates that it has started decomposing.

Clean up:

All glassware may be washed with soap and water, as sodium acetate is not hazardous. Make sure goggles are worn while washing the glassware, as there may be some slight acid or base residue.

Solutions for the Percent Yield Lab

Pre-lab: Boiling the water away is a good method for performing this lab, as water is the only product other than sodium acetate formed during this reaction.

Lab procedure:

Step 1: Units should be given as “grams”

Step 4: The solution will get much warmer due to the heat of solvation of sodium hydroxide. The solution may also appear temporarily hazy because of concentration gradients around the pellets.

Step 6: Both weights should be given in “grams”.

Questions:

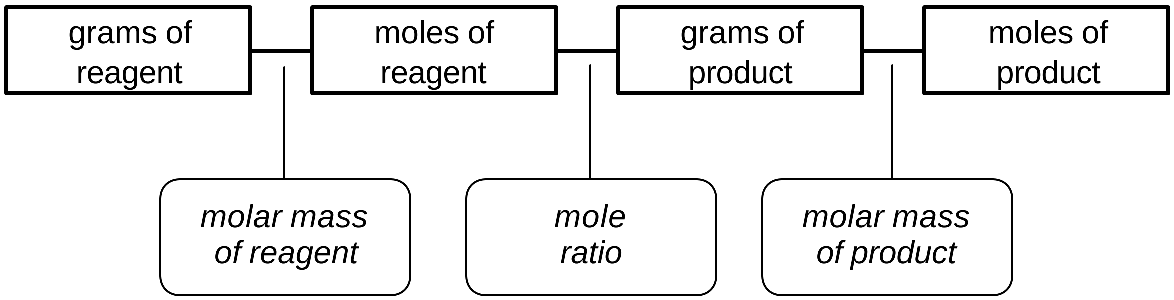
1) 6.83 grams
2) Their percent yield will be equal to their actual yield divided by 6.83, then multiplied by 100%. Typical yields are between 80-150%, depending on the amount of splattered solution and amount of leftover water remaining.
3) Percent yields less than 100% can usually be explained by splattered solution. Percent yields greater than 100% are explained by the presence of residual water in the sodium acetate. Percent yields of exactly 100% should be checked, as this is an unlikely answer – advanced students may recognize that 100% yield doesn’t indicate perfection but rather offsetting positive and negative errors.

Solutions to the Stoichiometry Calculations Worksheet

1) 30.3 grams
2) 13.9 grams
3) 7.4 grams
4) 193.5 grams
5) 79.9 grams

<u>Stoichiometry Calculations</u>

It can be hard to remember how to do stoichiometry calculations. Fortunately, there's an easy way you can remember what to do. Use the map below to get consistently correct answers:



To use this map, you need to identify where you are starting in the diagram and where you need to you. This is just like taking a long car trip – before you can take the trip, you need to look at a map and figure out how you're going to get to your destination. To make the necessary conversions, you need to perform one calculation for every box you travel through. For example, if you want to go from grams of reagent to grams of product, you'll do three calculations. For each calculation, the conversion factor you need to use is displayed between the box you're moving from and the destination box.

Let's try an example:

$$2\ FeO + O_2 \rightarrow Fe_2O_3$$

If we start with 40 grams of O_2, how many grams of Fe_2O_3 can we make?

Step 1: Make sure you're starting with a balanced equation

In this example, you were given a balanced equation. However, you might sometimes need to figure it out for yourself. Without a balanced equation, you're guaranteed to get the wrong answer!

Step 2: Find your starting and ending points on the map

In our case, we've been given 40 grams of O_2, which is one of our reagents, so we'll start in the "grams of reagent" box. We want to find the number of grams of Fe_2O_3, so we'll end in the "grams of product" box.

Step 3: Figure out how many calculations you need to do.

Our first calculation will be to jump from "grams of reagent" to "moles of reagent". The second calculation will be from "moles of reagent" to "moles of product". The final calculation will be from "moles of product" to "grams of product".

Step 4: Do the calculations by setting up t-charts

In calculation #1, we go from "grams of reagent" to "moles of reagent". To do this, follow the standard procedure for a t-chart: Put the thing you've been given on the top left, put the units of what you've been given on the bottom right, put the units of what you want on the top right, and put the conversion factors in on the right side. In our case, we were given "40 grams O_2" (put in top left), the units of which are "grams O_2" (bottom right). The units of what we want are "moles O_2" (top right). The unit conversion factor is the molar mass of oxygen, which is 32 grams/mole. This calculation is shown below:

$$\begin{array}{c|c} \text{40 grams } O_2 & \text{1 mole } O_2 \\ \hline & \text{32 grams } O_2 \end{array} \quad \mathbf{= 1.25\ moles\ O_2}$$

In the second calculation, we use the output from the first as what we've been given. Our conversion factor is the mole ratio, which is defined as the coefficient of what we want to find in the equation over the coefficient of what we were given. In this case, it is "1 mole Fe_2O_3" divided by "1 mole O_2". This calculation is shown below:

$$\begin{array}{c|c} \text{1.25 moles } O_2 & \text{1 mole } Fe_2O_3 \\ \hline & \text{1 mole } O_2 \end{array} \quad \mathbf{= 1.25\ moles\ Fe_2O_3}$$

In the third calculation, we use the output from calculation 2 as what we've been given. Our conversion factor is the molar mass of Fe_2O_3, or 159.7 grams/mole. This calculation is shown below:

$$\begin{array}{c|c} \text{1.25 moles } Fe_2O_3 & \text{159 grams } Fe_2O_3 \\ \hline & \text{1 mole } Fe_2O_3 \end{array} \quad \mathbf{= 199.6\ grams\ Fe_2O_3}$$

The final answer to this problem is 199.6 grams of Fe_2O_3.

All stoichiometry problems may be solved using this method, although it will not be necessary to do all of them using three calculations. Make sure you do only as many calculations as you need to, depending on how far apart your starting and ending points are. As with any trip, the fewer creative short cuts you make, the more likely you are to have a good time.

Percent Yield Lab

In class, we've been talking about stoichiometry, the predictionof reaction masses using chemical equations. Today, we'll get a chance to see how this works for a real chemical reaction.

In this lab, you will be mixing acetic acid with sodium hydroxide to form sodium acetate and water. The equation for this reaction is:

$$C_2H_3O_2H + NaOH \rightarrow H_2O + NaC_2H_3O_2$$

You will start with a known quantity of acetic acid and sodium hydroxide and will predict the mass of sodium acetate formed. You will then measure the amount of sodium acetate actually produced by boiling the reaction mixture to dryness.

Pre-lab question: Is boiling the water to isolate the sodium acetate the best way to measure the amount of sodium acetate formed? Explain why or why not.

Lab procedure:

1) Weigh a 100 mL beaker. The weight of your beaker was: ______________

2) Pour 25 mL distilled H_2O into a 100 mL beaker using a graduated cylinder.

3) Carefully place 5 grams of acetic acid into the beaker. Be careful – even though they smell the same, acetic acid is <u>much</u> stronger than vinegar.

4) Weigh approximately 3.3 grams of sodium hydroxide and place it into the beaker containing the water and acid. Stir until all of it has dissolved. Write your observations below:

5) Place the 100 mL beaker over a Bunsen burner and heat slowly until all of the water has boiled away. Be careful not to heat it past this point, because the beaker will shatter.

6) Once the bekaer has cooled for 5 minutes, weigh it again. The difference between this weight and the weight of the empty beaker is equal to the weight of sodium acetate you produced.

Weight of the beaker with sodium acetate: _______________

Weight of the sodium acetate (actual yield): _______________

7) Clean up: Wash all residue from the beaker with soap and water.

Questions:

1) From the weight of acetic acid you started with in this reaction, use your knowledge of stoichiometry to predict how much sodium acetate you should have produced (theoretical yield). Show your work!

2) Calculate your percent yield of sodium acetate for this reaction, using the actual yield you found and the theoretical yield you calculated in question 1 above. Show your work.

3) Was your percent yield exactly 100% If it was less than 100%, explain where you think the lost sodium acetate went. If it was more than 100%, explain why you have more apparent product than the law of conservation of mass allows. For full credit, your explanations should be complete!

Stoichiometry Calculations Worksheet

Use your knowledge of stoichiometry to answer the following questions:

1) How much calcium oxide is formed when 21.7 grams of calcium reacts with an excess of oxygen?

2) How much magnesium hydroxide is needed to react completely with 30 grams of nitric acid?

3) How much copper (II) sulfate is formed when 2.95 grams of copper is dissolved in pure sulfuric acid?

4) How much lithium oxide is produced by the reaction of 90 grams of lithium metal and excess oxygen gas?

5) How much carbon dioxide is formed when 23 grams of $C_{12}H_8$ is burned in an excess of oxygen gas?

Chapter 13 – Limiting Reagent Lab

Overview

Limiting reagent problems are extensions of simple stoichiometry problems. Instead of finding out how much of a product you can make from a given amount of reagent, they involve determining how much product can be made using an uneven mixture of starting materials. In this lab, students will learn how to solve limiting reagent problems and will apply this knowledge to the precipitation reaction between lead (II) nitrate and potassium iodide.

Teaching about limiting reagents

As mentioned above, limiting reagent problems are nothing more than stoichiometry problems in which you have an uneven amount of each reagent. As a result, one of the reagents (the limiting reagent) will run out before the other does, limiting the amount of product that can be formed.

It's important to emphasize to students that limiting reagent problems are done in exactly the same way as other stoichiometry problems they've solved. The only difference is that two sets of calculations need to be done instead of one. If your students can perform a simple stoichiometry problem, they should have no difficulties with limiting reagent problems.

In fact, students already know what a limiting reagent is. Give your students this example:

> *I really like cheese sandwiches. The recipe I use to make cheese sandwiches is to put one slice of cheese between two slices of bread. If I have 50 slices of bread and 10 slices of cheese, how many cheese sandwiches can I make?*

Most students will easily recognize that you can make 10 cheese sandwiches with the materials provided. Ask them why you can't make 25 sandwiches even though you have 50 slices of bread, and they'll easily identify that after 10 sandwiches, you've run out of cheese. Cheese, in this case, is the limiting reagent.

Once students understand the idea behind limiting reagent problems, they usually have a much easier time figuring out how to solve them. Let's see another example, this time using chemicals instead of food:

The synthesis of potassium iodide from potassium metal and iodine crystals is shown below:

$$2\ K_{(s)} + I_{2(s)} \rightarrow 2\ KI_{(s)}$$

If we start with 20 grams of potassium and 40 grams of iodine, solve to find:

a) What our limiting reagent is.
b) How many grams of potassium iodide should be formed.
c) How many grams of our nonlimiting reagent are left over after the reaction.

Parts a and b can be solved at the same time. To do this, calculate the number of grams of potassium iodide that can be formed using each of the reagents, assuming an excess of the other:

- If we assume we have an excess of iodine and 20 grams of potassium, we'll find that we should be able to make 84.3 grams of potassium iodide (do this calculation using the t-chart method outlined in Chapter 12).
- If we assume we have an excess of potassium and 40 grams of iodine, we'll find that we should be able to make 52.4 grams of potassium iodide.

Which value is right? Let's go back to the cheese sandwich example for help. When we solved this problem in our heads, we found that 50 pieces of bread corresponded to 25 cheese sandwiches, while 10 slices of cheese correspond to 10 sandwiches. In this case, the smallest answer is the one that's correct, and the reagent that resulted in this answer is the limiting reagent.

The same is true in the potassium iodide example. Since 52.4 grams is smaller than 84.3 grams, we can only make 52.4 grams of potassium iodide. The reagent that allowed us to generate this answer was iodine, our limiting reagent.

Now for part c – how much of the nonlimiting reagent is left over?

This calculation can be easily determined using the following formula:

$$\textbf{Quantity of leftover nonlimiting reagent} = \textbf{The original quantity of nonlimiting reagent} - \textbf{The original quantity of nonlimiting reagent}\left(\frac{\textbf{the smaller predicted quantity of product}}{\textbf{the larger predicted quantity of product}}\right)$$

I'll admit, it doesn't look like a very simple formula. However, let's see how our information plugs into it:

The original quantity of nonlimiting reagent (potassium) is 20 grams. The smaller predicted quantity of product was found in our limiting reagent calculation to be 52.4 grams. The larger predicted quantity was found to be 84.3 grams. As a result, the quantity of leftover limiting reagent is [20 – 20 (52.4 / 84.3)], or 7.57

grams. Though this calculation isn't trivial, most students find it's easier than doing a third stoichiometry calculation to determine how much potassium is needed to make 52.4 grams of potassium iodide, then subtracting this quantity from the 20 grams.

Let's break down what we did into easily learned steps:

1) We figured out how much product we could make, based on reagent 1.
2) We figured out how much product we could make, based on reagent 2.
3) We looked at the two answers and determined that the smaller of the two answers corresponds to the maximum amount of product that can be formed. This is because one of the reagents runs out after this much product is made, stopping the reaction.
4) After this reaction took place, we can determine the amount of nonlimiting reagent left over by using the equation on the last page.

Some students will believe that the reagent with the smallest mass will always be the one that is the limiting reagent. By discussing that this isn't always the case (in fact, it wasn't even the case in our example above), you can show them that they can't use the masses of the reagents as a shortcut to the right answer. After all, it's the number of moles of each compound that matters, not the mass!

Doing the Limiting Reagent Lab

Equipment:

- 1 liter distilled water
- 30 grams potassium iodide
- 30 grams lead (II) nitrate (usually just referred to as "lead nitrate")
- 12-250 mL beakers
- 2 spatulas
- 6 stirring rods
- 6 funnels (preferably plastic)
- 10 filter papers, sized to fit the funnels
- 2 balances
- 6-100 mL graduated cylinders
- 10 watch glasses
- laboratory oven (hotplate may be substituted)
- 6 forceps

Safety:

Lead (II) nitrate and lead (II) iodide are toxic. Students should wash their hands well after this lab to ensure against accidental ingestion. If students ingest any lead salts, induce vomiting and call for immediate medical assistance.

Nitrates are explosive when heated to high temperatures. For this reason, never heat any chemical in this lab over 100^0 C. Though the quantities used in this lab pose limited danger, it's better to be safe than sorry.

Take care to keep flammable liquids and other materials away from the hot oven, as this poses an ignition hazard. Never store anything on top of an oven except for unused oven racks.

Room destruction factor:

The yellow lead (II) iodide precipitate does a good job of coating most laboratory surfaces and can be hard to remove. Make sure you tell your students that they need to have it completely cleaned up before leaving their lab area.

How the lab works:

In this lab, student pour lead (II) nitrate solution into a potassium iodide solution. Both compounds are colorless, but combine to form a yellow lead (II) iodide precipitate. It's a dramatic reaction that always impresses the students. Once the precipitate has formed, it is isolated by filtration. This filtrate is then dried in an oven and weighed to find the yield. This actual yield will then be compared to the theoretical yield to find a percent yield for the reaction.

Students may have trouble recalling how to use filter paper to isolate a precipitate, so you may want to review the procedure before starting the lab. Make sure students are reminded not to pour the solution higher than the level of the filter paper in the funnel, as this will cause unfiltered solution to pass into the collecting beaker.

In step 10, students place their wet filter paper into the oven to remove water. Preheat the oven before starting this lab to speed this process as much as possible. If time is short, students can speed the process even more by pressing their wet filter paper flat between two paper towels to remove some of the residual water. It is also important that students be reminded to write their names on their filter paper in pencil to ensure they collect the right one when it is done drying.

The calculations in this lab are similar to those in the student handout and in the example we discussed earlier. If students have questions about these calculations, point them toward their notes before giving them a hand. Many students ask questions because they're afraid of being wrong, even if they basically know what they're doing.

What can go wrong:

- The lead (II) iodide precipitate gets through the filter paper and into the receiving beaker. To solve this problem, students need to refilter the solution.
- The filter paper is folded incorrectly, causing it to filter too slowly. Alternately, students sometimes poke at the filter paper and tear small holes in the bottom. If the solution is filtering too slowly, students just need to stick with it until the end. If they have torn holes in the filter paper, the solution will need to be refiltered.
- Student yields are greater than 100%. As with the stoichiometry lab (Chapter 12), this is most likely because there is residual water in the filtrate.

Clean up:

The lead (II) iodide precipitate should be collected and dried for later disposal by hazardous waste disposal personnel. Never throw away lead waste! The filtrate may be poured down the sink.

Solutions for the Limiting Reagent Lab

Pre-lab:

Steps 2 and 4: Units should be given in "grams".
Step 5: A bright yellow precipitate will form as lead (II) nitrate is added.
Step 7: Units should be given as "grams".
Step 12: Units should be given as grams.
Step 13: The amount of lead (II) iodide formed is calculated by subtracting their step 7 answer from their step 12 answer.

Questions:

1) 0.8 grams of KI can produce 1.39 grams of PbI_2.
1.0 grams of $Pb(NO_3)_2$ can produce 1.11 grams of PbI_2.
The smaller of these two answers, **1.11 grams,** of PbI_2 will be formed.
2) The limiting reagent was lead nitrate.
3) The amount of potassium iodide left over is 0.20 grams.
4) Percent yield is equal to actual yield divided by theoretical yield, times 100%. In this reaction, the percent yield will be equal to the grams of PbI_2 formed times 90.1.
5) If the yield was less than 100%, some of the lead (II) iodide probably passed through the filter paper. If the yield was more than 100%, there was probably residual water in the filter paper. If the yield was exactly 100%, check to make sure that the data hasn't been made up by your students, as this is difficult to achieve in this lab. Strong students may correctly indicate that a 100% yield is more indicative of offsetting errors than of perfection.

Solutions to the Limiting Reagent Worksheet

1a) ammonium nitrate
1b) 18.6 grams of ammonium phosphate, 31.9 grams of sodium nitrate
1c) 29.5 grams of sodium phosphate

2a) iron (III) phosphate
2b) 46.3 grams of calcium phosphate, 43.8 grams of iron (III) carbonate
2c) 54.0 grams of calcium carbonate

Limiting Reagent Handout

Before now, we've always done stoichiometry calculations in which we had exactly the right amount of both reagents for a chemical reaction to take place. For example, for the reaction 2 Na + Cl_2 → 2 NaCl, we have previously assumed that we would automatically have sodium and chlorine in a 2:1 ratio.

However, what happens if we don't have exactly the right proportions of each reagent? What would happen if we had 2 moles of sodium and only 0.5 moles of chlorine?

In the real world, this is a very common problem. After all, you can't expect that your chemical stockroom will have exactly 2 moles of sodium for every mole of chlorine on the shelf. As a result, we've come up with the idea of a limiting reagent. Basically, the limiting reagent for a chemical reaction is the one that runs out first, limiting the amount of product that can be formed.

Think of this in terms of a barbeque. Let's say we want to make hamburgers. When you look in your cupboard, you find that you have 10 hamburger buns. Unfortunately, your refrigerator has only 6 hamburger patties. How many hamburgers can you make?

The answer, of course, is six. After all, you can't make a hamburger without a hamburger patty. The same goes for chemical reactions – if you run out of one of the reagents (the limiting reagent), the chemical reaction stops.

Fortunately, limiting reagent problems don't require that you learn a new form of stoichiometry. All you need to do is to determine the amount of product that each reagent can form – the one that results in the smallest answer is your limiting reagent, and the small answer indicates the maximum amount of product that can be made using the reagents at hand.

Example:
For the reaction between sodium and chlorine, how many moles of sodium chloride can I make if I start with 2 moles of sodium and 0.5 moles of chlorine?

Answer:
Do two stoichiometry calculations to solve this problem. The first problem will determine how much product can be made with 2 moles of sodium (assuming plenty of chlorine) and the second problem will determine how much product can be made with 0.5 moles of chlorine (assuming plenty of sodium). The smaller of these answers tells you how much sodium chloride will actually be made, and the reagent that resulted in this answer is the limiting reagent!

For this example, 2 moles of sodium will form 2 moles of sodium chloride, and 0.5 moles of chlorine will form 1 mole of sodium chloride. As a result, chlorine is the limiting reagent, and the reaction will form 1 mole of sodium chloride.

Limiting Reagent Lab

In this lab, you will be mixing lead (II) nitrate with potassium iodide to produce solid lead (II) iodide and aqueous potassium nitrate. The equation for this reaction is:

$$Pb(NO_3)_{2(aq)} + 2\ KI_{(aq)} \rightarrow PbI_{2(s)} + 2\ KNO_{3(aq)}$$

You will be given a known quantity of both reagents and will have to predict how much lead (II) iodide precipitate will be produced. You will measure the amount of lead (II) iodide produced by filtering it from the aqueous solution.

Pre-lab: Is filtering the lead (II) iodide solid out of the solution the best way to measure the amount of precipitate formed? Explain why or why not.

Procedure:

1) Pour 75 mL of distilled water into a 250 mL beaker with a graduated cylinder.

2) Using a watch glass, weigh approximately 1.0 grams of potassium iodide on the balance. Write down the precise amount of KI below:

   The weight of KI used was: ___________________

3) Add the KI you measured into the distilled water and stir until dissolved.

4) Clean any remaining KI from the watch glass and weigh approximately 0.8 grams of lead (II) nitrate on the balance. Write down the precise amount of $Pb(NO_3)_2$ below:

   The weight of $Pb(NO_3)_2$ used was: ___________________

5) Add the lead (II) nitrate you measured into the distilled water and write your observations below:

6) Stir the mixture with a stirring rod to ensure that all of the lead (II) nitate and potassium iodide have dissolved. Let the mixture sit for two minutes.

7) Weigh a piece of filter paper on a balance and record its mass:

Weight of the filter paper: ___________________

8) Place a second collecting beaker below your funnel and filter the solution through the filter paper to collect the lead (II) iodide precipitater. This solution may be poured down the sink.

9) If there are still any chunks of lead (II) iodide still stuck to the 250 mL beaker, rinse it into the filter paper with a small amount of distilled water.

10) When most of the lead (II) iodide ahs been collected in the filter paper, remove the filter paper from the funnel with forceps and place it in the hot oven. This will cause the water to evaporate from the paper.

11) After heating the filter paper for 15 minutes in the oven, check to see if it is still wet. If it is, heat another 5 minutes before proceeding to step 12.

12) Weigh the filter paper and precipitate on the balance. By comparing the weight of the filter paper that contains the precipitate to the original weight of the filter paper, you can find the mass of lead (II) iodide present. When you are finished, place both the paper and precipitate in the waste beaker.

Weight of the filter paper containing precipitate: ___________________

13) Calculate the weight of lead (II) iodide produced using your data.

Weight of PbI_2 produced: ___________________

Questions:

1) From the weight of lead (II) nitrate and potassium iodide used in this reaction, calculate the amount of lead (II) iodide that should have been formed (your theoretical yield). Show both calculations!

2) What was the limiting reagent in this reaction?

3) How much of the other reagent was left over at the end of this reaction?

4) Calculate your percent yield of lead (II) iodide for this reaction, using the actual yield you found and the theoretical yield calculated in question 1. Show your work.

5) Was your percent yield exactly 100%? if it was less than 100%, explain where you think the lost lead (II) iodide went. If it was more than 100%, explain why you had more PbI_2 than predicted.

<u>Limiting Reagent Worksheet</u>

For the following reactions, find the following:

a) Which of the reagents is the limiting reagent?
b) What is the maximum amount of each product that can be formed?
c) How much of the other reagent is left over after the reaction is complete?

1) Consider the following reaction:

$$3\ NH_4NO_3 + Na_3PO_4 \rightarrow (NH_4)_3PO_4 + 3\ NaNO_3$$

Answer the questions above, assuming we started with 30 grams of ammonium nitrate and 50 grams of sodium phosphate.

2) Consider the following reaction:

$$3\ CaCO_3 + 2\ FePO_4 \rightarrow Ca_3(PO_4)_2 + Fe_2(CO_3)_3$$

Answer the questions at the top of this sheet, assuming we start with 100 grams of calcium carbonate and 45 grams of iron (II) phosphate.

Chapter 14 – Heat Capacity of Water Lab

Overview

The amount of energy it takes to raise the temperature of a chemical compound can be expressed by the equation $\mathbf{\Delta H = mC_p\Delta T}$, where $\mathbf{\Delta H}$ represents the amount of energy absorbed or released, **m** is the mass of the compound in grams, $\mathbf{C_p}$ is the heat capacity of the compound, and $\mathbf{\Delta T}$ is the temperature change. In this chapter, students will learn to perform heat calculations using this equation. In the lab, they will calculate the heat capacity of water.

Teaching heat calculations

The main problem that students have with heat calculations is neither the math nor the algebra – both are straightforward. The problem that students have is with the concept of heat capacity.

Heat capacity, as you know, is the amount of energy required to heat one gram of a substance by one degree celsius. Use examples to demonstrate this to students. For example, which would you rather use to pull a pan from a hot oven, an oven mitt or a sheet of aluminum foil? Most students will intuitively realize that the aluminum foil will transmit the heat easily while the oven mitt is a much better insulator. The reason: Oven mitts have a higher heat capacity than aluminum.

The equation we use to determine the amount of energy needed to heat a material is fairly intuitive once students understand heat capacity. The mass term (m) indicates that if you have more of a substance to heat, it will take more energy – a simple concept. The C_p term indicates that every substance requires a different amount of energy to heat. The ΔT term indicates that the amount of energy you add will depend on how hot you want to make it. All of these are easily understood, and students shouldn't have problems solving this equation.

Though we won't be using this in the lab, it's also important to know how much energy it takes to change the state of a substance. The equation used is $\Delta H = n\Delta H_x$, where "n" stands for the number of moles of the substance and the "x" subscript will determine whether the heat of vaporization (ΔH_{vap}, corresponding to boiling/condensing) or the heat of fusion (ΔH_{fus}, corresponding to melting/freezing) will be used. The idea here is roughly the same as the first equation: The amount of energy needed to change state will depend on how much of the substance is present (the "n" term) and the amount of energy it takes for one mole of a substance to undergo a change in state (the ΔH_x term).

Once students understand both equations, they can be put together to determine how much energy will be required to change both the temperature and the state of a substance. The best way to teach students how to do this is to treat the problem as a series of smaller steps corresponding either to temperature or state change – the answers to these problems can then be put together to get the final answer. A word of caution: Because the units of ΔH_x are measured in kJ / mol and the units of C_p are measured in J / g $\cdot$ ^{0}C, the answers will be in different units. Make sure students change either joules to kilojoules (or vice-versa) before adding the values together.

Doing the Heat Capacity of Water Lab

Equipment:

- 6-250 mL beakers
- 6 ring stands with rings
- 6 Bunsen burners
- 2 balances
- 6 thermometers
- 2 flint strikers
- 6 stopwatches or a wall clock with a second hand
- 6 beaker tongs

Safety:

Students will be working with Bunsen burners, so all of the usual precautions for working with Bunsen burners should be followed. As always, students should wear goggles during this lab.

Room destruction factor:

If fire safety procedures are followed, there is no danger of damage to the lab.

How the lab works:

In this lab, students heat room temperature water until it begins boiling. Using my estimate of how much energy a Bunsen burner produces during one second of operation, the mass of the water, and the change in termperature of the water, students can work backwards to find the heat capacity of water.

Students need to measure the temperature of the water before they start the heating process. Students sometimes forget to take an initial water temperature, making it impossible to determine the overall change in temperature.

This lab is interesting in that it's not shy about having several sources of experimental error. The estimated heat output value for a Bunsen burner is 1800 J/sec – clearly, this won't be the same for all Bunsen burners in all locations. This lab also pretends that all of the heat from the Bunsen burner goes straight into the water – it should be obvious that a large quantity of heat won't be absorbed by the water at all. This is a problem that most heat calculation labs have (notably the classic one where you find the heat of combustion of a nut by burning it under a soda can full of water). No matter how you set up this sort of lab, most of the heat won't be absorbed by the water at all. To get a truly good answer, you'd need to burn the nut in a bomb calorimeter.

This lab at least has the sense to ask students, "How would you change this experiment to increase the accuracy of your data?" You'll find that your students have numerous changes they'd like to make to this lab. After all, they're going to find that the observed heat capacity of water is many times the expected value of 4.184 J / g · ^{0}C. The most likely improvement that students will have for the lab is that the heat from the Bunsen burner will be enclosed and aimed at the can to avoid radiative or convective heat loss to the rest of the room. For advanced classes it may be interesting to have your students perform their revised experiments and compare the data from the two experiments.

<u>What can go wrong:</u>

Because the lab gives an inherently incorrect answer, it's difficult to really identify anything that can go wrong. Typically, students struggle with the inaccuracy of their data, but eventually come to realize that it's not due to anything they've done, but rather to the flawed experimental procedure they've been given.

<u>Clean up:</u>

The hot water can go down the sink.

Solutions for the Heat Capacity of Water Lab

<u>Prelab:</u> C_p = (32,000 J)/(100 g x 70^0 C) = 4.57 J / g · ^{0}C.

<u>Lab:</u>

Step 1: The mass of the water should be approximately 175 grams.
Step 3: The temperature of the water will be between 15 – 25^0 C.
Steps 4 and 5: The exact time (to the second) should be given.

<u>Calculations:</u>

1) The amount of energy is equal to the time (in seconds) times 1800 J. A typical value is 600 – 900 kJ.
2) The mass of the water is the same as in step 1 of the lab.
3) The change in temperature will be approximately 75^0 C, depending on the initial water temperature from step 1.
4) Typical values range between 35 – 65 J / g $\cdot$ ^{0}C.

Post-lab questions:

1) The main source of experimental error is that most of the heat generated by the Bunsen burner isn't absorbed by the water at all. Our calculations, on the other hand, make the erroneous assumption that all of the heat generated is absorbed. This causes a roughly tenfold error.
2) One way of improving the data is to thermally isolate the Bunsen burner such that the only place the heat could travel would be the water. Specific suggestions, such as placing a column around the Bunsen burner or using a wider beaker should be included.
3) Assuming exactly 175 grams of water and a temperature change of 75^0 C, the amount of energy would be 54,900 J.

Solutions for the Heat Calculations Worksheet

1) 5.00 kJ
2) 198.1 kJ
3) -1726 kJ

Heat Calculation Handout

In chemistry, "heat" and "temperature" are different. "Heat" is the amount of energy a chemical has, frequently measured in joules (J). Because we can't directly measure heat, we have to measure "temperature", which reflects how much kinetic energy an object has (as measured in degrees Celsius or Kelvins). In thermodymanics, the term "enthalpy" is used interchangably with "heat", as it avoids confusion between the terms "heat" and "temperature".

It takes energy to make the temperature of anything increse. The relationship between energy and temperature is shown by the equation:

$$\mathbf{\Delta H = mC_p\Delta T}$$

ΔH = change in energy (J)
m = mass of the material (g)
C_p = specific heat capacity (J / g $\cdot$ ^{0}C)
ΔT = change in temperature (^{0}C or K)

C_p has a different value for every state of every substance. For example, the heat capacity of liquid water is 4.184 J / g $\cdot$ ^{0}C, while that of steam is 1.87 J / g $\cdot$ ^{0}C.

It also takes energy to melt or boil any substance. The amount of energy required to melt or boil a substance can be expressed by the following equations:

$$\mathbf{\Delta H = n\Delta H_{fusion}}$$

$$\mathbf{\Delta H = n\Delta H_{vaporization}}$$

ΔH = change in energy (J)
n = number of moles
ΔH_{fusion} = the molar heat of fusion (kJ/mol)
$\Delta H_{vaporization}$ = the molar heat of vaporization (kJ/mol)

ΔH_{fusion} and $\Delta H_{vaporization}$ are constants and correspond to the amount of energy it takes to freeze (fuse) or boil (vaporize) one mole of a substance.

When doing heat calculations that involve both a change of state and a change in temperature, make sure the answers for both calculations are written in the same units before adding them together!

Finding the Heat Capacity of Water

In this lab, we're going to find the heat capacity of water using the equation $\Delta H = mC_p\Delta T$. Usually, we solve for the change in energy if we're given the mass of water, the heat capacity of water, and the change of temperature. In our lab today, we're going to use experimental values of ΔH, m, and ΔT to find the heat capacity of water.

Prelab:

1) If you've found that the temperature of 100 grams of water has been raised by 70^0 C and it took 32,000 Joules of energy to cause this temperature change, what is the heat capacity of water?

Lab:

1) Using a balance, place approximately 175 grams of water into a 250 mL beaker.

   The exact mass of water in the beaker is: ___________________

2) Set up a ring stand so that a ring with wire gauze is approximately 30 centimeters from the bottom of the stand. Place the beaker full of water on top of the wire gauze.

3) Find the initial temperature of the water in degrees Celsius: ____________

4) Place a Bunsen burner below the ring and prepare to light it with a flint striker. Write the time (to the nearest second) that you light the burner.

   The exact time we started this experiment was: ___________________

5) Heat the water until it starts to boil. Boiling is defined as the time that the water starts bubbling violently – if small bubbles form and slowly rise to the surface, it has not started boiling. Write the time (to the nearest second) that the water starts boiling.

   The exact time the water started boiling was: ___________________

6) Once you have collected data in step 5, you may stop the experiment and pour the boiling water down the sink.

Calculations:

1) **Finding DH**: The gas used to heat the water was methane. For every second that the Bunsen burner was running, approximately 1,800 Joules of energy are given off. Using this information and your data from steps 4 and 5, calculate how much energy was added to the water during this experiment.

DH for this experiment was: ____________________

2) **Finding m**: From your experimental data, what was the mass of water you used?

m for this experiment was: ____________________

3) **Finding DT**: Given the experimental data from step 3 and the fact that water boils at 100^0 C, calculate the change in temperature for water:

DT for this experiment was: ____________________

4) Using the equation on the first page of this lab and the answers from calculations 1-3, find your experimental value for the heat capacity of water:

We found the heat capacity of water to be: ____________________

Post lab questions:

1) The actual value for the heat capacity of water is 4.184 J / g · ^{0}C. Explain why your answer was different, citing possible sources of experimental error. This will count for a large portion of your lab grade, so be complete!

2) If you had to do this lab again, how would you change the experiment to increase the accuracy of your data? Your answer should be complete and precise.

3) How much energy should it have taken for the water to boil? In other words, use the true value of C_p, 4.184 J / g · ^{0}C, to calculate the amount of energy it should have taken for the water to start boiling.

Heat Calculations Worksheet

Use the following table of constants to solve the following problems:

compound	ΔH_{fusion}	melting point	C_p(liquid)	$\Delta H_{vaporization}$	boiling point	C_p(gas)
water	6.0 kJ/mol	0^0 C	4.184 J / g · ^{0}C	40.6 kJ/mol	100^0 C	1.87 J / g · ^{0}C
benzene	10.6 kJ/mol	6^0 C	1.7 J / g · ^{0}C	30.6 kJ/mol	80^0 C	1.05 J / g · ^{0}C

1) When 15 grams of ice at a temperature of 0^0 C has enough energy added to make liquid water at a temperature of 0^0 C, what is the change in enthalpy?

2) When 65.8 grams of ice at 0^0 C is heated to a temperature of 110^0 C, what is the change in enthalpy?

3) If we have 450 grams of water vapor and 95 grams of benzene vapor at a temperature of 845^0 C, what will the enthalpy change be if both are cooled to a temperature of 95^0 C?

Chapter 15 – Calorimeter Lab

Overview

Calorimeters are devices used to find heats of reaction. Chemical reactions are performed in an enclosed space thermally-isolated from everything but a nearby reservoir of water. The energy given off by the reaction causes the water temperature to increase. By using $\Delta H = mC_p\Delta T$, one can determine the amount of energy released by the process. In this lab, students will use calorimeters of their own design to find the molar heat of solvation of sodium hydroxide.

Teaching about calorimetry

Before starting the calorimeter lab, it's important that students have a good understanding of heat calculations. If your students can't find the heat of reaction using data you've given them, they're going to have an even harder time trying to make sense of experimental data.

Bomb calorimeters are the most common type used in beginning chemistry courses. In a bomb calorimeter, organic substances are placed into a metal capsule called a "bomb". The bomb is inserted into a tank of water that's outfitted with a stirrer and thermometer. Oxygen is added to the bomb until the pressure is much higher than atmospheric pressure and a spark is applied to the organic compound, causing it to burn. This, in turn, causes the temperature of the bomb, then the temperature of the water to increase. The increase in water temperature is related to the amount of energy given off by $\Delta H = mC_p\Delta T$.

The calculations needed to interpret calorimeter data are the same calculations needed to determine how much energy it takes to raise the temperature of a compound – after all, we are heating a compound (water) and it's easy to measure both the change in temperature and the mass. Students like this lab because they finally get to see what heat calculations are used for in the lab.

This lab doesn't require a bomb calorimeter, as students will build their own calorimeters from inexpensive and easily-obtained materials. Before performing this lab, you should spend some time in class discussing why it is important to make sure that the calorimeter is thermally-isolated from the environment.

Doing the Calorimeter Lab

Equipment:

- 72 Styrofoam cups (8 oz. or larger)
- 72 plastic cups (8 oz. or larger)
- 72 paper cups (8 oz. or larger)
- 6 – 250 mL beakers
- 6 watch glasses
- 24 paper plates
- 3 rolls of paper towels
- 3 rolls of aluminum foil
- 3 rolls of plastic sandwich wrap
- 3 rolls of clear plastic tape
- 1 box of paper clips
- 1 box of straws
- 6 scissors
- 6 thermometers
- 3 – 100 mL graduated cylinders
- 2 balances
- 2 spatulas
- 100 grams of sodium hydroxide pellets
- 6 stopwatches, or a wall clock with a second hand

Safety:

Students should wear goggles during this lab because sodium hydroxide is extremely corrosive to skin. If students notice that their hands feel slippery, have them dip their hands in household vinegar to neutralize the sodium hydroxide.

Room destruction factor:

When students build their calorimeters they sometimes have a tendency to leave small pieces of paper, styrofoam, plastic, etc. lying around the room. Remind students not to place trash in the sinks, as both aluminum foil and plastic sandwich wrap are excellent for plugging drains.

How the lab works:

In this lab, we move away from the pure chemistry of heat calculations and into the realm of chemical engineering. The goal is to build a calorimeter that can accurately measure changes in temperature due to a chemical reaction, using only random pieces of plastic tableware. We can all agree that the reaction is chemistry. Calorimeter design, on the other hand, is engineering.

As may be expected, the average high school student has limited experience with chemical engineering. As a result, give your students some suggestions to help them with their construction:

1) Good calorimeters are well insulated. The water inside the calorimeter needs to be as isolated from the rest of the room as possible to ensure that any chemical changes are due only to the chemical reaction being studied. A good rule of thumb is that whenever you want to insulate something from its surroundings, treat it as if it's hot coffee that you want to keep warm for a long time. Another thing students should keep in mind is that heat doesn't just travel through the sides of a cup – it also travels through the top and bottom, making it necessary to insulate those as well.

2) Good calorimeters are easy to use. If students put together a calorimeter with perfect insulation, they may find that it's impossible to measure the temperature change.

3) Their calorimeter should be large enough to contain the reaction mixture. If your students make a calorimeter that can hold only 50 mL of water, it will be insufficient for holding the required 200 mL of water in this lab.

Aside from providing this information, your role in this lab is to provide supplies for your students and ensure that nobody gets hurt. Students are creative when building their calorimeters, and some of the resulting calorimeters will give excellent experimental results.

The results for this lab are generally fairly accurate. The actual molar heat of solvation of sodium hydroxide is -44.5 kJ/mol. Typical student answers are between -35 kJ/mol and -45 kJ/mol.

What can go wrong:

- The calorimeter your students make is too difficult to use. Generally, this occurs when students make calorimeters with a large number of parts. Suggest to your students that simpler items work better than complex ones.
- Your students don't understand what they're graphing. The reason they're making a graph is so they can figure out the change in temperature for the reaction. When your students graph temperature versus time, they should find the temperature to be stable before and after the reaction. The difference between these two temperatures is the ΔT value they're trying to find.
- The calorimeter dissolves! This rarely happens, but may occur if paper cups are exclusively used. Remind your students that choosing quality materials is an important part of engineering and have them try again.

Clean up:

This lab generates basic waste. To dispose of it, add phenolphthalein to the waste and slowly add acetic acid until the color changes from pink to clear. When the color has changed, add sodium bicarbonate until it stops bubbling, then dump it down the sink.

Solutions for the Calorimeter Lab

Prelab:

1) 8368 J
2) If the reaction wasn't totally surrounded by water, some of the heat would escape the calorimeter, resulting in faulty data.

Section 1:

The students should make a sketch of their calorimeters, showing how all of the materials fit together. A common mistake in this section is to leave out the illustration of the thermometer.

Section 2:

Data table: Should be a table showing the relationship of temperature to time.
Graph: Should be a graph of the data in the table, including units, a title, and shown as a curve (not a connect-the-dots).
Heat of reaction: The students should use their data to find the heat of solvation. If 200 mL of solution is used, the heat of solvation should be equal to 836.8 times the change in temperature. An excellent value is -4,440 Joules – typical student answers are somewhat smaller.

Postlab questions:

1) The molar heat of solvation should be ten times the heat of solvation in Section 2, because 0.1 moles of NaOH were used. The correct answer is -44.5 kJ/mol, though this will likely be somewhat smaller.
2) Students should give a good reason why the molar heat of solvation is incorrect, based on their data and observations.
3) *Specific* suggestions should be provided for improving the calorimeter.
4) It would not be accurate, because some of the heat would be transferred to the gas and escape.

Solutions to the Calorimeter Worksheet

1) A chemical change takes place in a thermally-isolated container. The heat of reaction is transferred to water and can be measured by $\Delta H = mC_p\Delta T$.
2) This will cause the heat of reaction to be larger than the real value because the environment would be adding heat to the calorimeter.

3) “Heat of reaction” refers to the amount of heat given off in a specific reaction. “Molar heat of reaction” refers to the amount of heat given off when one mole of a substance reacts. Example: If 0.5 moles of a compound gives off 100 J when burned, the heat of combustion is 100 J and the molar heat of combustion is 200 J.
4) 375 x 4.184 x 18 = 28,240 J
5) 28,240 x 5 = 141,200 J

Calorimeter Lab

In this lab, we'll be measuring the energy change when a substance is dissolved in water. This energy change is referred to as the heat of solvation, or $\Delta H_{solvation}$. Because you will know the mass of water used, the specific heat capacity of water, and the change in temperature during this process, you can solve the equation $\Delta H = mC_p\Delta T$ to find the heat of solvation.

Prelab: Practice calorimeter problems

1) I have built a calorimeter that holds 100 grams of water. If the temperature of the water rises by 20^0 C during a reaction and the heat capacity for water is 4.184 J / gram $\cdot$ ^{0}C, what is the heat of reaction?

2) Why do we need to make sure that the reaction is completely surrounded by water? What would happen if it weren't?

Section 1: Building a calorimeter

In this section you'll design and build a calorimeter of your own. Your teacher will provide you with a list of meterials you may use in building your calorimeter. Keep the following tips in mind when designing your calorimeter:

- The water should be *thermally-isolated* from the rest of the room. *Thermal isolation* means that no energy from the water can enter the room, or vice-versa. Any material that keeps heat from traveling from one place to another is called *insulation*.
- You will need to place your thermometer in a location where you can read it while the reaction is taking place.
- You will need to build your calorimeter so it is easy to add water and sodium hydroxide pellets.
- You will need to hold 200 mL of water to contain the reaction.

Section 1 (continued):

Sketch your calorimeter below, making sure you label each part clearly:

Section 2: Finding DH for the solvation of sodium hydroxide

In this section, you will measure the molar heat of solvation for sodium hydroxide. To do this, you will use your calorimeter. The general procedure for this lab will be the same for everyone – however, you may need to make some slight changes depending on the design of your particular calorimeter.

1) Pour 200 mL of distilled water into a the reaction area of your calorimeter.

2) Make a data table (next page) into which you will put temperature measurements of the solution every 20 seconds.

3) Place the thermometer into the water and measure the temperature. This first measurement of temperature is your T=0 measurement.

4) Add 4 grams of NaOH to the water in your calorimeter and start the timer.

5) Take a measurement of the temperature in your calorimeter every 20 seconds and write them in your data table. When the temperature stops rising, take another 3 temperature measurements, then stop.

6) Using the data from your table, make a graph of temperature versus time. Space for this is provided to you on the next page.

7) The final temperature you found minus your T=0 temperature is equal to the temperature change of the water, DT.

8) Calculate the amount of energy released in this reaction using the equation on the first page of this lab.

Section 2 (continued):

<u>Write your data table in the area below. Make sure you use a ruler!</u>

<u>Graph (Temperature vs. Time):</u>

<u>Find the heat of solvation below. Show all of your calculations:</u>

Postlab questions:

1) Using the heat of solvation you found on the last page, find the *molar* heat of solvation of sodium hydroxide. Molar heat of solvation is defined as the amount of energy given off when one mole of a substance is dissolved.

2) Do you think that the value you got for the molar heat of solvation of sodium hydroxide is correct? If so, explain why. If not, explain why not.

3) If you could rebuild your calorimeter, how would you improve it? Explain the changes you'd make and why you think they'd make your answer more accurate.

4) Do you think your calorimeter would be accurate for measuring the heat of reaction for a chemical reaction that forms a gas? Why or why not?

Calorimeter Worksheet

1) Briefly describe how a calorimeter works.

2) If you were to do this experiment on a hot summer day in a room without air conditioning, how might this change your results? Explain.

3) Explain how "heat of reaction" and "molar heat of reaction" are different.

4) I did an experiment where I placed 375 grams of water into a calorimeter. When the reaction finished, the water temperature had increased by 18^0 C. What was the heat of reaction for this process? (The heat capacity of water is 4.184 J / g $\cdot$ ^{0}C).

5) The reaction in problem 4 involved the reaction of 0.20 moles of an unknown compound. Using the your answer from problem 4, what was the molar heat of reaction of the unknown compound?

Chapter 16 – Energy Diagram Lab

Overview

Energy diagrams show what happens energetically during the course of a chemical reaction. By showing the energy of the reagents, the transition state, and the products in a visual format, students gain a better appreciation for how energy relates to chemical processes. In this lab, students will apply what they've learned about energy diagrams to the synthesis of magnesium oxide from magnesium and oxygen.

Teaching about energy diagrams

Energy diagrams are a good way to introduce your students to thermodynamics. The very term "thermodynamics" is enough to convince students that what they're studying is too difficult to understand – fortunately, energy diagrams are simple enough to understand that your students should gain some confidence before moving on to more complex calculations.

The best way to teach your students about thermodynamics is through analogy. The analogy I use is that of driving a car over a large mountain. In this analogy, the car represents the evolution of the reagents over time, the starting point represents the reagents, the ending point represents the products, and the transition state represents the top of the mountain.

- High mountains will take longer to climb than small ones. This is related to energy diagrams in that reactions with high energy transition states generally proceed more slowly than those with low energy transition states.
- Cars with small motors won't be able to climb the hill at all. In the same way, reagents that combine with low energy will be unreactive because they can't reach the transition state in the first place.
- The presence of a tunnel allows your trip over the mountain to proceed much more quickly. Clearly, this corresponds to the presence of a catalyst in a chemical reaction. When a car travels through a tunnel to get through a mountain, the final destination will be the same as if it traveled over the top of the mountain, but the pathway it used to get there is much different. In the same way, the catalyzed transition state is much different than the uncatalyzed transition state, but the reaction eventually forms the same products.
- If the destination is a long way down the mountain from the top, the car will be moving much more quickly than if it was near the top. In the same way, if the products have a lower energy than the reagents, the reaction will be exothermic (corresponding to a high speed in a car). If the products have a higher energy than the reagents, the reaction will be endothermic

(corresponding to a much lower velocity). This even helps students to remember that exothermic reactions have a negative ΔH_{rxn} and endothermic reactions have a positive ΔH_{rxn}. If the car ends up at a lower altitude than it started, the the altitude is less than it was. Likewise, an exothermic reaction, in which the products have lower energy than the reagents has a ΔH_{rx} with a negative sign. OK – I'll admit the analogy is a little strained at this point, but students seem to buy into it.

Doing the Energy Diagram Lab

Equipment:

- 0.3 meter Mg ribbon
- 2 balances
- 6 ring stands with rings
- 6 clay triangles
- 6 crucibles (lids are not required)
- 6 crucible tongs
- 6 Bunsen burners
- 2 flint strikers

Safety:

The usual precautions should be used when working with Bunsen burners. Additionally, students should be reminded that crucibles should be cooled on the countertop and not with running water, as the water will cause them to shatter.

You should be very careful to ensure that each student is given only a small amount of magnesium ribbon, as large magnesium fires are impossible to control. In the very unlikely event that a magnesium fire gets out of control, do NOT use water to extinguish it. Instead, evacuate the school and call the fire department.

Room destruction factor:

If the usual fire safety rules are followed, there should be no problems.

How the lab works:

In this lab, students will be heating magnesium ribbon in a crucible until it ignites, signaling that the reaction of magnesium with oxygen forms magnesium oxide in an exothermic reaction. The magnesium ribbon will burn for approximately five seconds before going out, at which point the experiment is over.

There's more going on with this lab than meets the eye:

- The difficulty that you have in getting a reaction to start gives you information about the activation energy. For example, the ignition of a Bunsen burner requires little effort, suggesting that the activation energy for the combustion of methane is low. On the other hand, the decomposition of calcium carbonate into calcium oxide and carbon dioxide requires very high temperatures, indicative of a large activation energy. In this lab, it takes approximately 60 seconds for the magnesium to ignite, suggesting a high activation energy.
- The formation of magnesium oxide is exothermic NOT because the magnesium ribbon gets hot when heated over the Bunsen burner, but because the amount of energy released when it burns is greater than the amount of energy required to start the reaction in the first place. Similarly, students sometimes mistakenly believe that heating water until it boils is an exothermic reaction because the water becomes hot before boiling. However, the amount of energy the water possesses after it boils is identical to the amount of energy added to it, indicating an adiabatic process (which isn't a reaction at all!)
- It's possible to determine the magnitude of DH_{rxn} by observing the amount of energy released when the magnesium reacts with oxygen. Because a very large amount of energy is given off, it can be surmised that the reaction is extremely exothermic. (The actual value is -601.7 kJ/mol, confirming this assumption.)

What can go wrong:

Once in a while, the magnesium won't ignite. This typically occurs when the Bunsen burners aren't set to burn at a high temperature. By adjusting the oxygen intake at the bottom of the Bunsen burner, the flame can be made hot enough to ignite the magnesium.

Clean up:

Magnesium oxide and may be safely thrown away.

Solutions for the Energy Diagram Lab

Prelab:

1) -445 kJ (½ times -890 kJ/mol)
2) exothermic (you can tell by the negative sign)
3) 1.12 moles (-1000 divided by -890)

Experiment section:

1) The mass of the magnesium ribbon should be written in grams.
2) The magnesium will burn with a brilliant white light, giving off white smoke and ultimately forming a white powder.

Postlab questions:

1) $2\ Mg_{(s)} + O_{2(g)} \rightarrow 2\ MgO_{(s)}$
2) Exothermic – a great deal of heat and light were given off.
3) We needed to heat the reaction to a high temperature because it has a very high activation energy.
4) The amount of energy given off should be equal to:

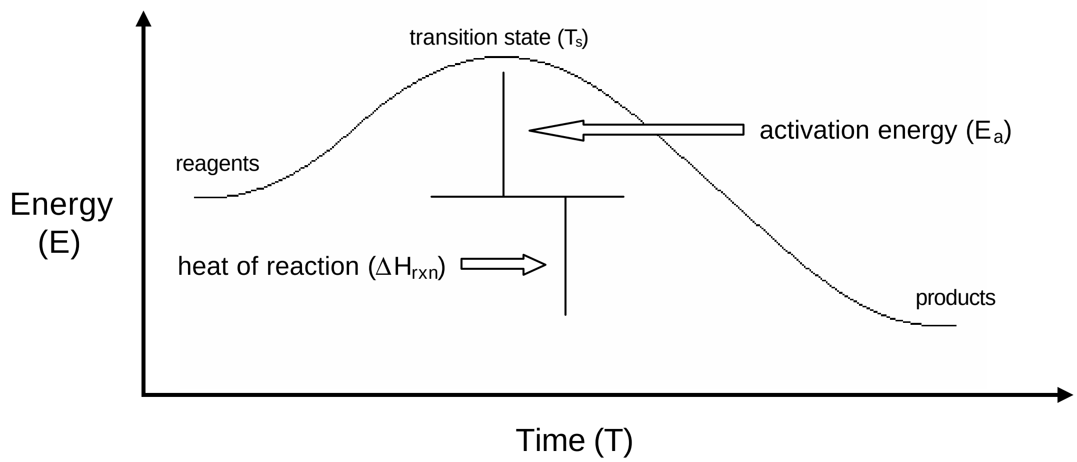
   (mass of ribbon/24.3) x -601.7 kJ
5) The energy diagram should show a highly exothermic reaction with a high activation energy.

Solutions for the Energy Diagram Worksheet

1) Endothermic, because the products have higher energy than the reagents.
2) It's a slow reaction because it has a high activation energy.
3) You could make the reaction proceed more quickly by adding a catalyst to decrease the activation energy. The reaction will also go more quickly if the reagents are heated.
4) Transition state refers to how the reagents are arranged with respect to one another. Activation energy refers to the energy required to orient the reagents in this manner.
5) They are spontaneous because they proceed to products without outside intervention. (Keep in mind that "spontaneous" doesn't necessarily mean "fast". For example, the combustion of gasoline is spontaneous because it will eventually form CO_2 and H_2O on its own, though not anywhere near as quickly as if there were an energy source present).
6) An energy diagram should be written that shows an exothermic reaction with a ΔH_{rxn} of -484 kJ/mol. The activation energy was not specified and cannot be determined for this process.
7) Hydrochloric acid is not a catalyst because it is consumed in this reaction and causes the formation of different products than if it were not present. Catalysts are never consumed in chemical reactions, and the overall equation for a catalyzed reaction is the same as that for an uncatalyzed reaction.

Energy Diagram Handout

Energy diagrams are illustrations that show the relationships between time and energy during the course of a chemical reaction. As you know, the reagents in a chemical reaction have some level of energy that changes as the reaction proceeds to the transition state, then finally the products. An energy diagram for an exothermic reaction is shown below:



The following are some important terms you need to remember when making energy diagrams:

- **reagents:** These are the chemicals that you start with in a reaction.
- **products:** These are the chemicals formed during the reaction.
- **heat of reaction (DH_{rxn}):** The difference in energy between reagents and products. The units of ΔH_{rxn} are Joules.
- **exothermic:** Exothermic reactions give off heat because the products have less energy than the reagents. The sign of ΔH_{rxn} is negative.
- **endothermic:** Endothermic reactions absorb heat (and feel cold) because the products have more energy than the reagents. The sign of DH_{rxn} is positive.
- **transition state (T_s):** The highest point on the energy diagram, representing the point at which the reaction is half-completed.
- **activation energy (E_a):** The amount of energy required for the reaction to take place. The higher the activation energy, the slower the reaction. The units of E_a are Joules.
- **catalyst:** Any chemical that causes the activation energy of a reaction to be lowered, speeding the rate of the reaction. Catalysts are not consumed during chemical reactions, meaning that one molecule of a catalyst can cause many molecules of reagent to react more quickly. If the activation energy is thought of as a hill, catalysts can be thought of as machines that dig tunnels to make it easier to move from one place to another.

Energy Diagram Lab

In class, we've talked about how energy diagrams show how the energies of the reagents in a chemical reaction change over time. You should by now be familiar with the ideas of endothermic reactions, exothermic reactions, activation energies, heats of reaction, and other thermodynamic terms.

In this lab, we're going to draw the energy diagram for the reaction of magnesium with oxygen to form magnesium oxide.

Prelab:

The combustion of methane can be described by the following equation:

$$CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O \qquad \Delta H = -890\ \text{kJ/mol}$$

1) How much energy will be given off when 0.5 moles of methane burn?

2) Is this reaction exothermic or endothermic? ________________________

3) How many moles of CH_4 would be needed to give off 1000 kJ of energy?

Experiment:

1) Find the mass of a 4 cm piece of magnesium ribbon: _________________

2) Place the magnesium ribbon in a crucible and place the crucible in a clay triangle over a Bunsen burner. Turn on the Bunsen burner and heat the ribbon until a chemical reaction takes place. Write your observations of the reaction below:

3) Clean up: The magnesium oxide powder can be thrown away in the trash.

Postlab questions:

1) Write the complete, blanced equation for the reaction of magnesium with oxygen to form magnesium oxide.

2) Was this reaction exothermic or endothermic? Explain your answer.

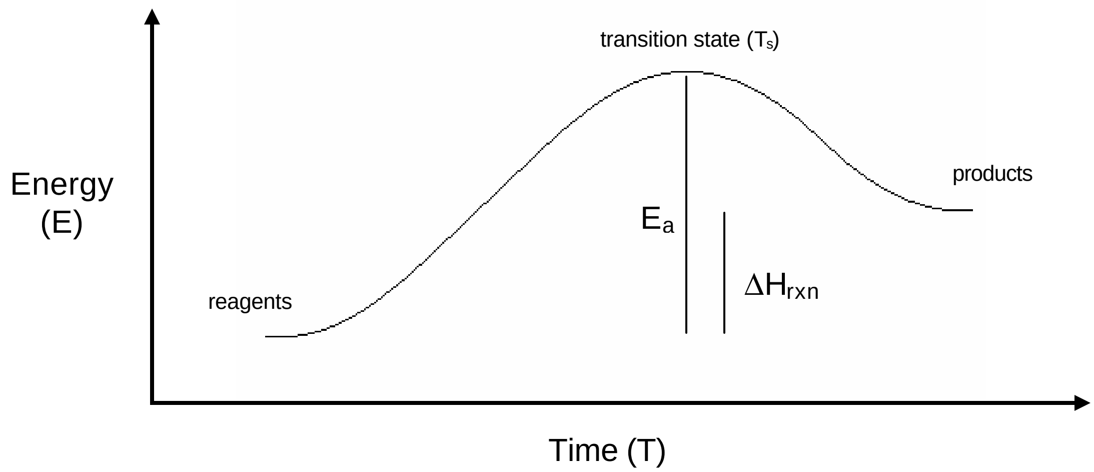
3) Explain why we had to heat the magnesium to such a high temperature before the reaction would take place.

4) ΔH_{rxn} for the process you observed in this reaction is -601.7 kJ/mol. Using this information and the mass of the magnesium you found in step 1 of the lab, find the amount of energy given off when you burned your piece of magnesium ribbon.

5) Sketch the energy diagram for this reaction, labeling all parts, including ΔH_{rxn}, E_a, products, reagents, and the transition state.

Energy Diagram Worksheet

Use the following energy diagram to answer questions 1-3:



1) Is this reaction exothermic or endothermic? How can you tell?

2) Based on the activation energy, is this reaction fast or slow? Explain.

3) How could you make this reaction proceed more quickly? Explain, adding to the illustration above, if needed.

4) Explain the difference between transition state and activation energy.

5) Explain why exothermic reactions are referred to as "spontaneous".

6) Consider the following reaction:

$$2\ H_{2(g)} + O_{2(g)} \rightarrow 2\ H_2O_{(g)} \qquad \Delta H_{rxn} = -484\ kJ/mol$$

Using this information, draw a complete energy diagram for this reaction, labeling all important parts:

7) The addition of hydrochloric acid to water makes iron corrode more quickly than water alone. The equation for this reaction is shown below:

$$Fe_{(s)} + 2\ HCl_{(aq)} \rightarrow FeCl_{2(aq)} + H_{2(g)}$$

In this reaction, is hydrochloric acid a catalyst? Explain why or why not.

Chapter 17 – Combined Gas Law Lab

Overview

The combined gas law describes the relationship between the temperature, pressure, and volume of a gas. It's called the combined gas law because it's a combination of Boyle's law and Charles' law, both of which were earlier attempts to explain how gases behave under changing conditions. In this lab, students will heat and cool a gas to determine the relationship between pressure, temperature, and volume. This will be accomplished by heating and cooling sealed test tubes and observing what happens as the temperature changes.

Teaching about the combined gas law

The combined gas law is typically taught after Boyle's law and Charles' law and explained as a method of explaining the interrelationship between the pressure, volume, and temperature of a gas. When your students learn the ideal gas law, they'll know just about everything they need to know about the relationships between pressure, volume, and temperature.

I've found, however, that students frequently have a hard time telling when to use each of the four above gas laws. The following are some guidelines that may help your students to use the correct gas law at the correct time:

- If "moles" or "grams" are mentioned in the problem, use the ideal gas law. The other laws don't use either unit at all.
- For all other problems, use the combined gas law. NEVER use Boyle's or Charles' laws!

The above guidelines may imply to you that Boyle's and Charles' laws are unimportant, which clearly is not the case – after all, they were the historical basis for the combined gas law and should be taught for this reason alone. However, students frequently confuse these equations, which causes problems when solving gas law problems. To cure this problem, students should be told to only use the combined gas law. An added benefit to students is that they only need to memorize two equations (the combined and ideal gas laws), much easier than memorizing four.

In order for this strategy to work, students need to know how to properly use the combined gas law. Simply put, if one of the variables doesn't change during some process (or isn't mentioned at all), just eliminate it when using the combined gas law. Thus, for constant temperature processes, the combined gas law is simplified to make Boyle's law.

When students finally understand which gas law should be used for each type of problem, many still make common errors that lead to incorrect answers. These problems include the following:

- Students solve problems using degrees Celsius instead of Kelvins. This may have the unfortunate effect of causing gases to have negative volumes and and pressures. Explain to students that the kinetic molecular theory of gases states that the kinetic energy of a gas is proportional to its temperature and that using the Celsius scale would cause the kinetic energy of a gas to be negative, which is clearly impossible.
- Students use the wrong units of pressure. For the combined gas law, the units of pressure don't really matter as long as they are the same for both sides of the equation. For the ideal gas law, either kilopascals or atmospheres may be used as long as the corresponding value for the ideal gas constant is also used.
- Students use the wrong units of volume. Again, this doesn't matter for the combined gas law. For the ideal gas law, volume should always be expressed in liters.

Doing the Combined Gas Law Lab

Equipment:

- 25 small test tubes (about the size of your index finger)
- 25 rubber stoppers that fit the test tubes.
- 6-10 mL graduated cylinders
- 15 rubber gloves
- 30 rubber bands
- 6 test tube tongs
- 6 Bunsen burners
- 2 flint strikers
- 5 lbs. crushed ice (not required, but helpful)

Safety:

As with any lab with a Bunsen burner, students should be cautioned that hot lab equipment remains hot for a very long time after the flame is extinguished. Goggles should be worn in this lab and students should NOT put on the rubber gloves at any time!

Room destruction factor:

Broken test tubes are about the worst that can happen with this lab – even this can be minimized by using test tubes made of heat-resistant glass.

How the lab works:

In the first part of the lab, students will stretch a piece of rubber glove over a test tube, secure it with rubber bands, and heat the closed end of the test tube. This causes the gas in the test tube to expand, eventually causing the latex to break with a loud popping noise. In the second part of the lab, student will heat a test tube over the Bunsen burner and seal it with a rubber stopper. As the gas inside the tube cools, the pressure inside the test tube decreases until the stopper is firmly stuck in the opening.

Because each section is essentially a different experiment, I'll discuss how each works separately:

Experiment one:

In step one, students find the volume of the test tube by filling it to the top with a graduated cylinder. The volume of the test tube will be equal to the volume of water required to fill it to the top. Some students may not understand how this works, so be prepared to explain it. *When your students pour the water out of the test tube, tell them **not** to dry the last few drops remaining in the test tube.* We'll discuss why in a moment.

In step three, students will stretch the rubber glove over the mouth of the test tube. Only a very small amount of rubber glove is required for this step, not the whole glove! Make sure the rubber is stretched very tightly over the mouth of the test tube and fastened with two rubber bands to ensure it won't pop off during heating.

Back to step one, and the reason we don't want students to dry the last few drops from the test tube. During this experiment, students will be heating a test tube over the Bunsen burner to observe the expansion of the gas inside. Unfortunately, there isn't much gas inside a test tube – certainly not enough to cause the rubber glove to burst. The vaporization of these last few drops of water ensures that the volume of the gas in the tube will increase enough for the rubber glove to burst. The same thing will happen if the water is completely removed, but the lab is nowhere near as dramatic.

Experiment two:

Before starting this section of the lab, make sure the test tube has cooled before attempting to remove the rubber glove from step one. Remind your students that hot glassware looks the same as cold glassware, making it difficult to tell when it has completely cooled.

In step three, it's important that your students moisten the rubber stopper to ensure that the seal between the stopper and the test tube is complete. If the stopper is not moistened, air can easily get into the test tube, causing this lab to fail. Additionally, students should be extremely careful when placing the stopper in the test tube, as the test tube will be extremely hot after 60 seconds of heating. For safety, students should hold the closed end of the test tube against the lab bench with test tube tongs before inserting the stopper to ensure that the hot test tube doesn't fall.

In step four, students should wait a few minutes before placing the hot test tube in the cold water bath to keep the glass from shattering. Incidentially, it doesn't matter if you use a cold water bath or an ice water bath because the temperature difference between the two is only about fifteen degrees Celsius – insignificant considering the very high initial temperature of the test tube. The cold water isn't necessary to make the lab work, but is recommended because it helps students to intuitively understand that the test tube was drastically cooled.

What can go wrong (in either part of this lab):

- The test tube shatters (either part of the lab). Have the students get another test tube and try again.
- The rubber glove melts to the test tube (experiment one). This is caused when students heat too close to the mouth of the test tube. The melted rubber is easy to remove, but students will have to redo this part of the lab.
- The stopper is sucked completely into the test tube (experiment two). This occurs when the stopper is too small for the test tube. Unfortunately, there's not much you can do about this but to shatter the tube to retrieve your stopper. Under no circumstances should you heat the tube to blow the stopper out, as it may explode.
- Students become agitated because they don't know the volume of the test tube in experiment two. It should be pointed out to them that even though they didn't find a volume in experiment two, they found the volume of the tube in experiment one – most likely it hasn't changed much.

Clean up:

Broken glass should be disposed of in a puncture-proof container. Everything else may be thrown in the trash.

Solutions to the Combined Gas Law lab

Experiment one:

Step 1: Volume should be given in mL.
Step 2: The temperature of the room will be approximately 298 K, depending on how hot or cold you keep your classroom.
Step 4: The rubber glove will expand with a popping noise.

Step 5: This happened because the volume of the gas in the tube increased as the temperature of the tube increased.

Step 6: The volume of the gas in the test tube will be 142% of its original volume.

<u>Experiment two</u>:

Step 5: The stopper will be partially sucked into the test tube, making it hard to remove. Eventually, it will come free with a popping or sucking sound.

Step 6: If a cold water bath was used, 0.366 atm; if an ice bath was used, 0.353 atm.

Step 7: If the pressure inside the tube was one-third of an atmosphere, this would make the stopper harder to remove and cause a popping or sucking sound as the air outside of the tube rushed into it.

<u>Post-lab:</u>

1) Experiment one showed that when temperature increases, volume increases. Experiment two showed that when temperature decreases, pressure decreases. Both findings agree with the combined gas law.
2) Nothing unexpected should have happened, because all results were predicted by the combined gas law.
3) If we had closed the tube with a rubber stopper in experiment one, either the test tube would have exploded when the pressure inside increased or the stopper would have been launched out of the tube. Both of these are very dangerous.

<u>Solutions to the "Combined Gas Law Problems" worksheet</u>

1) 1.19 liters
2) 268 mm Hg
3) 267 cm^3
4) As the temperature increases, gas molecules move faster, pushing out from their original positions. As a result, the final volume is greater than the original volume.
5) If temperature were in degrees Celsius, any problem where temperature moved between positive and negative values would result in either the volume or pressure of the gas being negative, a physical impossibility.

Combined Gas Law Handout

The combined gas law is an equation that describes the relationship between the pressure, volume, and temperature of gases. The combined gas law states that:

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

P_1 = the initial pressure of the gas
V_1 = the initial volume of the gas
T_1 = the initial temperature of the gas

P_2 = the final pressure of the gas
V_2 = the final volume of the gas
T_2 = the final temperature of the gas

As the above suggests, the combined gas law is used whenever we're trying to figure out what would happen when we change the pressure, volume, or temperature of a gas. For example, the combined gas law would be good if we changed the temperature of a gas and wanted to find the new volume, but would not be good for figuring out how many moles of gas there are in a sample.

Here are some simple rules you should follow when using the combined gas law:

1) Always use Kelvins as your unit of temperature. If you use degrees Celsius, your answer will be wrong. (Incidentially, this rule goes for all gas law problems, no matter what equation you use).

2) Make sure that all varialbes with a subscript of "1" stand for the property of a gas before it has undergone a change in pressure, temperature, or volume. Make sure all variables with a subscript of "2" correspond to the properties of a gas after it has undergone a change. For example, if a gas was initially at a temperature of 500 K and the temperature was increased to 750 K, T_1 is "500 K" and T_2 is "750 K".

3) If a variable is kept constant in the problem or not mentioned at all, then you can leave it out of the equation entirely. For example, if a problem doesn't state the temperature of a gas, the combined gas law reduces to $P_1V_1 = P_2V_2$, which is Boyle's Law.

Finally, remember that using the combined gas law is no different than using Charles' or Boyle's law. After all, the "combined" in the combined gas law stands for the fact that it's just a combination of the two.

Combined Gas Law Lab

The combined gas law allows us to determine the relationships between the pressure, volume, and temperature of gases. In this lab, we'll investigate these relationships and test the validity of the combined gas law.

Experiment one:

1) Obtain a small test tube and find its volume by filling it with water. When you have determined the volume, pour out the water.

   The volume of our test tube was: ______________________

2) Find the temperature of the room in Kelvins: ______________________

3) Stretch a piece of rubber glove tightly over the opening of the test tube and fasten it securely with rubber bands. The rubber glove should appear nearly transparent when this is done correctly.

4) Heat the rounded end of the test tube over the Bunsen burner using test tube tongs. Write your observations below:

5) Why do you think this happened?

6) What was the volume of the gas inside your test tube at the end of this experiment? Assume the pressure remained constant and the temperature in the tube was 150^0 C.

Experiment two:

1) Very carefully remove the rubber glove from your test tube. Use caution, because it may still be hot.
2) Using a Bunsen burner and test tube tongs, heat the closed end of the test tube for about 60 seconds.

3) Remove the test tube from the heat, and while it is still hot, firmly place a rubber stopper in the opening of the test tube. It's best if the end of the stopper is moistened, as this will ensure a good seal.

4) Allow the test tube to cool to room temperature. Once it has cooled for about five minutes, place it into an ice bath or run it under cold water.

5) Remove the stopper from the test tube. Write your observations of what happened when you removed the stopper:

6) Calculate the pressure inside the test tube after it cooled. Assume that the pressure of the hot test tube was 1 atm and that the temperature of the hot tube was 500^0 C. Also assume that the temperature of the test tube was 10^0 C if you ran it under cold water or 0^0 C if you used an ice bath.

7) Does your finding from # 6 explain your observation from # 5? Explain.

Post-lab:

1) Did your observations in this lab agree with what the combined gas law predicts? Explain.

2) Did anything unexpected happen when you did this lab? Explain why or why not.

3) In the first part of this lab, we heated the test tube after stretching a rubber glove over the end. In the second part, we cooled it down after it had been closed with a rubber stopper. Why do you suppose we didn't close the test tube with a rubber stopper in the first part of this lab? Explain.

Combined Gas Law Problems

1) If a gas initially has a temperature of 600 K and a volume of 2.1 liters, what will the volume of the gas be if I cool it to a final temperature of 340 K? Assume the pressure of the gas remains constant.

2) A gas initially has a volume of 22.0 liters and a pressure of 230 mm Hg. If we change the pressure so the gas has a final volume of 18.9 liters, what will the new pressure be? Assume the temperature remains constant.

3) A gas initially has a volume of 45 cm^3, a pressure of 2.3 atm, and a temperature of 150 K. If we decrease its pressure to 0.8 atm and increase its temperature to 310 K, what will the new volume of the gas be?

4) Why does the volume of a gas increase as the temperature increases? Explain in terms of the motions of the gas molecules.

5) The units of degrees Celsius and Kelvins are both the same size, meaning that if the temperature has increased by 1.000^0 C, the temperature has also changed by 1.000 K. If both units are the same size, explain why we have to do gas law calculations using Kelvins instead of degrees Celsius.

Chapter 18 – Gas Law / Stoichiometry Lab

Overview

Stoichiometry is often treated as a topic completely apart from the gas laws. However, whenever a chemical reaction is done with gases, it's important that the stoichiometric relationships between gases are well understood. In this lab, students will combine sodium bicarbonate with acetic acid, creating aqueous sodium carbonate, liquid water, and carbon dioxide gas. When the carbon dioxide from this reaction is collected, the actual yield of this reaction will be compared to the theoretical yield to determine how well this reaction succeeded.

Teaching about gas stoichiometry

One of the problems students frequently have is that they forget what they've learned as soon as they start learning the next topic. To a large extent, this sort of selective forgetfulness is fostered by standardized tests that treat every academic subject as being wholly separate from the others. As a result, students have trouble relating what they've learned earlier in the year to what they're currently studying. This goes double for topics learned in other classes.

Fortunately, there are things you can do to help students keep current with topics they studied earlier in the year:

- **Repeat yourself!** In this lab, sodium bicarbonate and acetic acid react to form carbon dioxide, which will be collected in a rubber glove. Students will remember this procedure from the conservation of mass lab (Chapter 10), and this familiarity will make them more comfortable in this lab.
- **Review ideas individually before tying them together!** In today's lab, students will be familiar with stoichiometry and the ideal gas law, but may need to have both reviewed before being comfortable enough to perform the lab. Five minutes of review time before the lab starts will save thirty minutes of time in the lab.
- **Make it fun!** Whenever you need to tie two ideas together, do it in an exciting and interesting way! I had a student that enjoyed this lab enough that he asked to borrow a rubber glove, rubber bands, and a beaker so he could show his parents. Not only were his parents interested, but he also got an A on the next quiz!
- **Use the magic words: "You already know about this, so it will be a piece of cake."** It's amazing – when you tell students that something isn't difficult, the students believe it and have fewer problems. Never, *ever*, tell students that a topic is difficult. The harshest words you should ever have for a topic are that "This might take an extra couple of worksheets to master."

This lab is the classic example of a review lab that ties two big ideas together. By the time you perform this lab, you will have already covered stoichiometry and the ideal gas law in great detail. Students simply need to mentally mix these ideas to understand the lab. With a little bit of review, as well as an explanation of how stoichiometry and the ideal gas law fit together (which is explored in the student handout for this chapter), your students should have no trouble with this lab.

Doing the "Gas Law / Stoichiometry Lab"

Equipment:

- 150 mL glacial acetic acid
- 1000 mL distilled water
- 150 grams of sodium bicarbonate
- 10 – 250 mL beakers
- 5 – 100 mL graduated cylinders
- 2 spatulas
- 2 disposable pipettes
- 10 latex gloves
- 25 rubber bands
- 2 balances
- Several large pots or buckets (or 3 tape measures)

Safety:

Acetic acid is an eye hazard and may cause some skin irritation, particularly on broken skin. Keep a bottle of saturated sodium bicarbonate solution handy for spills or skin burns. If acetic acid is introduced to the eyes, flush immediately with water and call for medical assistance!

Room destruction factor:

The product of this reaction, sodium acetate, tends to be spilled on the countertops, leaving threadlike crystals behind when the water evaporates. Have students wipe down the countertops several times with damp and dry paper towels to minimize the formation of these crystals.

How the lab works:

The procedure for this lab is identical to that for the conservation of mass lab. Briefly, sodium bicarbonate will be placed in the finger of a rubber glove, then added to a dilute acetic acid solution. The subsequent formation of carbon dioxide will cause the glove to inflate. In this lab, however, the volume of gas will be measured and used to find a percent yield for the lab. This is done using stoichiometric calculations and the ideal gas law.

In step 7 of the lab, students need to find the volume of the rubber glove. This can be done in two ways. The first (and best) method is to immerse the inflated rubber glove in a large container filled to the top with water. The volume of the glove will displace an equal amount of water, causing the container to overflow. When the container is refilled with a graduated cylinder, the amount of water required to refill it is equal to the volume of the rubber glove. It is not necessary to detach the rubber glove from the beaker – simply turn the beaker upside down and immerse the glove in the bucket.

The second method for finding the volume of the rubber glove is by measuring the circumference of the glove and solving for the volume. This is done by pulling the glove off the beaker and tying it with rubber bands (a very tricky maneuver). If one treats the rubber glove as a perfect sphere, a measuring tape can be used to find the circumference of the glove in centimeters. To convert circumference to radius, use the equation $c = 2\pi r$. To solve for the volume of the glove from the radius, use $V = {}^{4}/_{3}\pi r^{3}$, which gives the volume in cubic centimeters.

In order for your students to do the ideal gas law calculations required for this lab, provide them with the atmospheric pressure and temperature before starting the lab.

<u>What can go wrong:</u>

- The glove leaks. Depending on the size of the leak, the students can either patch the glove with adhesive tape or start the lab over.
- The glove has inflated so much that it’s impossible to fit in the water-filled bucket to find its volume. If this happens, you’ll either need to get a bigger bucket or find the volume of the glove using a measuring tape.

<u>Clean up</u>:

The waste from this lab is harmless, as sodium bicarbonate and acetic acid are used in quantities that neutralize one another. The greatest problem with cleaning this lab up is that the sodium acetate formed in this reaction is difficult to get completely off the countertop, tending instead to form long, thin crystals. Though highly soluble in water, it may take a few rinsings to remove them entirely. Fortunately, sodium acetate is harmless.

Solutions for the Gas Law/Stoichiometry Lab

Prelab:

Some students will say that rubber gloves are good for collecting carbon dioxide, while others will disagree. Be prepared to give credit for either answer, as long as proper justification is given.

Lab procedure:

Step 1: The weight of the flask should be given in grams.
Step 6: A complete description of the reaction should be give, with observations such as foaming and bubbling, fast inflation of the glove, and the formation of small bubbles in the solution even after most of the reaction is finished.
Step 7: The volume of the glove is usually in the area of 600 mL.

Questions:

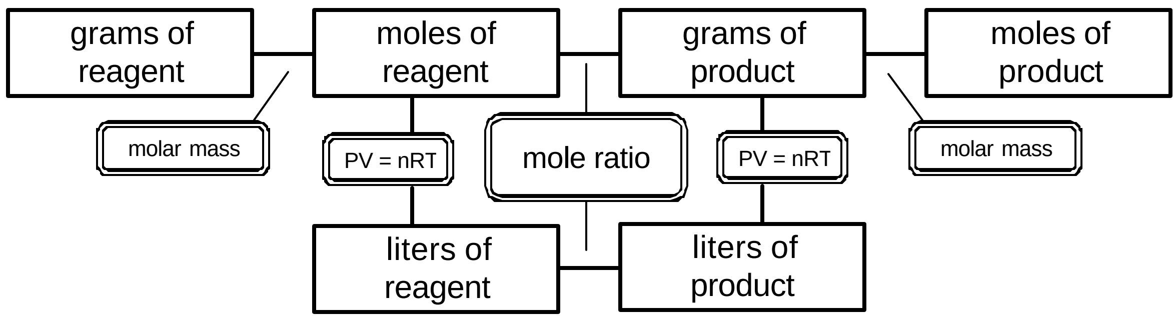
1) 0.125 moles CO_2
2) At 298 K and 1 atm, 3.06 L. Your answer will depend on local conditions.
3) To find percent yield, divide the volume from step 7 by the answer from #2 above, and multiply your answer by 100. A typical answer is 20-50%.
4) If the percent yield is less than 100%, it's likely that some of the gas in the glove leaked out. If the percent yield is greater than 100%, there are probably computational errors. If the percent yield is exactly 100%, students have probably invented their data.

Solutions to the Gas Law – Stoichiometry Worksheet

1) 45.9 grams
2) 14.1 liters
3) Good "real world" examples include any chemical reaction that involves the production of a gas, such as the combustion of gasoline or the production of soda pop.

Gas Law – Stoichiometry Calculations

We've already learned how to do stoichiometry calculations by following the necessary steps outlined on a map. To do stoichiometry calculations involving gas law equations, use the map below:



To use this map, you need to identify where you are starting in the diagram and where you need to end up. This is just like taking a long car trip – you need to look at a map and figure out where your starting point and destination are. To make the necessary conversions, you need to do one calculation for every road you travel on. For example, if you go from grams of reagent to liters of product, you'll need to do the following three calculations:

- Grams of reagent to moles of reagent. The conversion factor is the molar mass of the reagent.
- Moles of reagent to moles of product. The conversion factor is the mole ratio between the reagent and product, determined by looking at the coefficients on the chemical equation for the reaction you're studying.
- Moles of reagent to liters of product. To convert from moles to liters, use the ideal gas law, PV = nRT.

As with many car trips, there's another way to get from one place to another. By following the three calculations below, you can also go from grams of reagent to liters of product:

- Grams of reagent to moles of reagent. The conversion factor is the molar mass of the reagent.
- Moles of reagent to liters of reagent. To convert from moles to liters, use the ideal gas law, PV = nRT.
- Liters of reagent to liters of product. The conversion factor is the mole ratio between the reagent and product, determined by the coefficients in the chemical equation for the reaction you're studying.

Either way will give you the same answer, just as there are usually many ways to travel from one spot on a map to another.

Gas Law / Stoichiometry Lab

In class, we've been talking about more difficult stoichiometry problems, where we add liters to the list of units between which we can easily convert. In this lab, we'll get a chance to see how this works in a laboratory setting.

In this lab, you will be mixing sodium bicarbonate and acetic acid to produce sodium acetate, water, and carbon dioxide. The equation for this reaction is:

$$C_2H_3O_2H_{(l)} + NaHCO_{3(aq)} \rightarrow NaC_2H_3O_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

You will start with a known amount of acetic acid and sodium bicarbonate. Your job will be to predict how much carbon dioxide you should have produced and you will compare this theoretical prediction to the amount you collect in this lab.

Prelab:

Is collecting the carbon dioxide in a rubber glove the best way to find the volume of CO_2 formed? Explain why or why not.

Lab procedure:

1) Find the weight of a 250 mL beaker: ________________________________

2) Pour 60 mL of distilled water into the beaker using a graduated cylinder.

3) Carefully place 7.5 grams of acetic acid into the beaker. Caution: Although acetic acid smells the same as vinegar, it's approximately 50 times more concentrated. Use caution!

4) Place 10.5 grams of sodium bicarbonate into one of the fingers of a rubber glove. Use rubber bands to tie the finger shut.

5) Fasten the rubber glove to the opening of the beaker using at least two rubber bands. The seal you make between the glove and flask must be strong enough to keep gas from escaping!

6) When the rubber glove is securely fastened to the beaker, untie the finger of the rubber glove and let the sodium bicarbonate fall into the beaker. Write your observations of what happens below:

7) When the reaction stops, find the volume of the rubber glove using the method that your teacher described before starting this lab.

Volume of the rubber glove: ______________________________

8) Clean up: Wash the residue from the beaker and clean any spilled solution from your lab table. The rubber glove and rubber bands may be thrown away.

Questions:

1) From the weight of acetic acid used in this reaction, use your knowledge of stoichiometry to predict how many moles of carbon dioxide you should have produced:

2) From the number of moles of carbon dioxide you should have produced, use the ideal gas law to determine how many liters of CO_2 gas should have been produced (theoretical yield):

3) Calculate your percent yield of carbon dioxide for this reaction, using the actual yield you found and the theoretical yield you calculated in question 2. Show your work.

My percent yield for this lab was: ____________ *%*

4) Was your percent yield exactly 100%? If it was less than 100%, explain where you think the lost carbon dioxide went. If it was more than 100%, explain how you made more carbon dioxide than predicted. Your explanation should be complete in order to receive full credit.

<u>Gas Law – Stoichiometry Worksheet</u>

Use your knowledge of the ideal gas law and simple stoichiometry calculations to answer the following questions. Assume all reactions are done at a temperature of 25^0 C and a pressure of 1 atmosphere.

1) Consider the reaction between sulfur and oxygen to form sulfur dioxide:

$$S_{8(s)} + 8\ O_{2(g)} \rightarrow 8\ SO_{2(g)}$$

How many grams of sulfur will be required to make 35 liters of sulfur dioxide?

2) Consider the reaction between ammonia and sulfuric acid to form ammonium sulfate:

$$2\ NH_{3(g)} + H_2SO_{4(l)} \rightarrow (NH_4)_2SO_{4(s)}$$

How many liters of ammonia will be required to make 38 grams of ammonium sulfate?

3) When do you think it might be useful to convert between liters, grams, and moles? Explain your answer, giving “real world” examples.

Chapter 19 – Solutions and Molarity

Overview

Most chemistry takes place in aqueous or organic solvents. The behavior of solutions depends largely on the amount of solute present, which can be expressed in both qualitative and quantitative terms. In this lab, students will use the expression for molarity to figure out how many mL of a 2.00 M NaCl solution is required to get exactly 1.00 grams of NaCl.

Teaching about solutions

Solution chemistry is probably the largest topic covered during the course of a general chemistry class. Not only do students learn about concentration, but they also learn about equilibria, acids and bases, solubility, and colligative properties. Many students have problems learning the wide range of solution-related topics, making it more difficult to understand how chemicals interact with one another.

When initially teaching about solutions, it's important that you cover a few basic topics. I've included these below, as well as explanations that work well with students:

Q: What's a solution?

A: A solution is any mixture in which one compound dissolves another. It's a homogeneous mixture, meaning that every part of the mixture has the same composition as every other part. The chemical that is dissolved is the "solute" and the chemical that does the dissolving is the "solvent". Generally, solutes are solids and solvents are liquids, though there are exceptions.

Q: What's the difference between solutions and other mixtures?

A: Solutions are special types of mixtures in which the solvent pulls solute molecules apart from one another. Examples include salt water, where sodium and chloride ions are pulled apart and sugar water, where water molecules pull adjacent sugar molecules away from each other.

This is different from other mixtures. In heterogeneous mixtures, components don't mix well with each other. Imagine granite, wood, or jellybeans – these mixtures clearly have different components in them. These visible differences occur because the components have little interest in interacting with one another.

In suspensions, many small particles drift around in a liquid. At first, suspensions are uniformly cloudy, but eventually (hours or days later), the liquid and solid separate. Examples of suspensions are milk and mud.

Q: How is concentration expressed qualitatively?

A: Solutions can be described as unsaturated (more solute can still be dissolved), saturated (the maximum amount of solute has been dissolved), or supersaturated (more solute is dissolved than theoretically predicted). If a solution is said to be either unsaturated or supersaturated, very little information has been given about the true concentration – for example, a tablespoon of sugar in a cup of tea and a tablespoon of sugar in a swimming pool both create unsaturated solutions with very different concentrations. The term "saturated" indicates a precise concentration that can be looked up.

It is hard for many students to understand how a supersaturated solution can be stable. The answer, of course, is that supersaturated solutions aren't stable – the only reason they exist at all is that there are no small particles in the solution (nucleation centers) that dissolved molecules can stick to in order to form crystals. The addition of a single small crystal can cause a supersaturated solution to undergo very fast crystal growth until it is merely saturated. Supersaturated solutions are rarely observed in the "real world" because of this inherent instability.

Supersaturated solutions can be formed in two ways. The first way involves the slow evaporation of solvent from a saturated solution. The second method requires the preparation of a very hot saturated solution, which is then cooled to form a supersaturated solution (recall that the solubility of a solid decreases as temperature decreases).

Q: How is concentration expressed quantitatively?

A: Concentration is most often expressed by molarity (M), defined as the number of moles of solute per liter of solution. Students generally don't have trouble calculating molarities unless they have trouble making conversions between grams and moles or milliliters and liters.

Other units of concentration include molality (m, moles of solute per kilogram of solution) and parts per million (ppm, mg of solute per kilogram of solution). Though important, these are much less common units in a high school chemistry class.

Q: What are colligative properties?

A: Colligative properties are properties that change when the solute concentration changes. Melting point, boiling point, and vapor pressure are the three colligative properties usually discussed. However, other properties such as taste, color, and texture probably also qualify as colligative properties – at the very least, they make it easier for students to understand what colligative properties are.

Doing the Molarity Calculation Lab

Equipment:

- 1 L of 2.00 M sodium chloride solution (made with distilled water)
- 6-10 mL graduated cylinders
- 6-50 mL beakers
- 6 ring stands with rings
- 6 squares of wire gauze
- 6 Bunsen burners
- 2 flint strikers
- 2 balances
- 6 beaker tongs

Safety:

The main safety issue deals with the Bunsen burners. As always, remind students that lab equipment remains hot long after the heat has been turned off. Rubber gloves should not be worn at any time during this lab.

Students should also be careful not to heat the 50 mL beaker after the water has been boiled from the solution or the beaker will shatter.

Room destruction factor:

None. This lab won't hurt your room at all!

How the lab works:

In this lab, students will be heating a 2.00 M NaCl solution to remove the water – their goal is to obtain exactly one gram of solid sodium chloride. Before this is done, they will need to figure out how much of the sodium chloride solution is required. To do this, they will use the equation M =

moles/liters, with M = 2.00 M NaCl and moles = 0.17 moles of NaCl (equivalent to one gram of sodium chloride). It's not a bad idea to tell students that they may only use a maximum of forty mL of solution – this limits the amount of trouble that can be caused by miscalculations.

The hardest part of this lab for students to understand is the second step in the procedure where they have to predict experimental errors. Because students want to isolate exactly 1.00 grams of sodium chloride, they will need to identify what will cause errors with their procedure and compensate for these errors by using slightly more or less NaCl solution than the calculation indicates. These changes in the amount of solution should be no more than 2 mL, as only 8.6 mL of solution are being boiled in the first place.

Poor sources of error include "we might spill some solution", "the balance didn't work right", "the solution might not be exactly 2.00 M", and "our calculations might have been wrong." Spilling the solution doesn't introduce experimental error to this lab because it would require students to start over and the other sources of error are unlikely or unreasonable.

There are two reasonable sources of error in this lab that students should discuss. The first is that water will remain in the sodium chloride after it has been heated (much as in the conservation of mass lab). This error is particularly vexing because there's no way to compensate for it by adding or subtracting NaCl solution. After all, if the measured mass is 1.50 grams because of additional water, only 1.00 grams of this is actually sodium chloride – the use of less solution in this process may give a measured weight of 1.00 grams, but would result in less than 1.00 grams of sodium chloride. It is important that students know that the weight they measure at the end isn't necessarily all sodium chloride.

The second reasonable source of error is that some of the sodium chloride will splatter out of the beaker while the solution is being heated. A small amount of sodium chloride solution (~0.5 mL) should be added to compensate for this error.

<u>What can go wrong:</u>

- Students determine that they need 5 liters of solution to get 1.00 grams of sodium chloride solution. By limiting the amount of 2.00 M NaCl they can use to 40 mL, this problem is eliminated.
- The beaker shatters when heated over the Bunsen burner. Fortunately, the pieces can usually be collected and the data salvaged, so long as they didn't travel too far across the room.

Clean up:

The contents of all beakers can be safely rinsed down the drain.

Solutions to the Molarity Calculation Lab

Prelab:

0.501 liters (501 mL)

Procedure:

Step 1: Because M = 2.00 M, and there are 0.017 moles of NaCl in one gram, the amount of NaCl solution needed is 0.0086 L, (8.6 mL).
Step 2: A reasonable error includes splattering during the boiling process (requiring additional NaCl solution). Others should be considered, based on how reasonable they are.
Step 3: This is the sum of the answers from steps 1 and 2.
Steps 4 and 7: Masses should be given in "grams".

Postlab:

Answers greater or less than 100% are caused by an incorrect estimate of experimental error for this lab. Students should provide a complete explanation of how their estimates of the errors were incorrect. Answers of exactly 100% should be examined to ensure that the data hasn't been manipulated, as perfect accuracy is extremely uncommon.

Solutions for the Molarity Worksheet

1) Solvents are liquids that dissolve other compounds. Solutes are solids that are dissolved by other compounds. In Kool Aid, the solvent is water and the solutes are sugar, coloring, and pureed insect bits.
2) Solutions are homogeneous mixtures in which one compound (the solute) is pulled apart by the other (solvent). Suspensions are mixtures in which small particles of (usually) solid drift around in a liquid, forming a milky mixture. Unlike solutions, which are stable, suspensions separate back into their separate phases after minutes or hours. An example of a solution is grape juice and an example of a suspension is milk.
3) 0.365 M.
4) Add water to 0.0158 grams of calcum chloridee until the total volume of the solution is 48 mL.
5) 23.61 grams.

Solutions Handout

Solutions are one of the most important topics in all of chemistry. Most chemical reactions take place in aqueous solutions, including biochemical reactions in our blood. Because the amount of stuff dissolved in a liquid has a huge effect on how that liquid behaves, we need to know how to define concentration.

The first method for measuring concentration is qualitative, meaning that it describes things in words rather than numbers:

- **Unsaturated** solutions are solutions that can still dissolve more solute. If you place sugar in your iced tea you've made an unsaturated solution, because if you were to add more sugar, it would also dissolve.
- **Saturated** solutions are solutions that have dissolved as much solute as they possibly can. If you added sugar to your iced tea until it was saturated, any sugar you add after that point sinks to the bottom, never dissolving.
- **Supersaturated** solutions contain even more solute than saturated solutions. This happens either because some of the solvent has evaporated from a saturated solution or a hot saturated solution has been cooled. Because there is more solute in a supersaturated solution than is theoretically possible, these solutions tend to form crystals when disturbed.

The second method for measuring concentration is quantitative, meaning that you can find an exact numerical value for it. In this method, concentration is expressed as **molarity**, and given the symbol **M**. Molarity is found by solving the following equation:

$$\mathbf{M} = \frac{\textbf{moles of solute}}{\textbf{liters of solution}}$$

For example, 1 liter of solution containing 1 mole of sodium chloride is referred to as a "one molar sodium chloride solution", or a "1 M NaCl solution". In this case, "one molar" means that the solution has a molarity of one.

In some cases, you may have to calculate using grams instead of moles, or milliliters instead of liters. Remember: To find molarity, you must **always** use "moles" and "liters", or your answer will be wrong! Fortunately, it's easy to convert between moles and grams using the molar mass as a conversion factor. To convert from milliliters to liters, use the conversion factor 1 mL = 0.001 L.

Molarity Calculation Lab

In this lab, you will be using the expression M = moles / liters to determine how many milliliters of 2.00 M NaCl solution you'll need to isolate 1.00 grams of sodium chloride. This is more difficult than it initially seems because not only will you have to make a molarity calculation – you'll also have to make adjustments to your final answer to compensate for anticipated experimental error.

The thing to remember about this lab is that whenever you do a calculation, you're computing how things *should* occur if everything happens with 100% accuracy. After you've made a theoretical determination about what you should do, you'll need to change this value to compensate for imperfections in your lab technique.

Prelab:

How much 0.100 M calcium bromide solution would be required to obtain 10 grams of pure calcium bromide?

Procedure:

1) In this lab, you will be using a 2.00 M NaCl solution as your source of sodium chloride. Calculate how many liters of 2.00 M sodium chloride solution are required to obtain 1.00 grams of solid sodium chloride.

Amount of 2.00 M NaCl solution needed: ________________________

2) Whenever people do a lab, they make mistakes that affect their final result. To compensate for experimental error, you will need to make adjustments to the theoretical quantity of 2.00 M NaCl solution you calculated in step 1. If you believe an error will cause you to have an abnormally low answer, you will need to add a little extra NaCl solution to compensate for it. If you believe an error will cause you to have an abnormally high answer, you will need to remove some NaCl solution to compensate. Use your best judgement to decide how much NaCl solution should be added or subtracted.

In the chart below, indicate your expected sources of error and how much 2.00 M NaCl solution you'll add or subtract for each one:

Anticipated Error(s)	**Adjustment to the amount of 2.00 M NaCl solution to compensate for error**

3) Using your answer from step 1 and the adjustments you decided that you needed from step 2, write down the total amount of 2.00 M NaCl solution you'll need for this lab:

Adjusted quantity of 2.00 M NaCl solution required: ________________

4) Find the mass of a 50 mL beaker: ________________________

5) Using a graduated cylinder, measure the amount of solution you think will be required to get exactly 1.00 grams of NaCl. Place this solution in your 50 mL beaker.

6) Heat the beaker containing your solution over a Bunsen burner until all of the solvent has boiled away. Make sure that you don't heat the beaker after all of the solvent has been removed, because the beaker will shatter.

7) When the water has been removed, you should have a thin crust of sodium chloride on the bottom of your beaker. To find the exact mass of sodium chloride rpesent, you will need to weigh the beaker and compare it to the weight of the empty beaker.

The mass of the beaker after heating was: ______________________

The mass of the salt in the beaker was: ______________________

8) Clean up: Rinse the salt residue from the beaker with water and pour it down the sink.

Postlab:

Did you get exactly 1.00 grams of salt?

- If your answer was exactly 1.00 grams of sodium chloride, explain whether you believe you actually have 1.00 grams of sodium chloride present, or if you think something else may have occurred.
- If your answer was less than 1.00 grams of sodium chloride, explain where the rest of the sodium chloride went and how you could prevent this sort of error in the future.
- If your answer was greater than 1.00 grams of sodium chloride, explain why the mass is greater than it should be and how you could prevent this sort of error in the future.

Molarity Worksheet

1) What is the difference between a solvent and a solute? Explain, using an example from around your house.

2) What is the difference between a solution and a suspension? Give one example of each you might find around your house.

3) Give the molarity of 480 mL of a solution containing 28 grams of $CuSO_4$.

4) Explain how to make 48 mL of a 0.003 M calcium chloride solution.

5) How many grams of magnesium hydroxide are required to make 540 mL of a 0.75 M solution?

Chapter 20 – Solubility Lab

Overview

There are two main factors that affect the solubility of a compound in water. The first is the famous “like dissolves like”, which means that polar solvents tend to dissolve polar solutes. The other is temperature, as higher solvent temperatures lead to higher solubility of solids and lower gas solubility. In this lab, students will investigate “like dissolves like” by dissolving KCl in isopropanol and water.

Teaching about solubility

Water is frequently referred to as “the universal solvent”, because it’s good for dissolving a wide variety of solutes. However, it’s clear that water isn’t truly a “universal” solvent, as many objects don’t dissolve in water. Your house, for example, is unlikely to dissolve when it rains.

Students intuitively understand that some things dissolve in water and some don’t. However, the mechanics of solubility is something much less easy to understand. The following are questions that students frequently have about solubility, and their answers:

Q: Why does *anything* dissolve?

A: Compounds dissolve when it’s energetically favorable for solute molecules to be attached to solvent molecules, rather than each other. This explanation, though completely true, tends to confuse students who aren’t used to thinking in terms of thermodynamics.

Instead of speaking in terms of energy, let’s discuss what molecules “want” to do. Strictly speaking, molecules don’t have any desires at all, but this does a much better job of helping students understand why molecules dissolve than a discussion of thermodynamics. After your students understand solubility in terms of like and dislike, you can translate this into thermodynamics.

There are two main things that can happen when you place a solid into a liquid: Either it will dissolve or it won’t. If it dissolves, the solute molecules prefer to be stuck to solvent molecules than to each other. If it doesn’t, the solute molecules prefer to be stuck to one another than the solvent.

Why might molecules want to stay undissolved? Imagine a very small child clinging to its mother. If somebody comes up to the child and asks for a hug, the child will cling even more tightly to mom. Clearly, in this case, the child prefers to stay where it is because its interaction with mom is more valuable than its interaction with somebody else. *In chemical terms, some compounds consist of molecules or ions that have stronger attractions to each other than to solvent molecules. As a result, they will tend to remain as a tightly-bonded solid, rather than dissolve and become a solution.*

On the other hand, let's imagine your teenage daughter. If given the choice between spending Saturday night playing Parcheesi with you and dancing the night away with some "cute guy" from the local university, it's clear what she will do. *In chemical terms, this corresponds to a case where the molecules or ions have a stronger attraction to solvent molecules than to each other. As a result, these molecules tend to dissolve and form solutions.*

Q: What does "polar" mean?

A: Polarity is when there's more electrons in one part of a molecule than another. Asymmetric molecules are more polar than symmetric molecules. Molecules that are completely symmetric are completely nonpolar.

The question, then, is how to define "symmetry". If you're an inorganic chemist, there are many levels of symmetry defined by point groups. The definition of symmetry we're going to use is much simpler.

To determine how symmetric a molecule is (and thus, it's polarity), we need to do the following steps. We will use the examples of carbon dioxide and water.

- **Draw the Lewis structure of the molecule**. If students have trouble with Lewis structures, this would be a good time to review them:

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 ..    ..                 . .
: O = C = O :           H – O – H
                          . .
```

carbon dioxide   water

- **Change the Lewis structure so it's as asymmetrical as the VSEPR theory allows**. This rule allows us to simulate the 3-D structure of a molecule in two dimensions. Recall that the VSEPR theory states that the electrons in a molecule arrange themselves to they're as far apart from each other as possible. In the above Lewis structures, we've drawn these molecules so that the electrons on the central atom are as far apart from each other as is possible. In carbon dioxide, the oxygen atoms are shown 180^0 apart from one another and in water the electron pairs are shown 90^0 apart from one another (the 90^0 angles are the best approximation that we can make of a 3-D molecule in a 2-

D diagram). In carbon dioxide, it's clear that the molecule is fairly symmetric and can't be changed much (the only thing that can be done around the carbon atom is to switch the two oxygen atoms with each other – to move one oxygen atom within 90^0 of the other would violate the rules of VSEPR). In water, on the other hand, you can make the molecule less symmetric. Taking a look at the diagram, the hydrogen atoms are 180^0 apart from one another, making the molecule appear symmetric. However, if we switch one of the hydrogen atoms with one of the lone pairs on oxygen, (see the diagram below) we get a structure that both follows VSEPR (the electron pairson the molecule are still 90^0 apart) but isn't symmetric (because the hydrogen atoms are next to each other). This shows that water is asymmetric.

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     ..
H – O :
     |
     H
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In this diagram, the electron pairs are still 90^0 apart, but the hydrogen atoms are now also 90^0 apart from one another. Because VSEPR treats bonds and lone pairs as equivalent, this is allowed when drawing Lewis structures.

- **Imagine that the molecule is a wheel and give it a spin. The more wobbly the molecule, the more polar it is.** Treat the central atom in the molecule as being the hub of a wheel. Imagine spinning the molecule around this hub. If the molecule wobbles, the molecule is polar. The more a molecule wobbles, the more polar it is. Carbon dioxide, for example, is not wobbly at all because the oxygen atoms are 180^0 apart from one another. Water, on the other hand, wobbles badly because the hydrogen atoms are shown 90^0 apart from each other – if they were weights on a wheel, the wheel would wobble. As a result, water is polar and carbon dioxide is not. In the same way, boron trichloride is nonpolar, ammonia is somewhat polar, and hydrogen fluoride is extremely polar.

This little trick works for all molecules, though it's easier for some than others. For example, when working with large organic molecules, it's frequently difficult for students to draw the entire structure. As a result, I tell my students that if an organic molecule is too large to draw, assume it's fairly nonpolar.

Q: Why does "like dissolve like"?

A: Polar solvents dissolve polar solutes because polar solvents have positive and negative charges on them caused by differences in electronegativity. Polar solutes also have positive and negative charges – when a positive charge from a polar solvent lines up with a negative charge from a polar solute, they attract one another and the solute is pulled into solution. This is further helped along by entropy, which energetically favors any process that makes the universe more random.

Nonpolar solvents dissolve nonpolar solutes because attractions between the molecules in both the solvent and the solute aren't very strong. As a result, there isn't much force holding the molecules together, making it easy for the entropically-favored process of solvation to take place. Unlike the solvation of polar solvents in polar solutes, there really isn't much attraction between the solvent and solute molecules in nonpolar materials.

Why don't polar and nonpolar compounds dissolve each other? Polar solvent molecules don't grab on to nonpolar solutes because they'd rather be grabbing other polar solvent molecules. This also explains why nonpolar solvents can't dissolve polar solutes – after all, the attraction of positive and negative charges will be a lot stronger than the nonexistent attraction of neutral charge to positive and negative charges.

Q: Why can hot solvents dissolve more solute than cold solvents?

A: Whenever energy is added to a system, the particles in the system behave in a more energetic fashion. When you heat a solution, the solids tend to act more energetically, causing the particles to break free from one another. Likewise, the solvent molecules will have more energy with which to grab the solute particles. Both processes working together help to increase the solubility of products.

On the other hand, gases become less soluble as you increase a solution's temperature. When more energy is added to a solution that contains a dissolved gas the gas molecules gain energy, causing them to break free from solvent molecules. As a result, the solubility of gases decreases as the temperature of the solution increases.

Doing the Solubility Lab

Equipment:

- 250 mL pure isopropanol
- 250 mL distilled water
- 6-10 or 50 mL graduated cylinders
- 6 glass stirring rods
- 40 grams KCl (NaCl can be substituted if KCl is not available)
- 14-50 mL beakers
- 6 beaker tongs
- 4 grease pencils for labeling beakers (or masking tape and ballpoint pens)
- At least 2 balances
- 500 or 1000 mL beaker for storing waste
- 4 hotplates

Safety:

The primary safety concern in this lab deals with the step in which excess isopropanol is removed from the KCl sampler using a hot plate (step 4). Make absolutely sure that students pour as much of the isopropanol into the waste container before beginning this lab to minimize the risk of fire. Hot plates should also be kept at a low temperature setting and placed in the fume hood. As always, students should wear goggles.

Room destruction factor:

This lab is easy to perform and causes very little mess aside from spilled KCl.

How the lab works:

In this lab, students will be asked to find whether water or isopropanol is more polar based on how well each dissolves potassium chloride. The quantity of potassium chloride dissolved will be determined by allowing both solvents to pass over a fixed amount of KCl and measuring the amount of KCl remaining.

Students sometimes have problems using “like dissolves like” in the laboratory. In this lab, for example, some students may believe that because isopropanol and water both have some polarity, KCl will dissolve equally well in each. Though it’s true that both are polar, isopropanol is only very slightly polar, making it a poor solvent for KCl.

You may want to use an analogy to explain this phenomena to students. When epidemics of flu go around, people either get the flu or they don’t. The condition of “getting the flu” is like “polarity” – molecules are either polar or they are not. However, if you’re one of the people that gets the flu, it’s less clear how you’ll feel. You could be one of the lucky people that gets a runny nose for a day or two before the flu goes away, or you could be one of the unlucky people that’s in bed for a week with the worst flu of your life. This is how polarity works: Molecules are either polar or nonpolar, but there can be big differences between two polar molecules. One compound may be very slightly polar and be better able to dissolve nonpolar compound than ionic compounds because their polarities are more similar. On the other hand, another compound may be so polar that it completely ignores nonpolar compounds and vigorously dissolves ionic compounds. “Like dissolves like” deals with similarities in polarity – sometimes a nonpolar compound will be similar in polarity to a polar compound, allowing them to dissolve one another.

One question that gives students trouble is question 4 in the postlab, which asks whether the increase in boiling point over pure solvent due to the addition of KCl would be greater for isopropanol or for water. The typical answer is that the increase in boiling point would be greater for isopropanol because the boiling point of isopropanol is lower than that of water. Before doing this lab, it should be made clear to students that colligative properties depend on concentration, not on the temperature at which the solvent normally boils.

What can go wrong:

- Isopropanol dissolves KCl and water doesn't. This is caused when students mislabel their beakers.
- Both isopropanol and water dissolve KCl. Again, this is caused when students mislabel their beakers.
- The KCl you planned to use for the lab has clumped up in a big blob and you can't get it out of the reagent bottle. This occurs when water gets into the bottle, causing it to partially dissolve and recrystallize in one big chunk. To fix this, turn a laboratory oven up to about 70 degrees and heat the KCl overnight with the lid removed. This should remove enough of the water that the KCl can be easily broken up with a spatula. If this doesn't work, you may have to break the reagent bottle and break up the chunk with a mortar and pestle.
- The students can't figure out which beaker on the hotplate is theirs. Make sure your students are good about labeling their beakers to ensure the correct beaker can be found at the conclusion of the lab.
- The beakers still contain solvent after sitting on the hotplate for three minutes. This problem is easily solved by leaving the beakers on the hotplate until all of the solvent has evaporated.

Clean up:

There should be no problem with putting the waste from this lab down the sink.

Solutions to the Solubility Lab

Lab procedure:

1) KCl is more soluble in water than in isopropanol because water and KCl are both polar and isopropanol is nonpolar. "Like dissolves like".
2) The KCl will dissolve and the solution will turn momentarily cloudy.
3) The KCl will not dissolve, tending to clump up at the bottom of the beaker.
4 & 5) All weights should be given in grams.
6) The amount of KCl remaining will be equal to the answer from step 4 minus the answer from step 5. The amount of KCl remaining should be significantly smaller for the aqueous solution than for the isopropanol.

7) If students initially predict that KCl will dissolve in water, then their prediction was correct. If not, they should explain why they were wrong in terms of "like dissolves like".

Postlab questions:

1) Methane would dissolve better in isopropanol because both are nonpolar.
2) Polar solvents dissolve polar solutes because both the solvent and solute molecules have regions of partial positive and negative charge. When the partial positive charge on the solvent lines up with the partial negative charge on the solute (or vice-versa), the solvent will pull the solute into solution, dissolving it.
3) If we had used a Bunsen burner, it would have ignited the isopropanol.
4) Water would have a greater change in boiling point, because the water dissolved more KCl than the isopropanol did. The boiling point of a solution increases proportionally to the amount of solute dissolved in it.

Solutions to the Solubility Worksheet

1) Polarity refers to the uneven distribution of electrons within a molecule, causing regions of partial positive and negative charge. This uneven distribution of electrons is caused by differences in electronegativity between the atoms in the molecule.
2a) polar
2b) polar
2c) nonpolar
3a) nonpolar
3b) polar
3c) nonpolar
4) Sugar should be added when the water is still hot, because hot solvents are able to dissolve more solute than cold solvents.
5) Colligative properties are any properties that change as a function of solute concentration. These include melting point, boiling point, and vapor pressure.

Solubility Handout

To understand how solutions behave, we first need a basic understanding of what makes solute molecules dissolve in the first place.

When compounds dissolve, the molecules in the solid are pulled apart from each other by the solvent molecules. Any process that can help the solvent molecules to separate solute molecules will increase solubility.

One thing that helps solvent molecules pull solute molecules apart is the phenomenon that “like dissolves like”, meaning that solvents tend to dissolve solutes with similar polarities. Polar solvents dissolve polar solutes, and nonpolar solvents dissolve nonpolar solutes. Solvents that are polar aren’t able to dissolve nonpolar solutes.

Though everybody can agree that the saying “like dissolves like” is easy to remember, its meaning is a little more complex than a three word slogan would suggest. Let’s approach it bit by bit to see if we can understand why it works:

What is polarity?

Polarity is when electrons spend more time on one side of a molecule than another. Molecules are said to be polar when the atoms on one side of the molecule have higher electronegativities than the atoms on the other side. Polarity can be determined by examining the Lewis structures of chemical compounds – if the structures aren’t completely symmetric, the molecule has some polarity. If a polar molecule contains atoms with very dissimilar electronegativities, it’s said to be very polar, while if it contains atoms with similar (but not the same) electronegativities, it’s only slightly polar. See the diagram below for some examples:

```
    H
    |
H – C – H
    |
    H
```

Methane is nonpolar, because the hydrogen atoms are arranged symmetrically around the carbon atom.

```
   . .
H – N – H
    |
    H
```

Ammonia is polar, because the hydrogen atoms are not arranged symmetrically around the nitrogen atom.

```
 . .       . .
: O = C = O :
```

Carbon dioxide is nonpolar, because the oxygen atoms are arranged evenly around the carbon atom.

Why do polar solvents dissolve polar solutes?

All polar molecules contain areas with partial positive charges and partial negative charges. The partial negative charges correspond to electronegative atoms that are good at pulling electrons away from other atoms, while the partial positive charges are the atoms with low electronegativities. If a partial negative charge on a solvent molecule comes near the partial positive charge on a solute molecule, the two molecules will attract. If the attraction is strong enough, the solvent molecule will pull the solute molecule from the solid, causing it to dissolve.

Why do nonpolar solvents dissolve nonpolar solutes?

In nonpolar molecules, there are no partial negative or positive charges that allow the molecules to stick to other molecules. As a result, the interactions between nonpolar molecules are generally weaker than those between polar molecules. Why, then, would a nonpolar solvent grab onto a nonpolar solute?

The answer is that there's something else that causes nonpolar solutes to dissolve. In addition to the charges on molecules, there is a term called "entropy", that corresponds to the randomness a system has. To give an example, a group of pennies would have low entropy if they were all turned to the "heads" side, while having much higher entropy if they were randomly distributed between heads and tails.

For reasons you'll discuss later, any process that increases the entropy of a system is energetically favored. As a result, if solute molecules have a choice between existing in a solid form (which is relatively nonrandom) or a dissolved form (which is very random), entropy provides an energetic boost for the molecules to dissolve and become more random.

Why don't polar solvents dissolve nonpolar solutes?

Polar molecules have partial positive and negative charges, while nonpolar molecules do not. As a result, there is no real attraction between polar and nonpolar molecules, and they do not dissolve.

You may be wondering why entropy doesn't cause a nonpolar solute to dissolve even if it isn't attracted to a polar solute. After all, if it dissolved, wouldn't the solid become more random? It would, and entropy favors having a nonpolar solute dissolve in a polar solvent for the same reason that it favors a nonpolar solute dissolving in a nonpolar solvent.

The reason that nonpolar solutes don't dissolve in a polar solvent is because there's another force present in the solution. Remember how we said that polar molecules have partial positive and negative charges? Well, if we place a nonpolar solute in a polar solvent, it can't dissolve because the polar solvent molecules will be so busy grabbing onto each other that it's impossible for nonpolar solute molecules to even mix with them.

In other words, there are two things going on when you place a nonpolar solute in a polar solvent. Entropy is a driving force for the nonpolar solute molecules to dissolve, and the interaction between polar solvent molecules is a force that helps to keep the nonpolar solute molecules from dissolving. As it turns out, the energy provided by entropy isn't enough to force the solvent molecules apart from each other. As a result, the solid won't dissolve.

What else can cause molecules to dissolve?

Solid solutes can be made more soluble in water by increasing the temperature of the solution. What happens is that the energy given to the solvent molecules when the temperature increases makes it easier for them to pull the solute molecules apart. This phenonmenon can be observed when dissolving sugar in tea – it's much easier to dissolve sugar in hot tea than in iced tea.

When a gas is dissolved in a liquid, it becomes less soluble when the solution is heated. This is because gas molecules gain enough energy that they are able to break free from the solution entirely, an option that isn't possible with solid solutes. If you've ever boiled water, you may have observed small bubbles forming in the pot before the water even got very hot – these bubbles are not steam, but small air bubbles that can no longer stay dissolved because the water has become hot.

What happens when the concentration of a solution increases?

As the concentration of a solution increases, some properties of that solution change. Any property that changes when the solute concentration changes is called a colligative property.

One common colligative property is melting point. As the concentration of a solute increases, the melting point of the solution decreases. This is why salt is put on icy roads in winter – salt causes ice to melt because it decreases the melting point.

Another colligative property is boiling point. As the concentration of a solute increases, the boiling point of the solution increases. If you were to pour half a kilogram of salt into a pot, you'd find that the temperature of the water when it started to boil would be higher than 100^0 C.

A third colligative property is vapor pressure. Vapor pressure is the pressure exerted by evaporated liquid from a solution. As the concentration of a solution increases, the vapor pressure of a solution decreases.

Solubility Lab

In class, we've discussed the reasons that various compounds dissolve when placed into various solvents. In this lab, we'll make some predictions regarding the solubility of potassium chloride in isopropanol and water, followed by experiments to see if our predictions were correct.

Lab procedure:

1) Make a prediction: Will potassium chloride be more soluble in water or in isopropanol? Explain your reasons for this prediction.

2) Place 2.0 grams of potassium chloride into a 50 mL beaker with 10 mL of distilled water. Label this beaker as "Beaker A". Stir for 1 minute and write your observations below:

3) Place 2.0 grams of potassium chloride into a 50 mL beaker with 10 mL of isopropanol. Label this beaker as "Beaker B". Stir for 1 minute and write your observations below:

4) Determine which solvent is best at dissolving KCl by doing the following:

   - Very carefully pour the liquid from each solution into a waste container, making sure to leave the crystals behind in the bottom of the beaker.
   - When the liquid has been poured out, place each beaker on a hot plate and heat for three minutes to evaporate any solvent that remains.
   - Find the weight of each beaker, including the KCl residue:

Weight of Beaker A (with residue): ________________________

Weight of Beaker B (with residue): ________________________

5) Find the weight of the empty beakers by doing the following:

- Rinse out the beakers with water until the KCl has been removed.
- Dry the beakers with a paper towel, then by placing it on the hot plate for about 60 seconds.
- Find the weight of each empty beaker:

Weight of Beaker A (empty): ________________________

Weight of Beaker B (empty): ________________________

6) Using your data from steps 4 and 5, determine the amount of KCl that remained in each beaker:

Weight of KCl from Beaker A: ________________________

Weight of KCl from Beaker B: ________________________

7) Explain why your prediction from step 1 was correct or incorrect.

Postlab questions:

1) If we had used methane instead of potassium chloride, in which solvent would it have been more soluble? Explain.

2) Explain the process by which polar solvents dissolve polar solutes:

3) In step 4 of the procedure, we removed the solvent remaining in the beakers by placing them over a hot plate. However, it would have taken far less time to remove the solvent using a Bunsen burner. Why do you think we used the hot plate instead of the Bunsen burner? Explain.

4) If you were to boil a saturated solution of potassium chloride in water and a saturated solution of potassium chloride in isopropanol, which would have a greater change in boiling point over the pure solvent? Explain.

Solubility Worksheet

Answer the following questions using your knowledge of solutions and solubility:

1) Explain what polarity is.

2) Classify the following solvents as being either polar or nonpolar:

   a) water ________________________________

   b) ammonia ________________________________

   c) carbon disulfide ________________________________

3) Write "polar" in the blanks below if the compound would be more likely to dissolve in a polar solvent and "nonpolar" if it would be more likely to dissolve in a nonpolar solvent.

   a) C_6H_6 ________________________________

   b) Na_2CO_3 ________________________________

   c) KNO_3 ________________________________

   d) S_8 ________________________________

4) Adding hot water to tea bags until the liquid is medium brown makes unsugared iced tea. Ice cubes are then added to chill the beverage to a drinkable temperature. If you wanted to make sweet tea (tea with sugar), at what stage in this process would you add the sugar to ensure maximum sweetness? Explain.

5) What is a "colligative property"? Explain, using examples you learned about in class.

Chapter 21 – Acid and Base Properties

Overview

There are many ways to figure out whether something is acidic, basic, or neutral. Modern ways include pH probes, titrations, and indicators. Older methods involve studying the chemical and physical properties of the compound, such as taste, feel, color, and reactivity with metals. In this lab, students are going to guess whether ten household items are acidic, basic, or neutral using observable chemical and physical properties. They will then use more modern methods to either confirm or disprove their guesses.

Teaching about the properties of acids and bases

Many students believe they know what acids and bases are from what they've learned in middle school and horror movies. Acids are compounds that bubble when you throw things into them and are good at dissolving people. Bases are things that are really smelly, like ammonia. What more is there to know?

Students are sometimes surprised to find that many of the chemicals in their house are either acids or bases. What's more, students are surprised to find that *most* acids and bases can be safely handled and eaten.

This is not to say that the acids and bases found in chemistry classes are generally harmless. Typical acids and bases found in the classroom are very strong and very caustic. However, very few acids and bases that students are likely to encounter outside of the classroom are dangerous, and many are found in foods. After all, you don't wear goggles when cooking dinner, even if you're making something that contains vinegar!

Once your students get the idea that acids and bases are everywhere, it becomes easier to explain to them the relevance of acid and base chemistry. Once you've established this, it's easier to hit them with the complicated stuff.

Doing the "Acid and Base Properties" lab

Equipment:

- 15 watch glasses (cleaned before the lab with non-toxic soap)
- 30 *new* disposable pipets

- 250 mL of the following liquids in carefully-cleaned or new bottles: pickle brine (unknown #3), vinegar (unknown #4), cooking oil (unknown #6)
- 100 grams of each of the following solids in carefully-cleaned or new beakers: baking soda (unknown #1), grated non-toxic soap (unknown #2), baking powder (unknown #5)
- 10 grams of magnesium powder in a labeled bottle
- 100 grams of sodium bicarbonate in a labeled bottle
- 8 spatulas
- Electrical conductivity probe, if available
- As many of the following as possible: pH probe, litmus paper, universal indicator paper, cabbage extract indicator, phenolphthalein indicator.

Safety:

The substances used as unknowns are all non-toxic. However, students with chemical allergies should be discouraged from tasting any of the unknowns due to the very remote possibility that they might become ill. It's also wise to tell students to taste very small amounts of each chemical, as many have very unpleasant flavors. Students should only ingest unknowns during the first part of the lab to ensure they don't accidentially ingest the indicator solutions.

When setting up the lab, make sure that the unknown compounds never come into contact with dirty glassware, as students will be tasting them during the course of this lab. Any glassware not used for this lab should be put away to ensure that students don't accidentially place the unknowns into dirty glassware. All of the unknowns should be in a form intended for human consumption – reagent grade chemicals sometimes contain heavy metal contaminants.

The indicators in this lab are toxic and students should be warned against eating any of them. Most hazardous if phenolphthalein, which is a strong laxative. If students ingest phenolphthalein indicator solution, immediately summon medical attention. To minimize the risk of accidental indicator ingestion, make sure that all students have finished the first part of the lab before any are allowed to proceed to the second. Before starting the second section of the lab, tell your students to put on goggles and warn them against eating any anything.

Make sure your students clean the countertops when they've finished the lab to ensure that no indicators are present. When your students are finished cleaning up, make sure they all wash their hands to remove indicator residue.

If you're doing this lab with more than one class, make sure you have fresh unknowns for each class, as they are sometimes contaminated with indicator during the second part of this lab.

Room destruction factor:

The acids in this lab can be sticky, so make sure that all spills on the countertop or floor are promptly cleaned up.

How the lab works:

In the first part of the lab, students will determine if the six unknowns are acidic, basic, or neutral using simple chemical and physical tests. Students should be permitted to perform whatever tests they would like, as long as they pose no safety hazards. The most interesting tests they can perform are to mix the unknowns with the magnesium powder or sodium bicarbonate, as they will find that both tend to react well with acids.

Inevitably, some students will find that the tests will give conflicting results for some of the unknowns. For example, a compound may feel slippery (indicating a base) and taste sour (indicating an acid). When this happens, students are likely to come to you and complain that "this lab is impossible." No matter what, don't give students any hints about what they should do to resolve this problem – given time, all will realize that the cure for conflicting results is to simply perform more tests until one conclusion is more strongly supported than the other.

In the second part of the lab, students should be given time to do whatever tests they would like to determine whether the compounds are acidic, basic, or neutral. Inevitably, some students will find that a compound turns red in litmus (indicating an acid) and pink in phenolphthalein (indicating a base). Again, let the students figure it out on their own. The goal of labs is to make students think critically – by making your students interpret conflicting data, you're forcing them to think and giving them a good idea of what real science is like.

What can go wrong:

- The kids are paralyzed by conflicting data. Most students deal with conflicting data by performing more tests until the answer jumps out at them. Others stop working and start complaining that "this isn't fair". The only way of dealing with this attitude is to make it very clear that you won't be helping them with this lab, and that refusing to work won't result in a good grade. Eventually, even the most stubborn students respond, though they'll let you know they aren't happy.
- The color of the chemical compound obscures the color of the indicator in part 2. Have your students dilute the unknowns if this happens. After all, an acid diluted by a factor of ten is still an acid, and a base diluted by a factor of ten is still a base! If this still doesn't work, tell them to use their best guess based on their observations.

- All of the guesses your students made in part 1 are wrong. Let your students know that this isn't necessarily a big deal, as long as they can analyze their mistakes and figure out what they did wrong. As always, it's not the answer that matters so much as the thought that went into finding it.
- Your students can't make the lab equipment work properly. This is most often a problem when using pH probes and conductivity testers. Check to see if the equipment is broken. If it isn't, have them ask another group to explain it to them. This should teach your students to listen more carefully to your explanations in the future.

Solutions to the "Acid and Base Properties" Lab

Prelab:

By comparing the easily observable chemical and physical properties to the known general properties of acids, bases, and neutral compounds, you should be able to come to the correct answer.

Section 1:

- Unknown 1 is a base
- Unknown 2 is a base
- Unknown 3 is an acid
- Unknown 4 is an acid
- Unknown 5 is a base
- Unknown 6 is neutral

Reasons for all of these answers should be based on the easily-observable chemical and physical properties of acids, bases and neutral compounds

Section 2:

Answers are the same as for section 1, except that the reasons should be based on indicator or pH probe data.

Postlab:

Students should have a good analysis that includes a list of which unknowns they got right and good reasons why they identified some of the unknowns incorrectly.

Solutions for the Properties of Acids and Bases Worksheet

1) acid
2) neutral
3) base
4) acid
5) basic or neutral
6) acid
7) base
8) neutral
9) acid
10) neutral
11) acid
12) base
13) acid
14) neutral

Acids, Bases, and Neutral Compounds

Knowing the pH of a chemical is one of the most important pieces of data that you can collect about it, as its relative acidity or basicity has a strong effect on how the compound reacts.

Until a couple of hundred years ago, there was really no definitive way of determining whether a compound was neutral, acidic, or basic. Many of the tools we rely on today had not been invented yet, and the tools that had been invented weren't widely available. As a result, chemists had to devise other ways of figuring out whether something was acidic, basic, or neutral. These methods were based on the easily-observed properties of chemical compounds.

Below is a chart showing how people used to determine whether a compound was acidic, basic, or neutral. Keep in mind that these are general rules that may or may not be true for all acids, bases, or neutral compounds.

Property	Acid	Base	Neutral
taste	"sour" – makes your mouth pucker	"bitter" – tastes sharp	oily, sweet, or no taste at all
texture	sticky	slippery	slippery or sticky
smell	may burn nose	most have no smell	may have a strong "chemical" smell, or no smell at all
electrical conductivity	conducts electricity	conducts electricity	may or may not conduct electricity
reactivity	reacts with metals	reacts with organic compounds	varied reactivities

Nowadays, we use more modern methods to determine whether a compound is acidic or basic, including the following:

- Indicators: Chemicals that change colors in response to different pH values. Examples include litmus (red in acid, blue in base) and phenolphthalein (colorless in acid, pink in base).
- pH probes: These are electronic devices that tell you the pH of a solution. These are the most common way that chemists determine pH.
- titrations: If you know whether your compound is an acid or a base, titrations allow you to determine its strength by performing a neutralization reaction.

Acid and Base Properties Lab

In class, you've learned about some of the properties of acids, bases, and neutral compounds. In this lab, you'll be testing your knowledge of these properties to determine whether some common household items are acidic, basic, or neutral.

Prelab:

If you had no scientific equipment of any kind, how could you tell whether an unknown compound was an acid, base, or neutral? Explain.

Section 1: Acid, base, or neutral?

In this section of this lab, you will be given six substances. Your task will be to determine whether these materials are acidic, basic, or neutral based only on simple chemical tests and your five senses.

Make sure you write the evidence you used to make your determination in the space below each answer:

Unknown #1 was (circle one): *acidic* *basic* *neutral*
because (list reasons below):

Unknown #2 was (circle one): *acidic* *basic* *neutral*
because (list reasons below):

Unknown #3 was (circle one): *acidic* *basic* *neutral*
because (list reasons below):

Unknown #4 was (circle one): *acidic* *basic* *neutral*
because (list reasons below):

Unknown #5 was (circle one): *acidic* *basic* *neutral*
because (list reasons below):

Unknown #6 was (circle one): *acidic* *basic* *neutral*
because (list reasons below):

Section 2: Were you right? The experts (you) decide

In this section, it's your job to determine whether your guesses from part 1 were right. Using any available equipment, verify or disprove your earlier answers.

Unknown #1 was ___________________________ *because:*

Unknown #2 was ___________________________ *because:*

Unknown #3 was ___________________________ *because:*

Unknown #4 was ___________________________ *because:*

Unknown #5 was ___________________________ *because:*

Unknown #6 was ___________________________ *because:*

Postlab:

In the space below, analyze the accuracy of the predictions you made in section 1 in the following manner:

- For each unknown, indicate which ones you predicted correctly and which you predicted incorrectly.
- For each unknown you predicted incorrectly, explain why you think you made the mistake, based on the reasons you gave for that prediction in section 1.

If you run out of room below, feel free to use the back of this sheet.

Properties of Acids and Bases Worksheet

For questions 1-14, identify the compounds as "acidic", "basic", or "neutral" based on the properties given.

1) Compound A tastes bitter. ___________________________

2) Compound B doesn't conduct electricity. ___________________________

3) Compound C feels slippery. ___________________________

4) Compound D dissolves metals. ___________________________

5) Compound E has no odor. ___________________________

6) Compound F has a strong, acrid odor. ___________________________

7) Compound G turns litmus blue. ___________________________

8) Compound H is unreactive. ___________________________

9) Compound I feels sticky. ___________________________

10) Compound J feels greasy. ___________________________

11) Compound K turns phenolphthalein solution red.

____

12) Compound L dissolves grease. ___________________________

13) Compound M bubbles when baking soda is added.

____

14) Compound N is oily. ___________________________

Chapter 22 – Titration Lab

Overview

Titrations are used to determine the concentration of an acid (or base) by neutralizing it with a base (or acid) of known concentration and plugging this data into the equation $M_1V_1 = M_2V_2$. In this lab, students will make an indicator from red cabbage and use it to titrate an acid and base with unknown concentration.

Teaching about titrations

Students don't have too much trouble understanding the idea behind a titration. After all, if 0.25 moles of an acid are required to neutralize a base with an unknown concentration, it's pretty clear that there were 0.25 moles of base to start with. From titration data, it's easy to use this information to find the concentration of the base.

What students do have trouble with is understanding how indicators fit into the equation. Frequently, students will try to fit the quantity of indicator present into the equation $M_1V_1 = M_2V_2$, causing them to get incorrect (and sometimes nonsensical) answers. Students can be made to understand the nature of indicators by comparing them to the fans at a football game. If you couldn't see the plays in a football game, it would still be easy to see who was winning by observing the fans' reactions. However, the number of fans doesn't have any outcome on the game itself – they are simply observers that can reflect, but not influence, the game's outcome.

Aside from this, students are generally able to understand titrations easily, and quickly learn how to perform titrations in the lab and use their data to accurately determine the concentrations of acids and bases.

Doing the "Titration Lab"

Equipment:

- 1 head of red cabbage
- 6-150 mL beakers
- 6 forceps
- 6 beaker tongs
- 6 ring stands and rings
- 6 squares of wire gauze
- 6 Bunsen burners

- 100 mL each of the following solutions (all will be used in section 2): 0.5 M HCl, 0.5 M NaOH, 0.5 M NH_4OH, 0.5 M H_2SO_4, 0.5 M KOH
- 100 mL each of 0.75 M NaOH and 1 M HCl (section 3)
- 100 mL each of 0.25 M HCl and 0.1 M NaOH (section 4)
- 25 disposable pipets
- 10 watch glasses
- 1000 mL beaker for holding waste

Safety:

It is vital that students wear goggles, aprons, and closed-toe shoes because they will be working with strong acids and bases. Saturated sodium bicarbonate solution should be kept handy to neutralize acid burns, and white vinegar should be kept handy to neutralize base burns. If the eyes are burned with either, flush with water and immediately summon medical assistance. To minimize chemical accidents, acids and bases should be stored in separate locations.

Room destruction factor:

Although the student destruction factor for this lab is potentially very high due to the presence of acids and bases, classrooms are generally relatively unharmed.

How the lab works:

In section 1, students will make a red cabbage solution that they will calibrate in section 2. In sections 3 and 4, this indicator will be used to titrate an acid and a base. Each section of this lab will be discussed individually:

Section 1: Making an indicator

Section 1 is performed without difficulty simply by following the steps in the lab procedure. The only problem with this section is the stench of boiling cabbage – though the smell is initially very strong, it becomes easy to tolerate in a few minutes.

Section 2: Calibrating the indicator

In section 2, students will determine the color the indicator in acidic and basic solutions. Students sometimes forget to wash the pipet between solutions, causing contamination and weakening of the acid and base solutions. A reminder that the pipets should be cleaned before and after each use should solve this problem. Also make sure that students collect all waste in a 100 mL beaker for disposal.

Sections 3 and 4: Titration of a base/acid

The unknown base solution in section 3 is 0.75 M NaOH and the unknown acid from section 4 is 0.25 M HCl. Make sure each unknown is clearly labeled as "Unknown for section ___" so students will know which to use in each section.

In both sections 3 and 4, the unit of volume used is "drops" instead of "milliliters". Remind students that it doesn't matter what units of volume are plugged into the equation $M_1V_1 = M_2V_2$, as long as the units are the same for both the acid and base. To ensure that the size of each drop is the same, it is very important that students use the *same pipet* for both the acid and the base – otherwise, the relationship between the two is impossible to determine. Students should also be reminded to carefully wash the pipet after each use to avoid contamination.

All waste should be collected in a 1000 mL beaker for disposal.

<u>What can go wrong:</u>

- No matter how much titrant the student uses, the color of the solution never changes. This occurs because students have either forgotten to add indicator to the solution or are titrating an acid with another acid (or base with base). Have your students show you exactly what they did and the problem will probably not recur.
- "How do we know when the color has changed?" It can be difficult to see a color change against the black background of the typical chemistry lab table. To avoid this problem, have your students perform the titration over a white sheet of paper.

<u>Clean up</u>:

Cabbage extract may be poured down the sink. I recommend flushing the sinks with extra water to keep the odor of boiled cabbage out of your room.

All acid and base waste should be poured into a 1000 mL beaker. To neutralize this waste, add about 20 grams of sodium bicarbonate to the solution and stir. When dissolved, add 0.5 M acetic acid to the beaker until the bubbling stops. This will correspond to a slightly acidic solution that is safe put down the sink.

Solutions to the Titration Lab

<u>Prelab:</u> 0.58 M

Section 1:

Step 1: The cabbage leaf will appear purple at the ends and lighter toward the base. Some students may notice white veins in the flesh of the leaf.

Step 3: The leaf will be bleached, ranging in color from light purple to white.

1) The leaf looks as if the color has been completely removed.
2) Flower petals can also be used as indicators.

Section 2:

Procedure: The table should list each acid and base, and the color that each solution turned when the indicator was added.

1) Acidic solutions appear purple, while basic solutions are green/yellow.
2) You would have to repeat this experiment any time you discovered a new indicator in order to determine whether its color in acid and base.

Section 3:

1) 0.75 M
2) The same dropper needs to be used to ensure that the drops are of uniform volume. If they weren't, you wouldn't be able to use "drops" as a unit of volume in the titration calculation.

Section 4:

1) 0.25 M
2) The only difference is that the unknown was an acid rather than a base. The underlying concept that titrations involve the neutralization of one compound with another is the same.

Postlab:

You could not do a titration if both solutions had an unknown concentration because there would be two unknowns in the equation $M_1V_1 = M_2V_2$.

Solutions to the "Titration Worksheet"

1) 0.1 M
2) 2000 mL
3) Because the products of a titration are a salt and water, the pH of the solution should be exactly 7.
4) The amount of indicator in a titration doesn't matter because the indicator is a spectator, not a part of the titration.
5) Added water won't change the amount of acid or base present. As a result, there will be no effect on the titration.

Titration Handout

Titrations are neutralization reactions used to determine the concentration of an acid or base. For example, if you had an acid with an unknown concentration, you would add base to it until all of the acid had been neutralized. At the very point that all of the acid was gone, the amount of base that was added would be exactly equal to the original quanity of acid. Likewise, if you wanted to find the concentration of a base, you could neutralize it with an acid of known concentration.

The equation for titration calculations is:

$$\mathbf{M_1V_1 = M_2V_2}$$

M_1 = molarity of acid	*M_2 = molarity of base*
V_1 = volume of acid	*V_2 = volume of base*

Note: Any unit can be used for volume, as long as the same unit is used for the volume of both acid and base.

In all titrations, an indicator is needed to determine when the acid or base has been neutralized. Without an indicator, it would be impossible to determine when the titration should stop. The exact amount of indicator used is unimportant, as it isn't directly involved in the neutralization reaction.

Indicators do, however, provide the main source of error when performing titrations. When we do a titration, we assume that the indicator changes colors when the pH is exactly 7.00. However, indicators typically change colors at a slightly different pH, causing us to stop our titration at the incorrect time. The point at which the indicator changes color is called the "endpoint", because the color change tells us when to end the titration. The point at which the pH is exactly 7.00 is the "equivalence point" because it takes place when the amount of acid and base are exactly the same. For most titrations, there is only a very small difference in pH between the equivalence point and endpoint, making this a minor source of error.

Titration Lab

As you know, titrations are used to determine the concentration of an acid (or base) by neutralizing it with a base (or acid) with a known concentration. At the endpoint of the titration, you can determine the concentration of the acid or base with the equation $M_1V_1 = M_2V_2$.

In this lab, we will look at several acid-base titrations and how they work.

Prelab:

350 mL of 0.75 M HCl is required to titrate 450 mL of a NaOH solution with unknown strength. Using this data, what was the concentration of the NaOH?

Section 1: Making an indicator

Acid-base indicators are used in titrations to let you know when the solution has been neutralized. Although there are many man-made indicators available, there are also many that are naturally-occurring. We will be working with one of these natural indicators today.

Procedure:

1) Tear a few leaves of red cabbage apart into small chunks with your fingers. Examine one of these leaves and write your observations below:

2) Fill a 150 mL beaker halfway with cabbage leaves and add water until the beaker is three-quarters full. Using a ring stand and Bunsen burner, heat the cabbage mixture to boiling for about ten minutes.

3) You should notice that the water turns purple as the cabbage boils. When the water is very purple, remove the beaker from the heat and pull a cabbage leaf from the mixture with forceps. Examine the leaf and write a description of what it looks like below:

4) The purple juice that you have made is cabbage extract indicator. We will be using this indicator for the remainder of the lab.

Questions for section 1:

1) Did the cabbage leaf look different after it had been boiled? Explain.

2) What other sorts of naturally-occurring compounds do you think could be used as indicators? Explain.

Section 2: What does the cabbage indicator indicate?

In section 1, you made an acid-base indicator from red cabbage. The only problem with it is that you don't know what colors it is in acidic or basic solutions. Right now, you could put the indicator in a solution and it would turn some color – unfortunately, you currently have no way of knowing what that color means. To fix this problem, we will expose our cabbage indicator to common acids and bases to determine its color under varying conditions of acidity.

Procedure:

In a watch glass, combine three drops of cabbage extract with small amounts of each of the following: 0.5 M HCl, 0.5 M NaOH, 0.5 M NH_4OH, 0.5 M H_2SO_4, and 0.5 M KOH. Make sure to rinse your pipet each time you use a new solution. Write your data in a table below:

Questions for section 2:

1) In the spaces below, identify the color of the indicator in acidic and basic solutions.

Color of the indicator in acidic solutions: ________________________

Color of the indicator is basic solutions: ________________________

2) Do you think you would need to repeat this experiment each time you discover a new indicator? Why or why not?

Section 3: Titrating a base with an acid of known concentration

In this section you will determine the concentration of a base with an unknown concentration by performing a simple titration. You may be familiar with this sort of titration from simple titrations performed to find the pH of swimming pools.

Important note for this section: Make sure you rinse the pipet with distilled water before and after every use. Otherwise, it will become contaminated.

Procedure:

1) Place ten drops of “unknown base” into a watch glass with a pipet.

2) Add two drops of cabbage indicator to the watch glass.

3) With the same pipet you used for step 1, slowly add 0.1 M hydrochloric acid to the “unknown base” until the color of the indicator changes. When the color changes, you have reached the endpoint and are finished.

Questions:

1) Calculate the molarity of the base, using $M_1V_1 = M_2V_2$. Show your work.

2) Why did you need to use the same pipet for steps one and three?

Section 4: Titrating an acid with a base of known concentration

In this section, you will find the concentration of an acid by titrating it with a base of known concentration. The procedure you will use will be the same as that for section 3, except that you will be titrating an “unknown acid” with 0.1 M KOH. Again, make sure you rinse your pipet after each use to avoid contamination.

Procedure:

1) Place ten drops of “unknown acid” solution into a watch glass.

2) Add two drops of cabbage indicator to the watch glass.

3) Slowly add 0.1 M KOH with the same dropper from step 1 until the indicator changes color. When the color changes, you are finished.

Questions:

1) Calculate the molarity of the acid, using $M_1V_1 = M_2V_2$. Show your work.

2) Was there anything different about doing this titration than for the titration of a base in section 3? Explain.

Postlab section:

If you had an acid with an unknown concentration, would it be possible to titrate it with a base of unknown concentration? Explain.

<u>Titration Worksheet</u>

Use your knowledge of titrations to answer the following questions:

1) If it takes 0.5 L of 0.35 m HCl to neutralize 1.75 L of a base with unknown concentration, what is the concentration of the base?

2) How many milliliters of 0.50 M NaOH will be needed to neutralize 400 mL of 2.50 M H_2SO_4.

3) Titrations are also known as "neutralization reactions". Explain what this means in terms of the products formed in a titration.

4) If the instructions for a titration indicate that I should use 0.05 mL of an indicator in a titration and I accidentally use 0.5 mL instead, will I have to do the titration over again? Explain.

5) During a titration, I accidentally added 100 mL of water to the acid solution I was trying to titrate. Will this cause an error in my titration, or can I continue without any trouble? Explain.

Chapter 23 – Kinetics Activity

Overview

Kinetics is the study of the rates of chemical reactions. Reaction rates are a function of reagent temperature, concentration, and surface area – as each of these increase, the rate of reaction increases. In this activity, students will simulate the relationships of temperature and concentration to reaction rate by impersonating molecules under various conditions.

Teaching About Kinetics

Students usually like learning about kinetics, because they can imagine what molecules do on a very small scale. Though molecules are very small, it's easy to imagine them as small particles that collide with one another to undergo chemical reactions. The use of molecular models can be particularly useful in visually demonstrating the interplay of molecules in a chemical reaction.

To teach students about how temperature and concentration affect the rate of a chemical reaction, it's first important to put chemical reactions in terms that students can easily understand. A good model for chemical reactions is the bumper car ride at an amusement park. Most collisions between bumper cars result in an elastic collision, where the two bumper cars bounce away from one another unchanged. However, once in a long while, two bumper cars will collide with exactly the right energy and orientation to disable one of them. This is analagous to a chemical reaction: Many collisions between reagent molecules don't result in chemical reaction because the reagent molecules weren't oriented correctly or because they didn't have enough energy. Once in a while, everything lines up perfectly and the reaction forms product.

There are things we can do to speed the rate of a chemical reaction, much as there are things we can do to maximize the destruction of a bumper car ride. When children ride bumper cars, they learn very quickly that it's far more satisfying to allow some speed to build up before crashing into their little brother. Not only do fast bumper cars provide a lot of jarring energy to little brothers, but they maximize the possibility that the little brother will be stuck in an inert bumper car for the rest of the ride.

For molecules, we can maximize the chances of chemical reaction by increasing their speeds. To do this, we increase the temperature of the reagents, which causes the molecules to collide more quickly and violently. Generally, reaction rates double for every ten degree increase in temperature.

Another good way to destroy bumper cars is to pack a very large number of cars into a very small area. When this happens, the bumper cars will tend to collide more frequently than if they were more spread out. Though the bumper cars won't be traveling quickly, the sheer number of collisions will cause a considerable number of them to be disabled.

The same phenomena takes place with reagent molecules. If a very large number of reagent molecules are present in a very small place, they will tend to hit each other more frequently. Because a certain arrangement of reagent molecules is needed for a chemical reaction to proceed, a larger number of random collisions will also lead to a larger number of collisions where the reagents have the correct orientation to undergo a reaction.

Throughout the history of bumper cars, younger brothers have learned that they can minimize the frequency of collisions by driving alongside their parents. The idea is simple: If you're in a group, half of the bumper cars will hit the other person rather than you. As a result, it's only half as likely that your bumper car will be disabled by an attack by an older sibling.

The same is true with reagent molecules. If reagents are thrown into a reaction mixture as a big lump, the number of purposeful collisions will be decreased. The molecules at the surface of the solid will be hit by other reagent molecules and react quickly, but the reagent molecules in the middle of the solid won't undergo any collisions at all. As a result, it's possible to speed the rate of a chemical reaction by grinding each reagent into a fine powder, rather than leave it as a big clump.

The bumper car analogy can be extended almost indefinitely, even to explain higher order chemical reactions. However, after a point, the analogy gets a little bit silly. For example, in a fourth order reaction, four bumper cars will need to collide simultaneously in such a way…

Doing the Kinetics Activity

Equipment:

- As many blindfolds as there are students
- 4 poles, between 1 and 2 meters in height (wooden dowels can be purchased at craft or home-supply stores). One end of the pole should be covered with a rag or tennis ball.
- 45 meters of twine or string
- 1 stopwatch

Safety:

In this activity, your students will wander outside in a 10 x 10 meter square while blindfolded. Make sure you choose an area free of holes or other obstructions that may cause students to trip. Each of the poles should have something soft on the exposed end to ensure that students aren't impaled if they trip.

Students should be encouraged to hold their hands at waist level while performing this lab to ensure that they don't walk headlong into anybody. It is a very good idea to keep a close eye on this activity to ensure that the boys keep their hands to themselves during this activity. If you believe that harassment will be a problem, you may want to have the girls take the data for section 1 and the boys take the data for section 2.

Room destruction factor:

It's unlikely your room will be damaged, as this activity takes place outside.

How the lab works:

In this activity, your students will pretend that they are molecules within a beaker. Each collision they undergo with another "molecule" corresponds to a chemical reaction. By manipulating the speed at which the students walk and the number of students present at any given time, the rate of reaction will be observed to change. This activity does an excellent job of illustrating why temperature and concentration cause changes in reaction rate. Students also think it's a lot of fun.

In section 1, students will be doing three experiments where they simulate the motion of molecules at different temperatures. Because molecules in solution move in random directions, students will be blindfolded and told that whenever they encounter any object they should turn around and wander away in a random direction. During each trial they should count the number of times they bump into another person, or another person bumps into them. Each trial will last for 90 seconds.

To ensure that students really walk more quickly in each trial, say the word "step" once a second for the low temperature trial, twice a second for the medium temperature trial, and four times a second for the high temperature trial. To keep students from running around the square at very high speed for the final trial, have them take very ten centimeter steps during each trial.

After the data has been collected and filled out in the first section, your students will make a graph showing the dependence of the number of collisions on the temperature. I've found that this graph generally has an upwardly sloping curve, corresponding to an exponential increase in collision rate, though you may have different results.

In section 2, students will simulate the effect of concentration on the number of collisions. One-quarter of your students will be involved in the low concentration trial, one-half will take place in the medium concentration trial, and all will take place in the high temperature trial. Each trial will last exactly 90 seconds. To ensure that the "molecules" travel the same distance in each 90 second period, instruct them to take a step every half second.

When the data has been collected for section 2, it is again graphed to show the dependence of reaction rate on concentration. These graphs usually show a high exponential increase in the reaction rate, corresponding well to what one might expect from a concentration increase in a second order reaction.

What can go wrong:

The only thing that I've ever seen go wrong with this lab is that the students become talkative and rowdy. Classes with chronic behavior problems should probably not perform this lab.

Solutions to the Kinetics Activity

Prelab:

1) The rate of a reaction increases as temperature increases because the particles have more energy with which to make the reaction take place.
2) The rate of a reaction increases when the concentration increases because the particles bump into one another more often, causing more opportunities for the reaction to take place.

Section 1:

Trial data: Data should increase from trial 1 through 3.

Graph #1: This graph should show an exponential increase in the number of collisions as the temperature increases. The temperature should be on the x-axis and number of collisions on the y-axis.

Questions:

1) A the temperature increases, the number of collisions increases.
2) If the graphs showed a straight line, it demonstrates a direct proportion between temperature and reaction rate. If the graphs show a curve, it shows that the reaction rate increases exponentially as the temperature increases.

Section 2:

Trial data: Data should increase from trials 1 through 3.

Graph #2: This graph should show an exponential increase in the number of collisions as the concentration increases. The concentration should be on the x-axis and the number of collisions on the y-axis.

Questions:

1) As the concentration increases, the reaction rate increases.
2) Because the graphs show an exponential increase in reaction rate as the concentration increases, the data shows that increasing the temperature has a very large effect on the rate of a chemical reaction.

Post-lab questions:

1) Concentration change, because the curve rises more dramatically.
2) It would work equally well for gases, though "partial pressure" would be substituted for "concentration". Gases and liquids behave in very similar manners because molecules move randomly in both.
3) Students should believe this is a good simulation, as it does a good job of recreating the data that one would expect from a real solution.

Solutions to the Kinetics Worksheet

1) Increasing temperature means that the particles will bash into each other more frequently and with more energy. These frequent, energetic collisions will cuse the reaction rate to increase as temperature increases.
2) Increasing concentration means that the particles will bash into each other more frequently. A higher number of collisions increases the reaction rate.
3) Lower temperature results in decreased reaction rates. Because rotting and bacterial growth are chemical processes, it's best to have a low temperature.
4) Temperature dependence: The more energetic the collision, the more likely that the activation energy will be exceeded.
   Concentration dependence: The more collisions, the more likely that some of them will have energies exceeding the activation energy.
5) The rate of reaction will increase because more of the reagent molecules will be at the surface, ready to react. Increasing the surface area of a solid increases its reaction rate.

Factors Affecting Reaction Rates

Kinetics is the study of things that affect the rates of chemical reactions. Whenever scientists are trying to determine the shelf life of a food or medicine, they do kinetic studies in which they measure the rate of decomposition under various conditions. Likewise, scientists do similar studies whenever they need information about the rate of any chemical reaction.

There are three main things that affect the rate of a chemical reaction. By changing any of these three things you can change the rate of a chemical reaction:

1) **Temperature**

Increasing the temperature of a mixture of reagents causes the molecules to move more quickly. Because they move more quickly, they undergo more collisions with greater energies. Because the reagent molecules undergo these frequent, high energy collisions, they are likely to overcome the activation energy for the reaction.

2) **Concentration**

Concentration refers to the number of particles of solute per unit of volume, such as "moles per liter".

When there are a large number of molecules in a small area, they are likely to bump into each other frequently. Because molecules need to be next to each other to react, increasing the concentration of a solution increases the rate of reaction.

3) **Surface area**

If you have one kilogram of a solid chemical, it will react far more quickly if you grind it into a powder rather than use it as a giant crystal. In a powder, more of the molecules will be on the surface of the solid. Because molecules need to be at the surface to collide with other reagent molecules, powders tend to react more quickly than large solids. This is why the sugar for your iced tea comes in small grains rather than large clumps.

<u>Kinetics Activity</u>

In this activity we will investigate what happens to the rate of a chemical reaction when the temperature and concentration of the solution is increased. However, there's a twist: Instead of simply measuring what happens in a chemical solution, we're going to measure what happens when we increase *your* "temperature" and "concentration".

Prelab:

1) What do you think happens to the rate of a chemical reaction when the temperature is increased? Why?

2) What do you think happens to the rate of a chemical reaction when the concentration is increased? Why?

Section 1: Increasing the temperature of a chemical reaction

When the temperature of a solution is increased, the particles in the solution move more quickly. After all, if temperature is a measurement of kinetic energy and kinetic energy depends on speed, then high temperature corresponds to high speed. In this section you will see what happens to the rate of a chemical reaction when the temperature is increased.

<u>Trial 1</u>: Low temperature experiment:

Total number of collisions you experienced: ________________

Total number of collisions everybody experienced: ________________

<u>Trial 2</u>: Medium temperature experiment (temperature doubled from Trial 1):

Total number of collisions you experienced: ________________

Total number of collisions everybody experienced: ________________

Trial 3: High temperature (temperature quadrupled from Trial 1):

Total number of collisions you experienced: _______________

Total number of collisions everybody experienced: _______________

Graph #1: Temperature vs. total collisions experienced by the whole class

Use the chart below to graph the relationship between the temperature and number of collisions experienced by the entire class.

Questions:

1) What is the effect of temperature on the number of collisions? Explain, using your data.

2) Was your graph a straight line or a curve? What does this mean about the relationship between temperature change and rate of reaction?

Section 2: Increasing the concentration of a chemical reaction

When the concentration of a solution is increased, there are more particles present in every part of the solution. This increase in particles will affect the frequency of intermolecular collisions. For section 2, you will determine what happens to the rate of a chemical reaction when the concentration is increased.

<u>Trial 1</u>: Low concentration experiment:

Total number of collisions you experienced: ________________

Total number of collisions everybody experienced: ________________

<u>Trial 2</u>: Medium concentration experiment (concentration doubled from Trial 1):

Total number of collisions you experienced: ________________

Total number of collisions everybody experienced: ________________

<u>Trial 3</u>: High concentration (concentration quadrupled from Trial 1):

Total number of collisions you experienced: ________________

Total number of collisions everybody experienced: ________________

Graph #2: Concentration vs. total collisions experienced by the class

Use the chart below to graph the relationship between the concentration and number of collisions experienced by the whole class.

Questions:

1) What is the effect of concentration on the number of collisions? Explain, using your data.

2) Was your graph a straight line or curve? What does this mean about the relationship between concentration and reaction rate?

Postlab questions:

1) Which causes a more dramatic change in the reaction rate, temperature change or concentration change? Explain, using your graphs.

2) In this lab, we simulated a liquid solution. Would this simulation work equally well for a mixture of gases? Explain why or why not.

3) Do you think that this simulation was accurate? If your answer is “yes”, explain why you believe this is a good simulation. If your answer is “no”, explain how you believe this exercise could have been made better.

Kinetics Worksheet

1) Explain why the rate of a chemical reaction increases as the temperature of the reagents increases.

2) Explain why the rate of a chemical reaction increases as the concentration of the reagents increases.

3) For what reason do we put leftover food in the refrigerator? Why can't we just leave it out on the countertop?

4) Activation energy is the minimum amount of energy needed for a chemical reaction to occur. Explain how activation energy fits into the dependence of reaction rate on temperature and reagent concentration.

5) Chalk reacts with hydrochloric acid when the two are mixed? Will pounding the chalk into a powder increase, decrease, or do nothing to the rate of this reaction? Explain, giving specific reasons for your answer.

Chapter 24 – Radioactive Decay Lab

Overview

Atomic nuclei undergo a variety of nuclear reactions, including alpha emission (in which a helium nucleus is emitted), beta emission (in which an electron is emitted), and gamma emission (in which a gamma ray is given off). These reactions are caused by instabilities in the nucleus of an atom. The time it takes for 50% of the nuclei in a sample to undergo radioactive decay is called the “half-life”. Half-life can be used for many things, among which carbon dating is probably the most well-known. In this lab, students will determine the theoretical vs. experimental half-lifes of the “heads” state of pennies during random flips.

Teaching About Radioactive Decay

Nowadays, students aren’t as familiar with concepts like “radiation” and “half-life” as students of twenty years ago. Back in the bad old days of the cold war, most people had heard of “half-lives” and knew that they had something to do with how long radioactive fallout would make the earth an unpleasant place to live. Last year I asked my students if they’d heard of the concept of half-life and was told that it was the name of a video game they liked to play. Times have changed!

Though the easing of tensions between the nuclear superpowers has made the world a better place, it has made our jobs as chemistry teachers a bit harder, as we have to teach all nuclear concepts from scratch. One can only hope that this is a permanent handicap to our teaching.

<u>Types of radioactive decay:</u>

Teaching students about the different types of radioactive decay is easy. After all, if your students can do simple addition and subtraction, they can easily figure out that losing ^{4}He from ^{235}U results in the formation of some element with an atomic mass of 231. With very little training, students can write the extra “2” below the helium and “92” below uranium to find that the identity of the element is thorium, element 90.

Some teachers discuss many types of radioactive decay, including electron capture, positron emission, etc. I prefer to stick with good old alpha, beta, and gamma radiation. They’re simple to remember, and are most likely the only types of radiation that students will need to know unless they become nuclear chemists. Likewise, I would avoid talking about quarks and other sub-subatomic particles to your students, as it’s unlikely that they will be able to make any sense out of this topic without advanced calculus and quantum mechanics.

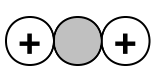
Radioactive decay:

Though students are easily able to learn that atoms decay, they don't always understand why very large atoms tend to fall apart in the first place.

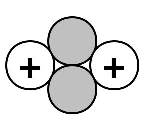
To explain why decay occurs, place the positive ends of two magnets together on a table and ask your students what will happen when you let them go. They will accurately predict that the magnets will fly apart from one another because like charges repel each other. Ask them what would happen if you placed a very tiny piece of tape between them. Some students will continue to believe that the magnets will fly apart, while others will indicate that the strength of the tape should be enough to hold the magnets together. This model of an atomic nucleus is a pretty good one – one nuclear force tries to keep all atomic particles together, while their charges cause them to repel.

Clearly, it will be difficult to hold a bunch of positive charges together with tiny pieces of tape alone. That's where the neutrons come in. Pull the two magnets apart and place a marble between them. The marble will help to separate the positive charges – this additional separation weakens the repulsion enough that the tape will definitely be strong enough to hold the nucleus together. This is what happens in an atom.

An added benefit of this analogy is that it explains why isotopes occur. As your students should remember, isotopes are forms of the same element that have different numbers of neutrons. In other words, there are different numbers of spacers being used to hold the protons together in each isotope. We can see how this works by looking at the diagram below:



In this diagram, one neutron, represented by the gray circle, serves to keep the two protons separated from one another.



In this diagram, two neutrons keep the protons separated from one another.

In both of the above cases, the two positive charges of the two protons are separated from one another, stabilizing the nucleus. Because the nucleus can be stabilized by two combinations of neutrons, there are at least two isotopes of this element.

This analogy also does a good job of explaining why no elements with an atomic number greater than 83 form stable isotopes. One can imagine that if you were to hold 84 magnets together, no combination of neutrons would allow the tape to be strong enough to keep them together, as there would just be too much positive charge in one place. Students don't have a hard time understanding this concept – in my experience, many are surprised that any nuclei are stable at all!

Half-life

How long does it take for any particular atom to decay? Nobody knows. How long does it take for half a radioactive sample to decay? One half-life. How can you find half-life? You need to take measurements – there's no way to predict half-life based on the periodic table.

Students frequently get hung up on the idea that half-life is a statistical value and not an absolute value. Expect to hear questions like, "If you had a radioactive sample with a long half-life, is there a chance that every atom will decay simultaneously?" The answer is yes; all of those atoms *could* decay at once. Will they? Probably not, because statistically, very few atoms decay in a short period of time for an isotope with a long half-life. Many students have problems with the abstract thinking required to understand this.

Here's an analogy that students might understand:

> *I've noticed that 10% of my students need to use the restroom during a typical class period. Although it's possible that all of the students might need to use the restroom in one day, it's never happened yet.*
>
> *Johnny sits in the third row of the class. What time will Johnny get up to use the restroom today? It's impossible to predict if he'll need to use the restroom at all, much less when he'll need to use it. For any student in my classroom, you cannot predict if they'll be among the 10% that need to use the restroom. Statistics tell us what is likely to happen, but cannot tell us specifics.*

Half-lives are useful for determining how long ago something died. This is because the concentration of atmospheric ^{14}C is kept steady by natural processes, and this ^{14}C is incorporated into the tissues of living things. When something dies, the relative amount of ^{14}C will decrease in its tissues as it undergoes radioactive decay to make other elements, while stable ^{12}C remains in tissues forever. From the proportion of ^{14}C to ^{12}C left in a previously-alive sample, it can be determined how long ago that organism died.

Doing the Radioactive Decay Lab

Equipment:

- 400 pennies
- 6 closed containers, such as cigar boxes or Tupperware containers

Safety:

There is nothing hazardous about this activity.

Room destruction factor:

It is impossible to damage your room with this activity.

How the lab works:

Students will place between 30 and 60 pennies heads-up in a container. They will then shake the container and remove any pennies that have landed on tails. The number of remaining heads will be counted and this number placed in a table. After all pennies have flipped from heads to tails, this data will be graphed and used to determine the half-life of a penny, in flips.

This lab should be simple – the half-life of a penny is one flip, because there is a 50% probability that the penny will land on tails on any given flip. Unfortunately, our sample is small enough that we won't have exactly 50% of the pennies flipping over each time we shake the box. There is a good chance that for some flips, ony 30% will flip over, and for others, 70%. The goal here is to get students to realize that while half-lives are given as constant numbers, it's just a statistical value and may not always be 100% accurate for very small samples. However, despite the small sample size of this lab, students should get an answer that's very close to 50% with their data.

Half-life can be detemined by examining the graphs that each student makes. Strictly speaking, the halflife should be determined by figuring out how long it takes for ½ of the pennies to flip from heads to tails. For example, if the students start with 32 pennies, the first experimental half-life will occur when the graph indicates that there are only 16 pennies left. The problem with this method is that unless the data is perfect, the number of flips for each half-life will be a decimal such as 0.8 or 1.3. Clearly, neither of those numbers corresponds to an observable value. There's nothing wrong with this – simply have the students find the interval from 32 to 16 pennies, from 16 to 8 pennies, and so on until all of the pennies have flipped over. The values for each half-life should then be added together and averaged to find the average half-life of “heads” in the pennies. This calculation will be performed during this lab in the “procedure” section. The answer to this calculation will still not be a whole number of flips, so it will be important to point out that the error in this lab is due to the very small sample size.

What can go wrong:

The main error seen with this lab is that a group finds the average half-life of heads to be two flips. This is probably an error with how they’ve determined half-life rather than a data error. However, because the number of pennies used in this lab is so small, there is a chance that this corresponds to their real data. If this happens, explain the significance of small sample size in statistical analysis – namely, the smaller the sample, the lower the quality of data.

Solutions for the Radioactive Decay Lab

Prelab:

The half-life is one flip, because there’s a 50/50 chance that the coin will turn over in any flip.

Procedure:

1) A number between 30 and 60 should be given.
3) A number with exactly half the value of the number in #1 should be given.
4) This table should be filled out. The values in the “Pennies that were heads” column should decrease with each entry. There will most likely be many blanks at the end of the table because the pennies all flipped over.
5) This graph should be filled out with the vertical axis labeled “pennies that were heads” and the horizontal axis labeled “number of flips”. The graph should be labled clearly and data should be plotted so it takes up the maximum area of the graph. All points should be connected with a curve.
6) The “number of flips” should be approximately 1. If the numbers listed are all whole numbers, either the graph has been read incorrectly or the data has been falsified.
7) The values in the table should be averaged – the answer should be ~1.
8) The values for all groups should be averaged – the answer should be ~1.

Questions:

1) It will most likely not be the same, because when you use a small statistical sample it’s harder to get good data.
2) The class average should (but may not) be better than individual averages, because there is a larger sample size and errors from each group tend to cancel each other out.
3) The observed half-life deviates more from the norm when there are fewer pennies.

4) It is impossible to say, as half-life is a statistical value, useful for populations of events and not individual events.

Solutions for the Nuclear Chemistry Worksheet

1) Atoms may be less stable because there may be so many protons in the nucleus that it cannot stay together. There may also be too many or too few neutrons, causing the nucleus to become unstable.
2) Alpha decay: Helium nucleus
   Beta decay: electron
   Gamma decay: gamma ray
3) Half-life is the amount of time required for 50% of the atoms in a radioactive sample to decay.
4) 15.63 grams.

5a) $${}^{135}_{54}\text{Xe} \rightarrow {}^{131}_{52}\text{Te} + {}^{4}_{2}\text{He}$$

5b) $${}^{135}_{54}\text{Xe} \rightarrow {}^{135}_{55}\text{Fr} + {}^{0}_{-1}\text{e}$$

5c) $${}^{135}_{54}\text{Xe}^{*} \rightarrow {}^{135}_{54}\text{Xe} + {}^{0}_{0}\gamma$$

6) No. To become radioactive, one would need to ingest nuclear material.
7) 20 minutes.

<u>Nuclear Decay Handout</u>

Nuclear decay is a process in which the nucleus of an atom breaks apart, giving off smaller particles. This occurs when the nucleus becomes unstable. Nuclei become unstable when there are too many or too few neutrons in the nucleus, or because the nucleus has extra energy it needs to get rid of. Whatever the cause, there are three main types of radioactive decay:

1) <u>Alpha decay</u>: This is when a particle containing two neutrons and two protons is emitted from a nucleus. This particle is given the same symbol as helium because the nucleus of a helium atom also has two protons. And example of the emission of an alpha particle (α) is shown below:

$$^{175}_{80}\mathrm{Hg} \rightarrow {}^{171}_{78}\mathrm{Pt} + {}^{4}_{2}\mathrm{He}$$

2) <u>Beta decay</u>: This is when a neutron in the nucleus is converted into a proton and an electron. The electron is called a beta particle (β) and is emitted from the nucleus as shown below:

$$^{220}_{85}\mathrm{At} \rightarrow {}^{220}_{86}\mathrm{Rn} + {}^{0}_{-1}\mathrm{e}$$

3) <u>Gamma decay</u>: This is when energy is given off from the atom in the form of a gamma ray (g). The nucleus of the atom keeps the same number of protons and neutrons, but is more stable because it has given off its excess energy. An example of gamma emission is shown below:

$$^{233}_{93}\mathrm{Np}^{*} \rightarrow {}^{233}_{93}\mathrm{Np} + {}^{0}_{0}\gamma$$

The * over the Np means that Np has extra energy it needs to get rid of.

Half-lives:

A handy thing to know about a radioactive element is the amount of time it will take for it to decay into different elements. Unfortunately, you cannot tell how long it will take for any radioactive atom to decay because there are no internal clocks that determine the instant it will fall apart. However, we can determine the average length of time it will take for 50% of the atoms in a sample of a radioactive element to decay – this length of time is called the "half-life".

Some elements have half-lives greater than 10^{10} years, while others have half-lives around 10^{-20} seconds. More typically, half-lives for most elements range from 0.1 second to 10^8 years.

One way that the age of some artifacts or fossils can be determined is through radiocarbon dating. In this method, the amount of ^{14}C left in a sample from when it was alive can be used to determine the artifact's age. Because ^{14}C has a half-life of 5715 years, we can figure that if there is only half as much ^{14}C as there should be in a sample, it's 5715 years old. Likewise, if there was only 1/8 as much ^{14}C as there ought to be, this would suggest that the artifact was three half-lives, or (5713 x 3) = 17,145 years old.

The amount of a radioactive sample left over after a given period of time can be calculated using the following equation:

$$\textbf{amount of sample remaining} = \frac{\textbf{original amount of sample}}{\mathbf{2}^{\frac{\text{number of years elapsed}}{\text{half-life of sample}}}}$$

Radioactive Decay Lab

When you have a sample of radioactive atoms, you can expect that some will decay in a short time while others take much longer to decay. However, if you look at any single atom in the sample, it's impossible to guess when it will decay.

Fortunately, statistics allows us to predict what will happen to the sample as a whole. Although we can never tell when any atom in a sample will decay, we can state the period of time it takes for 50% of the atoms in a sample to decay. This time is called the "half-life". As a result, if we have a radioactive sample that weighs 1000 grams, we know that after one half-life we'll have 500 grams left because 50% of the atoms in this sample will decay. For example, it takes 19.3 seconds for 50% of the atoms ina sample of ^{10}C to decay and 1.4×10^{17} years for 50% of the atoms in a sample of ^{50}V to decay. Another way of expressing these statements is that ^{10}C has a half-life of 19.3 seconds, while ^{50}V has a half-life of 1.4×10^{17} years.

In this lab, you'll be investigating something more familiar than radioactive isotopes: The flipping of a coin. As with atoms, the term "half-life" refers to the time needed for 50% of the sample to do something.

Prelab:

Let's imagine that coins are radioactive atoms. "Heads" means that the atoms have not decayed and "tails" means that they have. Keeping this in mind, what is the half-life of "heads", in number of flips? Put another way, if there are two pennies that are heads, how many flips will it take for one head to remain?

Procedure:

1) Pick a number at random between thirty and sixty: ___________________

2) The number from step 1 corresponds to the number of pennies that you should get from your instructor. When you have obtained your pennies, you should put them in a box **heads side up** at your lab station.

3) You will be shaking the box to randomize the orientation of the pennies. Before you do this, make a guess about how many of the pennies will still be "heads" when you are finished shaking the box for the first time:

I predict there will be______ pennies left heads up after one shake.

4) Shake the box and remove all the pennies that came up as tails. Keep doing this until there are no pennies left as heads, removing the pennies that are tails after each shake. Record your data on the table below:

Turns	Pennies that were heads after each turn
0	
1	
2	
3	
4	
5	
6	
7	
8	
9	
10	
11	
12	
13	
14	

In the "Pennies that were heads after each turn" column, write the number of heads after the number of turns in the column on the left. In the space that corresponds to "0 turns", write the number of pennies you started with, since all of them started out as heads.

5) Graph your data on the table below. Label both the x- and y-axis of your graph and draw a best-fit line for your data. *Do not connect the dots!*

6) Using your experimental data, determine the value for each half-life in flips. For example, if you started with 80 pennies, the first experimental value for half-life would be the number of flips it took for 40 pennies to flip over, leaving 40 heads. The second experimental value for half-life would be the number of flips it took for only 20 heads to remain. Note: This may not be a whole number – estimate the number of flips to the nearest tenth.

Half-life	**Number of flips**

7) Using the data from the table above, find your overall average half-life for heads in flips. Show your work.

Our average half-life for heads was ____________ flips.

8) Combining the data from all lab groups, find your class overall average half-life for heads in flips. Show your work.

The class average half-life for heads was ____________ flips.

Questions:

1) Is your average half-life for heads the same as what you predicted before the lab? Why or why not?

2) Is the class average half-life for heads closer to the theoretical average than yours? Explain why or why not.

3) Did you notice anything strange about the observed half-life when you got to the last few pennies? Explain, citing your data where appropriate.

4) ^{194}Hg has a half-life of 520 years. How long will it take for **one** atom to decay? Explain your answer.

<u>Nuclear Chemistry Worksheet</u>

1) Explain why some atoms are stable and others are not.

2) What are the three main types of radioactive decay? What particles are given off during these processes?

3) What is “half-life”?

4) An artifact contained 1000 grams of element X when it was buried. Element X is radioactive with a half-life of 200 years. If this artifact is dug up 1200 years after burial, how many grams of element X would you expect it to have?

References

When I sat down to write this book, I figured I'd just have a seat at my computer and polish up all the lessons that have worked well in my classes. As it turns out, that's pretty much what I did.

However, no writer knows everything possible to write a book. That's where the following references come in. When you find yourself needing to answer difficult questions, I recommend consulting the following sources, as they've all been hugely valuable to me while writing this book, and also for my teaching in general:

- The CRC Handbook of Chemistry and Physics, by the Chemical Rubber Company. Any physical constants in this book originated there. This is a fantastic book that everybody should own. If you don't own it, you're not prepared to teach.
- Physical Chemistry, by P.W. Atkins. If you have any questions about thermodynamics, quantum mechanics, or statistical thermodynamics, this book probably has the answer.
- Inorganic Chemistry, by D.F. Shriver, P.W. Atkins, and C.H. Langford. If you need to know how molecules behave under unusual circumstances, this is the source for you. It's easy-to-read and very comprehensive.
- The *WebElements* website, put together by Dr. Mark Winter at the University of Sheffield, England. If you need to know anything about the elements, this is the place to go. This site is updated frequently and is of consistently high quality. *WebElements* can be accessed at http://www.webelements.com/.

Appendix – Laboratory Safety

As every science teacher already knows, safety is the most important thing in the laboratory. In nearly every lab we do, there exists the small but very real chance that somebody could be seriously injured. Above all else, our jobs are to keep our students healthy and happy.

Anybody who's spent time in a high school chemistry lab is well aware that students can do a wide variety of dangerous things without thinking about it. I've heard about students who were blinded while looking into boiling test tubes and others who set themselves on fire with a Bunsen burner. The bottom line is this: Students are going to get into trouble in ways that are hard to predict, and it's our job to be prepared when this happens.

Wearing safety goggles is the single most important thing we can do to ensure laboratory safety. Unfortunately, it's also the thing that students like doing least. No matter what the laboratory activity, students need to wear their goggles, as very small quantities of many chemicals can cause permanent eye damage. In my class, failure to wear goggles results in instant failure in a lab. I realize this is very harsh, but it's also necessary to ensure that students take this rule as seriously as I do.

At the beginning of the school year, provide each student with a set of laboratory rules and make sure they understand them! I want to make it very clear that laboratory rules are a good start, but for some labs you may need to incorporate others. Before doing any lab, identify all safety hazards yourself. Never count on anybody else to do this for you (not even this book!) as different people have different thresholds for what they consider to be safe. Always use your best judgement and always be twice as restrictive as you think necessary. After all, you are the only person responsible for the safety of your students!

Flinn Scientific does a nice job of promoting safe laboratory practices. On their website, they provide a wide variety of safety resources to help both you and your students understand what is safe and what is not. To access their safety information, visit http://www.flinnsci.com/homepage/sindex.html.

Another good source of information is provided by the Department of Chemistry and Biochemistry at the University of California at Santa Cruz. For safety rules, emergency procedures, and a really disturbing picture of a chemically-damaged cornea, visit http://www.chemistry.ucsc.edu/Projects/Safety/.

One final source I'd like to mention is the Duke University Department of Chemistry Lab Manual, located at http://www.chem.duke.edu/safety/. It's particularly good for explaining what to do in case of serious emergency, and I personally like the section outlining what type of rubber glove is best for each process in the laboratory.