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Revolutionary Chemistry Text

Jennifer L. Carhart
Author: Jennifer Carhart

Nicholas Kildahl
Advisor: Prof. Nicholas Kildahl

Revolutionary Chemistry Text

Open almost any introductory chemistry text today and right away certain problems with the format and wording become evident. These problems present learning barriers for the student, so instead of using their time to understand the material, they battle with trying to understand the way the material is presented. On a higher level, and further into the text, even more problems present themselves, such as barriers between different fields of chemistry, and making the science look archaic. Noted problems with current textbooks include the following:

- Doesn't grab the reader – the text is dull, and often leaves out any application to the real world that chemistry actually has
- Layout – Often the order in which authors lay out their material does not follow a logical flow for most students and even many chemists
- No purpose – Again, little application to real life situations for chemistry is presented, and also seldom explained is why Chemistry is an important subject matter to learn
- Static – Many texts give the impression that chemistry is a static science set in its ways and that it never changes
- Wording – Often language and examples chosen can confuse or even bore the student; the way material is presented is key to student learning and understanding
- Boundaries – Frequently boundaries between organic, inorganic, physical and other branches of chemistry are imposed on the student, and thus give the impression that inorganic concepts cannot be applied to organic chemistry, etc.

The problems listed above are major ones, and have been enough to turn many students away from continuing their education in chemistry, and makes it one of the "harder" subjects to learn. In changing the way chemical information is presented, I hope to enable people to not only learn chemistry more easily, but care about the subject matter more. My main goals in writing this book are as follows:

- The Why – I hope to give a good explanation as to why Chemistry is important and why the student should take the course.
- Interactive – Real life examples of how chemistry is applied in daily life, examples that the students can contemplate themselves and relate to
- Break the barriers! – Try to tear down the barriers between different fields of chemistry, as has already been done in research
- Layout – Present the information in a more logical flow, and make sure all latter chapters are backed by fundamental basics laid out in earlier chapters
- Experimental to Theory – My main focus and goal in presenting the material is to show that chemistry stems from experimental observation. Often current texts give the impression that the theory was created and then experimental evidence was found to support it. This order is wrong and the impression needs to be rectified.
- Scientific Method – In conjunction with the previous goal, I hope to present the material in the format of the scientific method. Thus, the student is exposed to and gets in the habit of viewing chemistry in this way, without having to actually

> go through the mechanics of memorizing the steps. It will not be another “thing” they have to learn, but rather will be ingrained in the way they learn.

Obviously it is impossible to write the perfect textbook, and problems with my format will also present themselves. However it is my hope that a book such as this will finally open the eyes of many students to the magic of chemistry, and open doors for them to walk through and explore themselves.

Chapter Outlines

I. Chapter 1 – From Lead to Gold to… Teflon!
  A. How Chemistry applies to life
    1. Teflon story and how chemical discoveries impact everyday life
    2. Chemical applications to food, clothing, music and technology
    3. Modern chemical research impacting life in the future
  B. Why Learn Chemistry
    1. How observation leads to discovery
    2. Learning chemistry teaches students how to learn
    3. "Hands on" science to understand how processes work
  C. How Society Uses Chemistry
    1. Impact that new knowledge has on a global scale
    2. Chemical knowledge used for both positive and negative purposes
    3. Why the word "chemical" has a negative connotation today
  D. Still… So what?
    1. Learning the ability to analyze and understand your world
    2. New minds opening new doors and helping the barrier revolution
    3. Employment opportunities
    4. Many fields have a chemical foundation

II. Chapter 2 – The Construction of the Universe
  A. What is this stuff we're made of?
    1. Basic principles of matter and energy
    2. Phases of matter
    3. Conservation laws
    4. Introduce the atom
  B. From the largest to the smallest
    1. Universe is made up of particles
    2. Basic atomic structure
    3. Interactions between atomic particles
  C. Will the real Mendeleyev please stand up?
    1. Introduction to the periodic table
    2. What the different columns indicate (ex. Alkali)
    3. General idea of why the elements are grouped as such

III. Chapter 3 – One Universe, One Ruler
  A. How big, how heavy, how long?
    1. Measurements for Mass, Volume, Length etc.
    2. Conversions
    3. Proper ways to read laboratory instruments
    4. Significant figures and calculations with sig. fig.
  B. Atomic Society
    1. Atomic mass and isotopes
    2. Formulas and molecular mass
    3. % composition and mass spec.

C. Putting on the White Coat
    1. Ways of analyzing and problem solving in the lab
    2. Uncertainty, precision and accuracy revisited
    3. Actual application of principals learned

IV. Chapter 4 – You Are One Attractive Element!
A. Let's Have a Bonding Experience
    1. What are bonds
    2. Ionic
    3. Covalent
    4. How to predict molecular formulas
B. Stoicheometry from Percent Composition
    1. Percent comp to formula
C. A Clue Sherlock! From Data to Structure
    1. Percent comp and mass spec data
    2. Finding formula and weight
    3. Finding structure

V. Chapter 5 – Why Don't We Get Together and React!
A. Did You Just Hit Me?
    1. What are reactions
    2. How to write reactions
    3. Balancing reactions
B. Moles Do Not Dig in Dirt! – More Stoichiometry
    1. The mole
    2. Molar ratios in reactions
    3. Going from a mass of reactant to finding the mass of product
C. Are You Gunna React? Or Do I Have to Make Ya?
    1. Endo/exo thermic
    2. Free energy and spontaneity
    3. Enthalpy and Entropy

Chapter 1 – From Lead to Gold to… Teflon!

Section A – How Does Chemistry Apply to Life?

It is known as the most slippery substance on earth. You find it everywhere: cars, hospitals, clothing, spacecraft, computers, atomic bombs and even in the kitchen. Yet this versatile and useful product was discovered unexpectedly. On April 6th 1938, a young chemist, Roy J. Plunkett, was trying to find a new gas that would act as a non-toxic refrigerant. One experimental batch of Freon gas had been left overnight. Plunkett and his assistant, Jack Rebok, were going to analyze the gas to see if it was experimentally what they wanted. However, they found no gas to come out when the valve was turned. Thinking the gas had leaked out completely, Plunkett weighed the cylinder, only to find it had the same mass (mass is a measurement of weight) as before. Upon opening the cylinder they discovered a greasy white powder. It was subsequently shown that the substance was a polymer with the name polytetrafluoroethylene (PTFE). Obviously not a very catchy name, Plunkett shortened it to "Teflon," which is now a household name. The story of Teflon and other chemical discoveries and applications, whether they are accidental or not, make us aware of the diversity and value of chemical knowledge to our world.

Feel the pages of the book you are reading, the floor underneath your feet, the texture of your skin. The entire world is made of chemicals. The American Heritage Dictionary defines a chemical as a substance with a distinct molecular composition that is produced by or used in a chemical process. What a narrow definition! Chemicals are not only "special" substances only used by chemists, everything is a chemical. From the water we drink to the clothes we wear to the oxygen we breathe, there is not an object in this world that isn't a chemical. Chemicals do have a distinct composition, just as we all have our own fingerprints. And they are used in processes, but then the act of breathing is a chemical process. Simply put, chemicals are combinations of atoms made up of the known elements on the periodic table. That includes you.

Ok, you say, but so what? If we know that everything is a chemical, then why bother studying it? The reason it is important to study chemistry is many-fold, but one of the most important reasons is due to the heavy involvement chemistry plays in our lives. Teflon has already been mentioned, letting your car run smoothly, your computer run cooler, Gore-Tex clothing stay dry, and even helps to make cleaning those scrambled egg pans from breakfast a breeze. What other applications does chemistry have? Chemistry is vital in meeting the basic human needs of food, clothing, and shelter. It is equally vital to the advancement of modern technologies such as computer chip manufacturing, space exploration and energy production. In short, chemistry is vital! From the inner metallic heat sinks on the chips to the plastic polymers that house everything, your computer is the result of chemical research. Astronauts can only survive if they understand the composition of the air required and how to recycle and best use the chemical material their body generates. Common cotton clothing is colored using dyes (chemicals!), and can now be waterproofed or made flame retardant. Buildings are made taller and with heightened stability due to the discovery of new metal alloys and lighter materials. The kitchen offers a plethora of products resulting from chemical development. Food has been enhanced via chemical research to increase nutritional value or shorten the time and ease the preparation of cooking. In fact, the act of cooking is simple chemistry; your kitchen is a laboratory!

Even more vital to the world than what chemistry has already produced is what can be discovered and produced in the future. Medicinal laboratories are working year-round to find cures for terminal illnesses. Processes to recycle and reuse materials with increased return must be found for our planet. New fuel sources and energy harnesses, which are ways of better capturing the energy from sources such as the sun, which need the right chemical composition of the solar cells to capture the light for transformation into energy, are being researched in order to preserve our natural resources. The future in chemical research will be of the utmost importance to the survival of not only our species, but the rest of the species on earth and the planet itself.

Section B – Why Study Chemistry?

Even with all the reasons given above for the importance of chemistry in our world, obviously not everyone becomes a chemist. Why then, should you study chemistry if you're interested in another subject or career path? Simply put, the study of chemistry builds observational and analytical skills that can be used in any field of study, as well as in everyday life.

Consider, for a moment, a simple product that has found many applications in the modern world: Velcro. While taking his dog for a walk, Georges de Mestral, a Swiss Mountaineer, noticed the burrs from a plant sticking to his dog's fur, and to his own pants. When he got home, he looked at the burrs under a magnifying glass, and found tiny hooks all over them. Based on this observation, he invented the hook and loop system of fasteners, which is today known as Velcro. He observed the natural world around him, analyzed how the hook and loop system worked, and created his own version after several experimental trials. The discovery of Velcro represents the heart of chemistry: observation leads to discovery. It illustrates a logical procedure that chemists often use to look at the world. Suppose we observe something interesting in the environment, for example, you notice one day that there is soap leaking down the side of your bathroom sink. Then, experimentation and information collection follows. Where could the soap possibly come from? Quick deduction should lead to the soap dispenser on the edge of the sink. But why is the soap coming out of the soap dispenser? Does the soap appear everyday? What was the weather like on days that the soap appeared? What material is my dispenser made out of? Analysis of the results often reveals a trend in the information gathered. When I used a plastic soap pump dispenser, soap would appear on the days there was high barometric pressure. Days with high temperature also seemed to correlate with the soap appearing. If I used a ceramic pump dispenser, no soap appeared at all. Finally, an end result, sometimes a discovery, sometimes a confirmation of the observation, is reached. With plastic soap pump dispensers, high pressure and temperature will lead to soap leaking from the dispenser. This observational and analytical process is honed in chemistry, but certainly is not restricted to application in chemistry. Analysis of great works of literature, understanding the political factors that shape history, even finding out what is making that funny noise in your car, are all problems that can be approached in the same way. A study of chemistry can provide you with skills of analysis useful in many fields.

Although opening the possibility to develop observational and analytical skills, there is much more to chemistry than that. Chemistry is very much a "hands on" science, with our knowledge of the world stemming from experimentation. If observation and discovery are the heart of chemistry, then experimentation is the body. Taking an active role in getting to the core of a problem is another useful tool for everyday life. Since even a shallow study of chemistry involves experimentation and personal participation, we begin to understand that through participation we can better influence our world. Whether it is participation in government, sports teams or any social activity, getting involved makes any activity more enjoyable and personal. Observation is important, but only if backed by experimental participation. In addition to reinforcing the idea of participation, the study of chemistry builds and refines problem solving skills. The ability to not only see the problem, but actually "get your hands dirty" and solve it is a skill everyone should have. Consider a problem with something in your daily life; for example, your doorbell to your house isn't working. One way to deal with the problem is to call someone else to fix it. But now you are dependant on other people, and have gained no solution of your own. Another approach however, is to systematically test each component of the doorbell system. But how should we test it to find out which component is not working right? First, remove the battery, leaving the rest of the system intact, and try testing it. If the battery is working fine, put it back and check the wiring. If the wiring shows that the circuit is complete, try the button mechanism. This method of problem solving, where one component of the problem is tested, and the others are kept constant, is a very commonly used method in scientific experimentation. Once the problem has been found, action can be taken to correct it. Now you have solved the problem yourself, not only giving you independence and a feeling of satisfaction that you solved it yourself, but you may have even learned something about the way circuits work! The ability and desire to learn and solve problems on your own is a crucial skill to survival in today's world.

Understanding how to benefit from a mistake or accident is yet another skill chemistry provides. Think back to the Teflon discovery. If Plunkett had looked at the gas canister as a failed experiment and discarded it, the world might not have Teflon. Many scientific enhancements to the human world have come about this way; a researcher got an unexpected result, but rather than toss it aside, s/he explored it, and found something very useful to human society. A good grasp of the idea "when life gives you lemons, make lemonade" is important for every facet of life, giving you the ability to make the best of any situation (and in Plunket's case that turned out to be quite profitable!).

Being a "hands on" science also makes chemistry interesting. It is one thing to learn about acid base reactions, but captivation by the bubbling fizz created when you pour baking soda into vinegar encourages us to get involved. The ability to conduct experiments in a laboratory is one of the most fascinating aspects of chemistry. With the proper knowledge, new compounds can be discovered, new medicines, even better and simpler ways of creating existing compounds. Reading Shakespeare won't make you write like a Pulitzer Prize winner; it takes years of practice, and even then some people seem to "lack talent." However, studying chemistry, even for a short period, gives you the power not only to reproduce what you've learned but to discover, even create, new things.

Section C – Chemistry and Society

Obviously all the previously mentioned reasons for the study of chemistry show the importance of chemistry to the individual. Chemical products have been used on a world-wide scale in every part of life. But how do new discoveries affect society? How does chemistry affect the earth?

In today's modern society, with the instant transmission of information across continental and multilingual boundaries, it is foolish to think that any new discovery in any scientific realm, particularly one with important economic or environmental properties, will not be exploited and researched to the extent of human capacity. Chemical applications can have extremely positive impacts on the entire world. Globally, chemistry stretches into all aspects of international affairs, from the materials used to house relief packages for 3rd world countries, to the manufacture of pharmaceutical compounds and refinement of fuel sources that play such a vital role in global economic associations. Increased access to chemical knowledge allows pooling of resources and ideas from different nations, allowing for a variety of perspectives on discoveries or methods of looking at how the chemical world works so that the scientific community as a whole might gain a better understanding of a specific finding. A chemical discovery made in one country may provide relief to another. For example, the chestnut trees in the United States were nearly all wiped out due to a tree fungus known as blight. The chestnut trees in China, however, were discovered to contain substances in their bark, which made them resistant to fungal infestation under normal conditions. Thus some salvation efforts for the American chestnut tree are underway by incorporating the genes that produce these chemicals in the Oriental chestnut tree.

As we have seen, chemical information can have a vast positive impact on a global scale and even on international relations. However, chemical discoveries can have a negative impact on the earth as well. It often happens that when a negative sociological event results from a chemical discovery, the chemist is blamed. It needs to be clarified that chemists, like most scientists, are interested in understanding the world around them; they seek an understanding of the processes observed in nature, and look for answers to the question "what happens if…" How chemical information is used depends on society, and what is happening on a global scale at the time. It is not necessarily the responsibility of the discoverer. For example, in the late 1800's, the only source of nitrates (compounds containing nitrogen) were from soil deposits in Chile. However, with growing agricultural needs, it became quite clear that the deposits in Chile would not be enough to feed the world. Thus, in 1909, Fritz Haber, the German physical chemist, created a process to synthesize ammonia. This process was further developed by Carl Bosch to make it economically viable. Both chemists received the Nobel Prize in 1918 for their work in this field. This chemical discovery was then put to use by society in two ways. In a positive light, ammonia is the main component of fertilizer. Without large quantities of fertilizer to make crops more productive, world hunger would be a much larger issue than it is today. The importance of fertilizer for the increase of plant biomass has effected immigration, industrial growth, and population trends in various countries all over the world. However, this process was also used to manufacture munitions for war. Germany was particularly dependant on the Chilean nitrates to make explosives. When the First World War broke out, the Allies cut off imports to Germany from the South American continent. Thus, the Haber-Bosch process was exploited to make munitions for Germany in the First World War. Conceivably, without the process, Germany would have been forced to surrender earlier than it did. Although Haber did suggest using his process for munitions manufacture, and thus is looked upon negatively by the allies, he was doing what he thought was right for his country. The ethics of scientific discovery and who is to "blame" for the exploitation of certain discoveries is a growing field of scientific research in the world today.

Surely incidents such as this give the field of chemistry a bad name. And certainly, when someone utters the word "chemical," images are invoked of harmful compounds, poisons, industrial waste, environmental pollutants, and any other toxic substance the individual has had contact with. People tend to think of the hole in the ozone layer, the pollutants in the world's marine and freshwater ecosystems, toxic substances in food, and global warming. The word, "chemical," and even the field of chemistry, have negative images in modern society. Think back to the dictionary definition given earlier, and how incorrect and narrow it was! And granted, as mentioned above, many chemical discoveries are taken by society and used to hurt each other, or help the human race while hurting the planet, and other negative means. But chemistry, the science of understanding what particles make up the world and how those particles interact, cannot be blamed for the negative exploitation of this understanding by individuals or governments. When a beneficial discovery is made, a new drug that cures a debilitating disease, for example, does anyone say: "Oh well done chemistry, hooray!" Of course not. So why does the world spurn chemistry for the bad effects on civilization it has been used for? It's simply illogical. As we have seen, chemical discoveries have both positive and negative applications. However the study of chemistry should not be held responsible for the negative applications any more than it should be credited for the positive ones. It must be kept firmly in mind that chemistry is an objective science, concerned with discovery and understanding. How the world then uses the knowledge obtained is a whole different realm of study.

Section D – Your Future with Chemistry

You may be saying to yourself, "Ok, Ok already, I get the point. Chemistry is important. But I'm still not convinced I need to study it." If problem solving or the excitement of discovery does not pique your interest in chemistry, there are still many reasons to study this branch of science.

One particular importance, not only to future scientists but to the realm of science as a whole is having new minds in the field. With new generations and new perspectives, we can obtain new outlooks on chemical procedures, ideas and analysis, both new and old. Constantly looking at the world at different angles is essential to the growth of chemistry as a whole. But there is even a greater reason for new minds to enter the field of chemistry. Current research is blurring the boundaries between branches of chemical study that have been in place for hundreds of years. Organic, inorganic, physical and biochemists are sharing research results, and using the knowledge of one field to further progress in the other. And although this "barrier revolution" is very common in modern research, the teaching of chemistry in high schools and universities remains compartmentalized by the old boundaries. Separate courses in inorganic, organic, quantum mechanics, and biochemistry are given, each with its own separate language, as if they were completely different branches of science. But they are all chemistry! Hopefully, as new minds enter the field of chemistry, we can begin to break down these old barriers and learn from each other. Certainly nature makes no distinction among the various chemical branches, and we should follow suit.

As stated in earlier sections, chemistry helps with observation and analysis. Perhaps you would just call someone to fix your doorbell, just let someone else handle the problem and be on your way. But there are certain issues we all must face in life: what will I eat today, what will I do to make a living; will I have a warm bed for myself at the end of the day? Even if you retain no chemical knowledge what-so-ever after studying chemistry, the skills built learning this science will remain with you for your entire life. What person in the modern world would not benefit from a logical and systematic approach to tackling their daily problems?

If all these reasons still don't have you convinced, consider this: science as a field is constantly growing and expanding as humans quest for a deeper understanding of how their world works. Expanding fields mean new researchers are required to fulfill the manpower needed to further research. In short, there is a job waiting at the end of a study of chemistry. Certainly even the most chemically skeptical person can identify with getting a job. Not only is the science growing, but it is a science fully ingrained in modern society, so anxiety about it shutting down is resolved. But research is not the only career path you have to take if you study chemistry. Academia, from 1st or 2nd grade all the way to college, is a rewarding career, helping other people better understand their world and how it works. Commonly in Europe, and starting in the United States, a degree in chemistry opens multiple career paths. Just as stated in this chapter, chemistry spurns analytical skills that employers love to see in their employees. Many fields from teacher to bank manager to governmental positions understand the value of those skills built.

In addition to a pure chemical degree leading to multiple doors, chemistry is either a foundation stone or a compliment for many other branches of science. Veterinary or human medicine, botany, geology, physics, biology (and all its many branches!), meteorology, astronomy and many others all require basic, or sometimes in depth, understanding of chemistry. Without chemistry, you could not have been eating a microwave dinner while watching a videotape of the lunar landing. Chemistry truly is a science that underlies every facet of modern life.

Chapter 2 – The Construction of the Universe

Section A – What is this stuff we're made of?

In order to obtain an understanding of chemistry, we first need to have a basic understanding of what our world is made of. In this chapter we will look at what makes up the world: *matter, not-matter*, and *energy*. To understand these concepts, we first have to understand a few terms. Once we understand these terms, we'll be leaps and bounds closer to understanding what our universe is comprised of. The key terms are listed below.

Element
Atom
Compound
Molecule
Mixture
Pure Substance

Over the period of human history it has been discovered that the universe is comprised of 100 fundamental substances. These fundamental substances we call *elements*. You may have heard the word "element" or "elementary" before, most of the time as a synonym for "basic." And indeed, an element is one of the basic building blocks of the world around us. Some examples of elements are iron (found in tools like hammers and nails and in skyscraper frames), oxygen (found in the atmosphere and SCUBA diving tanks), aluminum (found in cars and kitchen foil), and carbon (found in diamonds, pencil lead and charcoal). But what makes an element basic or fundamental? What would an element look like if we magnified our view of it a 1000 times? A million times? A billion? Extreme magnification of an element in its solid form reveals tiny "spheres" aligned in specific patterns. These spheres are called *atoms*. You can think of atoms like Lego building blocks. There are many different kinds of atoms, like there are different shapes and colors of Legos. What makes an element fundamental is that all the atoms that make it up are identical; an element is made of "units" that are all the same shape and color. What makes one element different from another is that the "blocks" are different. In other words, elements differ from each other because their atoms differ. Naturally, atoms aren't colored like Lego blocks, but they do differ in size and in how they interact with other atoms, whether they are the same or different. Are elements the only substances in the universe? Why do we have so many chemical names if there are only about 100 elements? Diversity of our universe results from combinations of elements to make other substances. And there are a huge number of possible combinations, with new ones being discovered or created everyday! Let us take a look at some of these combinations.

Substances that contain two or more elements, combined according to a set of "rules" that we will explore later in this book, are called *compounds*. Compounds surround us in everyday life. Substances such as water (made up of the elements oxygen and hydrogen), sugar (carbon, hydrogen and oxygen), salt (sodium and chlorine), Teflon (fluorine and carbon), and DNA (carbon, hydrogen, oxygen, nitrogen and phosphorus) are compounds. And what would we see if we magnified our view of a compound?

Would we see our atomic Legos of all the same shape and size like in the element? No, we have already said that compounds are combinations of different elements. So are the different atoms all mashed together in a messy heap? No again. An extreme magnification of a compound again reveals basic building units, but now each unit is an aggregation of atoms of the constituent elements. Further, these atoms are attached to each other in a very regular pattern. And, just as the atoms are all the same in an element, these units are all the same throughout the compound. We call these units *molecules.* A molecule is a specific combination of atoms. A molecule of water, for example, is made of one oxygen atom and two hydrogen atoms. These atoms always arrange themselves in the same way, as shown below in figure 2.1.

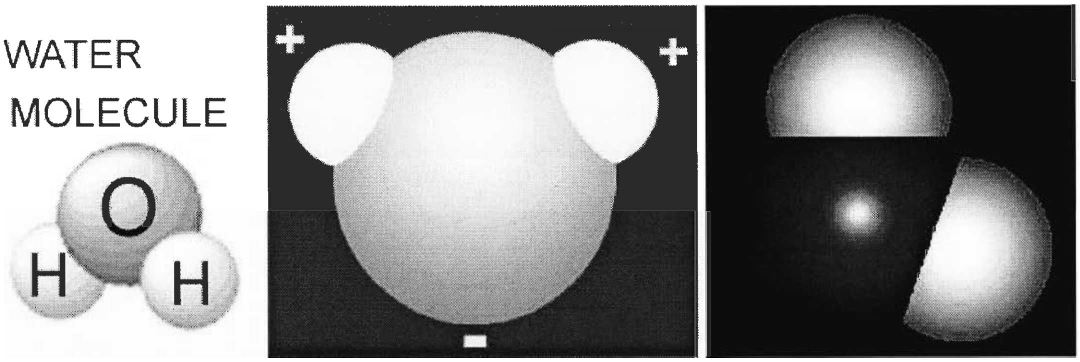


Figure 2.1 – Different representations of the water molecule.

As you can see, there are several ways we can depict molecules pictorially, but you should notice that in all these pictures, the basic structure is the same. All the macroscopic objects you see before you are made up of trillions of molecules.

Sometimes, however, we can combine elements and/or compounds together in a different way. Imagine taking two different atomic Legos and instead of connecting them together, just take the different colors and shapes and put them into a bucket together. What should we call that now? We call this type of substance a *mixture*. Mixtures are comprised of two or more elements or compounds blended together, but not bonded to each other in any sort of fundamental unit. Extreme magnification of a mixture may reveal what looks like an organized unit, but a quick look in any direction will reveal different looking units. In other words, mixtures contain two or more types of fundamental units. The other difference between mixtures and compounds or elements is that mixtures can be separated without having to dismantle any major organization. When we have our atomic Legos in the compound, to separate them back into elements we have to take the Legos apart, break them down. But in the mixture of Legos in the bucket, all we have to do is reach in and separate the different shapes and colors.

The last term we should understand is *pure substance*. In this chapter we have been using the word "substance" as a synonym for "stuff." Pure substances, however, are as their name implies: pure. What we mean by pure is that all the fundamental units within a pure substance must be the same. Thus, elements and compounds are pure substances, consisting of atoms and molecules, respectively. Mixtures are not pure substances because they have more than one fundamental unit.

The word, matter, traditionally refers to any object, from your cat to a cloud in the sky, which takes up space and also has mass. Fair enough, we can imagine our atomic Legos taking up space. But what about mass? The definition for mass traditionally is the amount of matter in an object. But wait a minute. If matter is defined by having mass, and mass is defined as the amount of matter in an object, how do we get a sense of what we are actually talking about? Obviously this circular definition is confusing and therefore incomplete. Matter consists of elements, compounds and mixtures. If you understand that pure substances are matter, then the concept is much easier to grasp. Anything made up of the 100 fundamental atoms is matter.

At this point you might be asking yourself whether there is anything that is NOT matter! Yes. Think about light for a moment. Does light take up space in the universe? Well, perhaps some theories about light suggest it might. Does light have mass? Ah ha! Light has no mass and therefore is not matter. So what do we call things that aren't matter? In the case of light and other electromagnetic radiation (from gamma rays to radio waves), we associate it to something called a "field". Fields are regions in space that have a particular "influence" imposed on them. We are sitting in several fields right now. For example, the gravitational field is pulling us to the center of the earth. Gravity is not matter; it has no volume or mass. The needle of a compass points north because the needle aligns with the magnetic field of the earth. Magnetism is another field and therefore not matter. For now, you can very loosely think of matter as being something you can touch, or interact with. You can't get a hand around gravity any more than you can get a hand around the human sense of justice or dignity. Other phenomena that are not considered to be matter are the human soul, the spirit, even things like ghosts! None of these things is matter.

There is in our universe "stuff" that people often think of (and that its name suggests) is another form of "not matter": antimatter. The concept of antimatter arose from scientists trying to predict the movement of some of the particles of "normal" matter. They found that their models only predicted correctly if they assumed that the particles had mirror images, but negative mirror images. Think about playdoh or clay, and roll it out on a table and take a cookie cutter. Cut a shape in the clay and hold it up. Now look down. No matter what shape you cut in the clay, a negative mirror image will be left behind. This is the basic idea behind antimatter; for all matter, there exists a negative copy. The importance of antimatter and how it relates to matter will be discussed later in this book.

Now that we have seen what matter is (and is not!), we can begin to talk in more depth about the different forms of matter. Although we concluded that both cats and clouds were types of matter, obviously they are different in many ways. One of these differences we can look at is known as their state, or *phase*, of matter, meaning what form of matter they are in. The three most common states of matter are: solid, liquid and gas. Solids are rigid forms of matter; the atoms or molecules are packed very closely together, restricting their motion from place to place, allowing them only to vibrate next to each other. A hammer is a solid form of matter. Liquids are fluid forms of matter; the atoms have more freedom of motion than in solids, and can slide around one another. Water is an example of a liquid. Gases are also a fluid form of matter; here however, the atoms have even more freedom of motion than in liquids, and often can pass one another without touching. The air in a balloon is an example of a gas. There is a fourth phase of matter called plasma, but we do not need to discuss that phase just yet. The three common phases, solid, liquid and gas, also exhibit differences in the shape they assume. Solids have their own distinct shapes. Look at your hammer. If you were to put it inside a balloon, would it become balloon-shaped? Of course not, it would just be a hammer inside a balloon. Solids retain their shapes unless changed by the input or release of energy. A liquid, on the other hand, takes the shape of whatever container it is held in. If you put some water into the balloon, does it stay in the shape of a cup inside the balloon? No, naturally it takes the shape of the balloon. Liquids have no shapes of their own; their shapes are defined by the containers they are in. Like liquids, gases also take the shape of whatever container they are in. If the balloon is filled with air and we change the shape of the balloon, the gas also changes shape. The difference between the gas and the liquid, however, is that the gas occupies all of the available space in the balloon. If you only fill the balloon halfway with water, the water does not expand to fill the entire balloon. But air does. Filling the balloon halfway with air does not mean that the air stays in one half of the balloon while the other half is empty. The gas expands to fill the whole balloon; that is the difference. Besides the differences in the freedom of motion and in shape of the states of matter, there is another difference between the three states. This is volume, or the amount of space a phase of matter takes up. Just because matter has volume does not mean that it is the same for all states, however. For example, assume we have an ice cube. Now we put it in a balloon and seal it so that only the ice cube occupies the space of the balloon, that is, nothing else other than the ice cube is in the balloon. How much space does the ice cube take up? Not very much. Now suppose we melt the ice cube. How much does the volume inside the balloon change? Not very much, since the change in volume between the solid and liquid phases of matter is not that extreme. Now suppose we boil the water inside the balloon. Will we see a volume change? Yes, the balloon expands since the gas particles are father apart than in the liquid phase, resulting in a larger volume. It's important to understand that while the volume change between a solid and liquid phase is not large, the volume change between a liquid and a gas is. Gases take up considerably more "space" than liquids and solids. The differences between phases of matter may seem rudimentary, but they are important in understanding how different substances interact.

The last concept discussed in this chapter is the companion of matter: energy. You might look at the lamp sitting next to you and say, "Oh yeah, I know what energy is." But scientists define energy as the ability to do work. While we might only think of energy as electricity, energy of many different types exists in the universe, responsible for everything from feeding you to traveling the globe to making wind. Some common types of energy are: chemical energy, thermal energy, electromagnetic energy, kinetic and potential. We will take a closer look at energy later in Chapter X. Keep in mind that whenever we talk about energy being "stored" we mean its ability to do work, like lifting an object in space, or moving through a circuit. It is also important to understand what scientist call "conservation laws." These "laws" simply state that in any process, energy and matter are conserved. All this means is that we cannot create or destroy energy or matter. The atoms in your body have been in our solar system since it was formed. Truly, we are all made of stars.

Section B – Taking a Closer Look

In the last section we introduced elements and compounds. We discovered what we would see if we could magnify them, and found that both are made of fundamental units. Specifically, we found the fundamental unit of the element to be the atom. But what would we see if we were to magnify the atom? Is there an even smaller fundamental unit than the atom? If we were to magnify the atom a billion times, we would see not one, but three different units, or particles, that make up the atom. These particles are called the *proton*, the *electron* and the *neutron*. Let us take a closer look at each particle, and see how they are different from each other, and what basic function they serve in the atom.

The proton is a stable particle that sits at the center of an atom. This center is referred to as the *nucleus*. The proton carries a charge of 1.6 x 10-19 Coulombs. The charge is defined to be positive. However, to avoid writing this awkward number each time we describe the proton, we often reduce the number down to simply +1 charge units. Protons, because they are very small units of matter, have mass. The mass of a proton is 1.673 x 10^{-27} kilograms. This number is very small indeed, and compared to us, a single proton would seem weightless. But remember, all those proton start to add up, and certainly contribute to the weight of everything around us.

Another subatomic particle is the electron. The electron is also a stable particle, but it does not stay in the nucleus with the proton. Instead, the electron surrounds the nucleus in patterns known as *electron clouds*. We will look at this concept more closely in a minute. Like the proton, the electron has a charge. The charge of the electron is equal to that of a proton, only negative. Numerically, this number is 1.6 x 10-19 Coulombs. Once more, we use an easier representation of this charge by saying the electron has a -1 charge units. The electron is also matter, and has a mass (when not in motion) of 9.11 x 10^{-31} kilograms. As you can see, the mass of an electron is much smaller than the mass of a proton.

The third particle is the neutron. Like the proton, the neutron is also stable and resides in the nucleus. However, the neutron does not have a charge like the protons and electrons do; it is an electrically neutral particle. Its mass is 1.675 x 10^{-27} kilograms, almost the same as that of the proton.

Now that we know what the three particles are, we should explore further how they are arranged in atoms. This arrangement, or structure, is similar for all atoms. The simplest atom, hydrogen, has no neutrons. It consists of a nucleus of one proton and one electron in a cloud around the nucleus. All other atoms have neutrons as well as protons clustered together in the nucleus, which constitutes nearly all of the mass, or bulk, of the atom. The electrons then "orbit" around the nucleus as depicted below in figure 2.2.

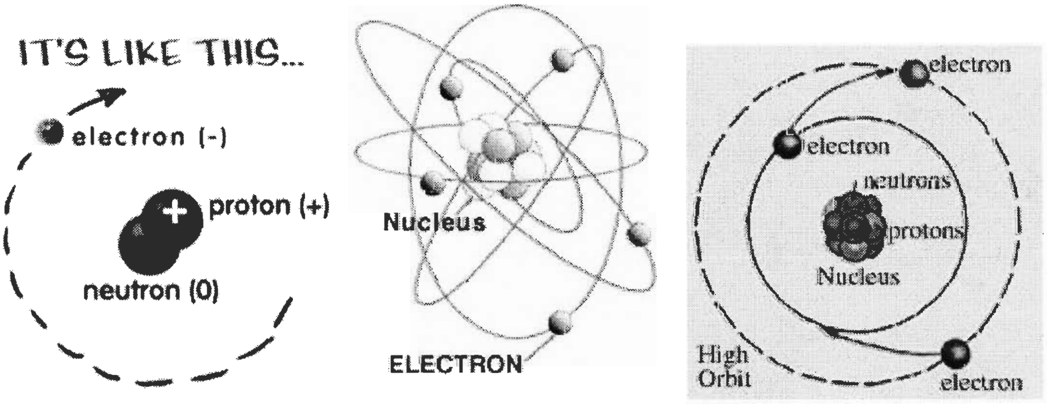


Figure 2.2 – Some representations of basic atomic structure

However, it is very important to understand that the electrons do not stay on a single orbit. The picture above is an over-simplified 2 dimensional way of representing a

3 dimensional object. In reality, electrons "fly" around the nucleus in spherical layers that we already mentioned, called electron clouds. Imagine a tennis ball. The very center of the ball can be considered to be the nucleus. The green fuzzy outer layer of the ball is one electron cloud. If you strip that layer away to the first layer of rubber, that shell is another electron cloud. The next rubber layer is another, and another, and so on. Atoms can have multiple electron clouds, but they only have one nucleus. Some additional representations of the atom model are given below in figure 2.3.

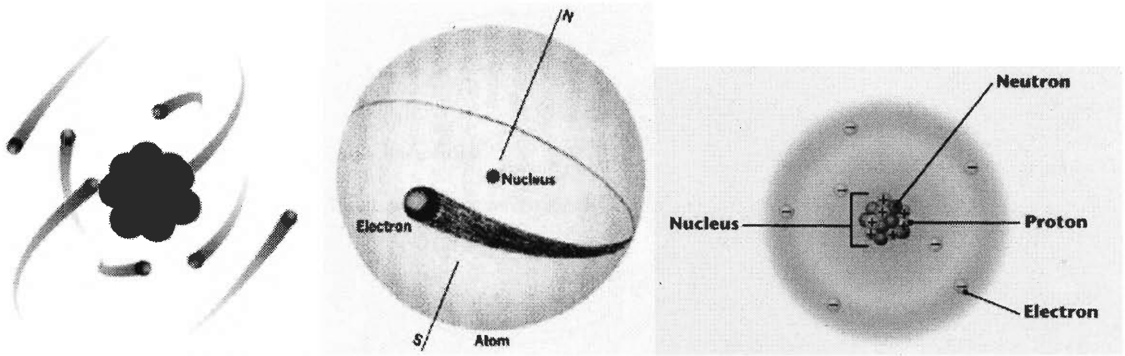


Figure 2.3 – Some representations of the nature of the electron cloud of atoms

Remember in the last section we talked about elements being different from each other because the atoms are different. What makes atoms differ is the number of each type of particle in the atomic structure. For example, hydrogen and helium are different elements because the hydrogen and helium atoms are different. Hydrogen atoms have one proton and one electron. Helium atoms, however, have two protons, two electron and two neutrons! The different number of particles is also responsible for the "size" difference between atoms; and we will look at this more closely in Chapter X

One curiosity arises from this model of the atomic structure however. What keeps the electron circling the nucleus and not collapsing into it or leaving it? Scientists have postulated many theories about why the atom is stable, and why the subatomic particles behave as they do. However, no clear cut and solid conclusion can be found from these theories; there is no simple explanation for the stability of the atom. In a nutshell, we still do not fully understand why the subatomic particles behave the way they do.

Section C – Will the real Mendeleyev please stand up?

Now that we understand that there are different atoms, corresponding to different elements, we need to figure out a way to keep track of them. Very early on, it was enough to simply keep a list of the known elements. But as time went on, and more and more elements were identified, this method was unsatisfactory. The simple list was becoming complex too quickly. How were chemists going to decide how to classify the elements?

The first problem was to universally decide on a way to distinguish one element from another. We have already seen that atoms (and thus elements) differ in the number of subatomic particles they contain. Should we then focus on just one of the particles in order to find a suitable method of classification? Should they all be considered? At the time of this process, subatomic particles had not yet been discovered. How then should chemists organize the elements? Elements were initially given an "order" based on *atomic mass.* As we stated in section A, the atomic mass of an element is based mainly upon the number of protons and neutrons in a given atom. Atomic mass was the way scientists chose to arrange the elements because it was one of the only properties of the atom known. In absence of the knowledge of subatomic particles, mass was the only way to distinguish between the elements. This method proved to be the best workable system of ordering the elements.

But although the elements had an order, they were still only organized as a list. However, scientists were beginning to discover that certain elements behaved very similarly to other elements. This introduced a good way in grouping some elements with other elements, providing another system of classification. But the groups of similar elements did not adhere to the order of atomic mass. Which system was right? In 1863, a French geologist, A. E. Béguyer de Chancourtois decided to arrange the elements by increasing atomic mass. He also decided to wrap the list around a cylinder so that several sets of similar elements lined up, thus solving the dilemma of which system of classification to use. This now gave the system a 3-dimentional nature however. Around the same time, an analytical chemist from England named John A. R. Newlands was also arranging the elements in a similar cylindrical fashion. He noticed that the chemical groups repeated every eight elements. He named this the octave rule, and compared it to a musical scale. However, some skeptical members of the English Chemical Society considered this absurd, and unfortunately his work was ignored for years. Newlands' research, however, did make it to the desk of a German chemist known as Lothar Meyer. In 1869 he arranged the elements 2 dimensionally in a table. This table included both the order of atomic number as well as the grouping of elements by properties of the elements and was accepted widely for many years. But it left no room for new elements, and there were still exceptions to the ordering system. However, only one year earlier a Russian chemist named Dmitrii I. Mendeleev proposed a similar table, which was scientifically much more useful. Mendeleev's table arranged the elements according to atomic mass, but also put the elements into groups, or families, that reflected similarities in how the elements reacted which each other. Unlike Meyer's table, Mendeleev's table followed the "octave rule" the Newlands had proposed. The groups also left gaps on the table, which were later filled in when new elements were discovered. Although about 130 years old, Mendeleev's system of arranging the elements is still used today. We know this system as the Periodic Table of the Elements.

A detailed periodic table is given in the front cover of this textbook. But for a rough idea, let us look at the table below in figure 3.4.

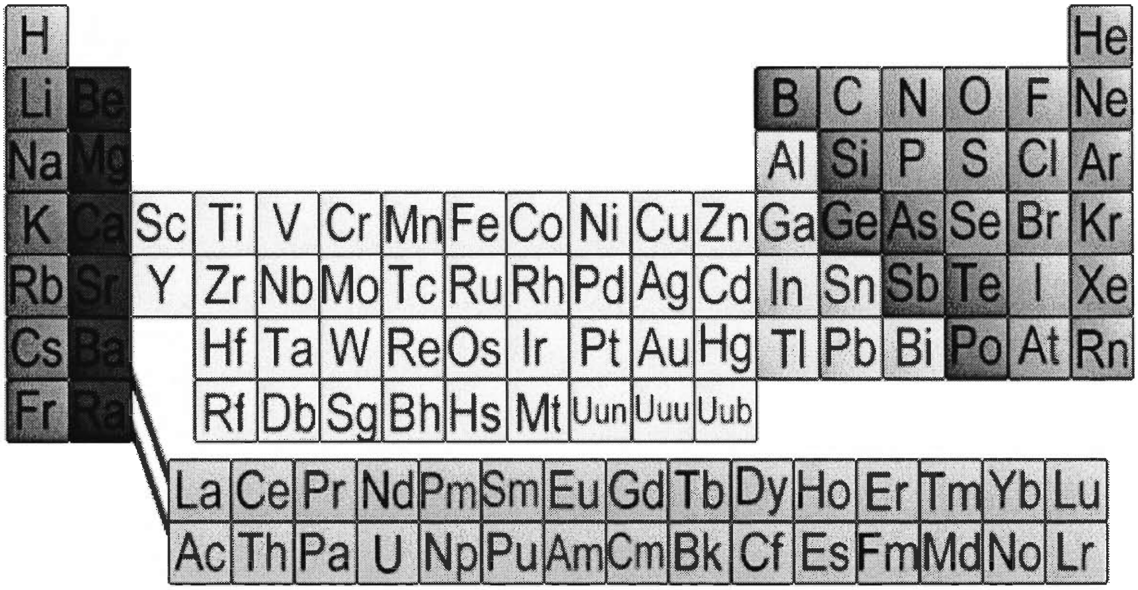


Figure 2.4 – Simple Periodic Table of the Elements

The first thing to note is that the table has horizontal rows which go from left to right across the table. We refer to these rows as *periods* (hence the name *Periodic* table). The periods all have one thing in common which each other. Remember when we spoke earlier about the electron clouds in different atoms? We stated that atoms could have several electron clouds. All the elements in a period of the table have the same number of electron clouds in their atoms. They have a different number of electrons, but the number of "shells" containing those electrons is the same across a period. The next thing we should look at is the columns of the table. These run from the top to the bottom of the table, and are called *groups*. The groups are similar to each other because they all have the same number of electrons in their outermost electron cloud. For example, in the first group of the periodic table all the elements have one electron in the outermost electron cloud, and the second group of elements all have two electrons in the outermost "shell." This business of electron clouds and where the electrons are within each cloud is very important to how different elements interact with each other as we will explore later.

The groups define collections of elements that share very similar chemical properties. For this reason, they are also called *families*, and some are given special names. Often the elements within a family can be substituted for each other in chemical reactions. Let us explore some of these families and why they were given a specific name. The first family of the table is referred to as the Alkali metals. This family includes the entire first group of the periodic table except for hydrogen. In figure 3.4 the alkali metals are shown in blue. All the alkali metals exhibit similar behavior in their interactions with other substances. For example, throwing an alkali metal into water generates a very violent reaction, often accompanied by an explosion and fire! (Do not try it at home!) The alkali metals get their name for two reasons. First, these elements are very commonly found in naturally occurring bases, or "alkalis." Although we commonly today use the word "basic" to describe a base, chemists use the word "alkaline." Since the elements of this family are found in alkaline substances, they were called alkali. The metal part of the name comes from their shiny appearance and malleability, which is the ability to be deformed and molded. Another family is known as the halogens. This family is the second to last group on the table, shown in pink on figure 2.4. The halogens all have seven electrons in their outermost electron cloud, and often combine, or bond, easily with the alkali metal family. The noble gases is the last group of the periodic table, colored brown in figure 2.4. Because all the elements in this family have the maximum number of electrons in their outmost shell, they are difficult to combine, or react, with other elements. This leads to another name for this family, the inert gases. Simply, the elements of this family occur most commonly as gases and are not found combined with other elements in nature. Another important grouping of elements is actually made up of many families. This group of families is called a "block" and is known as the transition metal, represented in figure 2.4 by the yellow block of the table. These metals tend to be strong and are used in a variety of structural and commercial applications. For now, understand that the transition metals are different because each element within the block can form complex molecules that the other families on the periodic table cannot.

Now that we have a good general sense of how the elements are arranged, let us take a closer look at how to read the information given to us in the table. Below in figure 2.5 is a basic single cell of the periodic table corresponding to a specific element, carbon.

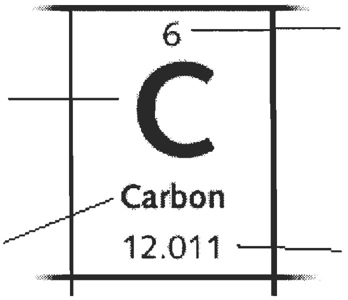


Figure 2.5 – Simple cell of the Periodic Table of the Elements – element of carbon

Here we see various letters and numbers displayed. Depending on the periodic table you look at or use, more numbers and even symbols will also be in the cell, but most periodic tables have at least these 4 entries in each cell. The top number is the atomic number. We have already defined the atomic number as the number of protons within the nucleus of an atom. Since atoms typically have an equal number of protons and electrons, this number also tells us how many electrons are in an atom. The next entry is the atomic symbol. In this example, the atomic symbol for carbon is the letter C. This symbol is an abbreviation for the element so scientists do not have to write out the entire name when diagramming reactions, writing out complex structures and so forth. Underneath the atomic symbol is the most commonly used and widely accepted name of the element. Again, here we have carbon. Lastly, we come to the atomic mass, mentioned earlier as one of the criteria scientists formerly used to organize the periodic table. Although it is not used anymore as the primary organizational feature of the table, atomic mass is still vitally important in analyzing compounds and substances. We will take a more in-depth look at atomic mass in the next chapter.

III. Chapter 3 – One Universe, One Ruler

Section A - How Big, How Heavy, How Long?

In Chapters one and two, we spoke of the importance of chemistry and looked qualitatively at some fundamental concepts. At this point, we have gone as far as we can in our quest for knowledge in a qualitative manner. In order to get a better understanding of chemistry, we need to be able to quantify things as well. Thus, we must discuss the systems used to measure mass, length, volume, time, and other physical properties of substances and their behavior. One such system has been officially adopted by scientists. It is called the system of *SI Units* and will be used commonly throughout this book. There are many properties we can look at the units for, but for now let us concentrate on units that describe four of these properties, mass, volume, length and time. You may be very familiar with units for these properties, but let us examine the scientific units for each in more depth.

We looked at mass earlier when we were describing matter. Mass is the amount of matter in a given object. The mass of an object never changes. However, often when we think of mass, we think of weight. Weight and mass are related, but they are different! Weight is a measure of the effect of gravity on an object. Weight changes with changes in gravity, but mass never changes, since no matter what planet you go to, you will still have the same amount of matter. The standard unit for mass is the kilogram (kg). Objects like this book and our bodies are easily measured in kilograms. However, often we deal with much smaller objects, which are less than a kilogram in mass. Instead of using large decimals and fractions, scientists use different fractions of the unit. These units all measure mass, and are related to each other by factors of 10. The table below shows the different units for mass.

Table 3.1 – Units for measuring Mass

1 kilogram (kg) = 1000 grams (g)
1 gram (g) = 1 gram (g)
1 milligram (mg) = 1 x 10^{-3} grams (g)
1 microgram (μg) = 1 x 10^{-6} grams (g)
1 nanogram (ng) = 1 x 10^{-9} grams (g)

lthough the gram appears to be the basic unit for mass, the SI Unit system uses the kilogram. All measure mass, and we will soon see how to convert between the different types.

The next property we should consider is volume. We also mentioned volume earlier when we looked at phases of matter. Volume is the amount of space an object takes up in our universe. All 3-dimensional objects have volume. The standard unit of volume in the SI Unit system is the liter (L). Just like mass though, often we might work with volumes that are smaller or larger than a liter. Thus we also have different fractions of the units for volume as well, shown below in table 3.2.

Table 3.2 – Units for measuring Volume

1 kiloliter (kL) = 1000 liters (L)

1 liter (L) = 1 liter (L)
1 milliliter (mL) = 1 x 10^{-3} liters (L)
1 microliter (μL) = 1 x 10^{-6} liters (L)
1 nanoliter (nL) = 1 x 10^{-9} liters (L)

In practice we use the milliliter more often than the liter and the gram more often than the kilogram. However, when reporting results or using the data collected in equations, SI units are typically used in order to keep the results universal.

Another property we should define is that of length. Length is important in measuring distances between atoms, length of particle waves, and other behaviors of substances. Length is defined as the linear extent in space from one end of an object to the other, along the longest dimension. What that means is length measures an object from one end of the object to another end, in a straight line. The standard unit for length is the meter (m). Like volume and mass, we have other forms of the meter to measure different lengths. These are given below in table 3.3.

Table 3.3 – Units for measuring Length

1 kilometer (km) = 1000 meters (m)
1 meter (m) = 1 meter (m)
1 millimeter (mm) = 1 x 10^{-3} meters (m)
1 micrometer (μm) = 1 x 10^{-6} meters (m)
1 nanometer (nm) = 1 x 10^{-9} meters (m)

The last property we will look at for now is time. Time is important in measuring the speed of a reaction, the speed of waves, and any particle or substance that is in motion. Time is hard to define, and we can look at time as a continuum of events that move the future to the present then into the past. The standard unit of time is the second (s). Measurements of time, like the other properties, have multiples of the second that are convenient for measuring different ranges of times. Table 3.4 below shows the proper measurements for time.

Table 3.4 – Units for measuring Time

1 kilosecond (ks) = 1000 seconds (s)
1 second (s) = 1 second (s)
1 millisecond (ms) = 1 x 10^{-3} seconds (s)
1 microsecond (μs) = 1 x 10^{-6} seconds (s)
1 nanosecond (ns) = 1 x 10^{-9} seconds (s)

The kilosecond is rarely used, due to the more common use of other units, such as hours, that we will look at below. The second is the most common measurement of time experimentally, and also the accepted unit for measuring time in the SI Unit system.

In addition to the scientific and SI accepted units given above, there are other units for these four properties. These were common in early science, and are still used in some parts of the world today. However, although it is important to recognize and be able to convert these units, we need to realize that in the scientific realm we normally only use the units above. The following tables show other units for the properties listed.

Table 3.5 – Other Units

For Mass:

1 pound (lb) = ~454 grams (g)

1 ton (t) = ~907,185 grams (g)

1 ounce (oz) = ~28 grams (g)

For Volume:

1 cubic centimeter (cc or cm^3) = 1 mL = 1 x 10^{-3} liters (L)

1 gallon (gal) = ~3.78 liters (L)

For Length:

1 inch (in) = ~0.0254 meters (m) = 2.54 cm

1 foot (ft) = ~.305 meters (m)

1 mile (mile) = ~1609 meters (m)

For Time:

1 minute (min) = 60 seconds (s)

1 hour (hr) = 3600 seconds (s)

1 day (day) = 86,400 seconds (s)

1 year (yr) = 31,536,000 seconds (s)

Now that we have seen what the units are for measurements of mass, volume, length, and time, we need to understand how to change from one unit to another. We call this process unit conversion. Imagine you have two friends who are having an argument over who ran a longer distance in a day. One friend states he ran 3 miles. Your other friend says she ran 5 kilometers. Who is right? In order to find out, we have to find a way to compare the two different units of measurement. In other words, we have to convert one of the units so that both distances are measured the same way. Let us convert the miles into kilometers. When you do conversions, it is helpful to set them up in a systematic way, so that all units are converted correctly and nothing is over looked. For this example, we need to first convert miles to meters, and then meters to kilometers. We can set this up in the following way:

$$\frac{3 \text{ miles}}{1} \quad \frac{1609 \text{ meters}}{1 \text{ mile}} \quad \frac{1 \text{ kilometer}}{1000 \text{ meters}} = 4.827 \text{ kilometers}$$

This settles the argument: the friend who ran 5 kilometers ran a longer distance than the friend who ran 3 miles. You should notice in the setup above that the dividing horizontal lines are showing equality of the values above and below them. 1 mile = 1609 meters; 1 kilometer = 1000 meters. We saw these values in the tables listed above. It is important to use a systematic method or setup like this in order to keep all the units and calculations straight. You multiply across both the top and bottom of the dividing line, and then divide the top value by the bottom value. Thus, to get the 4.827 kilometer answer, we did the following calculation: (3 x 1609 x 1)/(1 x 1 x 1000). Using this method, the mile unit on the bottom cancels the mile unit on the top, the meter units also cancel, and we are left with kilometers, which is the unit we wanted. This example may seem rudimentary, and thus this elaborate setup unnecessary, but conversions can get much more complex, and an organized method of conversion becomes crucial then. Let us look at a more complex example. Pretend you have gone to Europe and fallen in love with a new Ferrari sports car in an automobile showroom there. The sticker on the car says that the Ferrari gets 8 kilometers per liter of gasoline. Not used to seeing gas mileage in these units, how would you be able to tell how many miles to the gallon the sports car gets? Again, we can set up a systematic unit conversion and find out.

$$\frac{8\ \text{kilometers}}{1\ \text{liter}} \quad \frac{1000\ \text{meters}}{1\ \text{kilometer}} \quad \frac{1\ \text{mile}}{1609\ \text{meters}} \quad \frac{3.78\ \text{liters}}{1\ \text{gallon}} = 18.79\ \text{miles/gallon}$$

That gas mileage is not wonderful, and may affect whether or not you decide to buy the car. In this more complex conversion, we are still using the same method as we did in the last example. (8 x 1000 x 1 x 3.78)/(1 x 1 x 1609 x 1) gave us the number 18.79. When we canceled the units on the top with the units on the bottom, we were left with miles on the top and gallons down below. Thus this tells up that for every 18.79 miles the car travels, it will use one gallon of gasoline.

We have seen four different physical quantities (mass, length, volume and time) and we have learned the units used in measurements of these quantities. Further, we have seen how to convert between two different units for the same quantity. Now we shall discuss how these measurements are made. Are there magical labels on every object that tells us what the mass, volume, length and age (time) for each dimension is? Unlikely. We obtain these measurements using appropriate tools, or instruments. Some of these tools are very common to us and we use them often. Length, for example can be measured with a ruler, and a watch certainly serves to measure time. But how do we carry out other measurements? How do we measure the volume of a liquid, or the mass of a book? Scientists have devised all sorts of tools to measure various properties of matter and non-matter in our universe. We have already said that length is measured with a ruler. What does a ruler look like? This may seem like a simple question, but really think about it. Rulers measure length because they are straight edges. And we defined length earlier as being the linear extent in space from one end of an object to the other, along the longest dimension. We choose the straight edge of the ruler to measure length, since the shortest distance between any two points in space will always be a straight line between them. Now look at the surface of the ruler. It has markings on it, right? These markings are called *graduations*. The markings systematically increase and are all the same distance apart. The space between the 3 inch mark and the 4 inch mark is exactly the same as the distance between the 6 inch mark and the 7 inch mark. Because volume is the space an object takes up in space, we can use a ruler to measure volumes of some structures. Consider a box, or any rectangular prism. If we measure its height, width and depth in meters, then multiply these three dimensions by each other, we obtain the volume in cubic meters. Cylinders, spheres, cones, and other perfect geometric shapes can all have their volumes measured in this way. But what about objects that have odd shapes? How would you measure the volume of your sunglasses? Or a handful of sand? How about the volume of water in a puddle in the street? Volume can also be measured in liters. Unlike cubic length volumes which can use the ruler for measurement, liter volumes must use a different tool. Scientists use a tool called a graduated cylinder to measure volumes. A picture of a graduated cylinder is given in Figure 3.1.

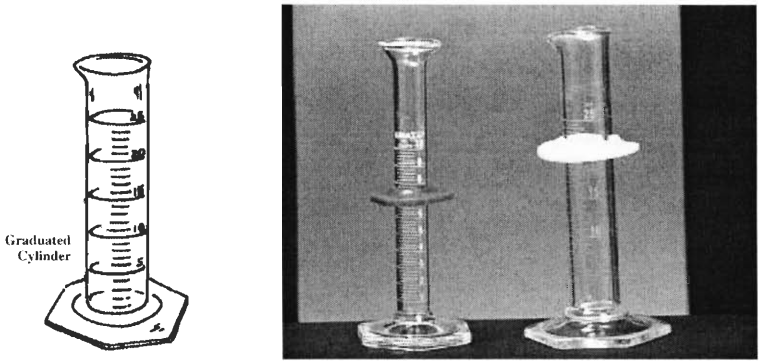


Figure 3.1 – Graduated Cylinders

Just like the ruler, graduated cylinders have markings (graduations) on them indicating the measurement for the object in question. If you wanted to know how much liquid was in your cup of tea, you could pour the contents of the cup into the cylinder and read the volume by finding the graduation mark at the liquid surface. If you wanted to find the volume of your sunglasses, you could read a volume of water, submerge your sunglasses in the water, find the new volume, and then take the difference to find the volume of just the sunglasses. It is important to know the correct way for reading a graduated cylinder, however. Because of the nature of some liquids, they often form a curve on the surface within the cylinder. This is illustrated below in figure 3.2.

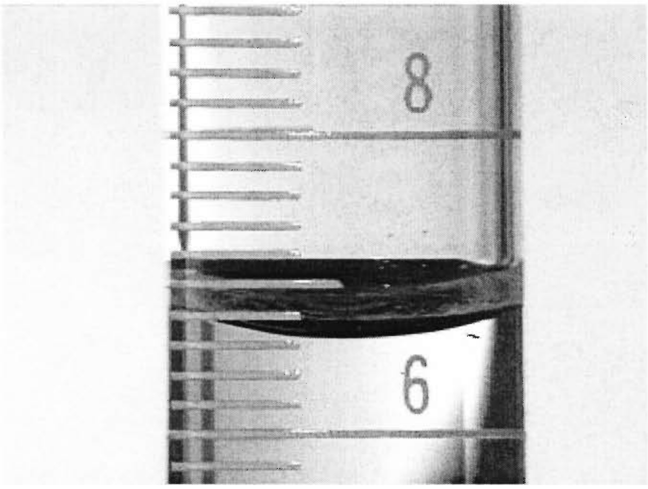


Figure 3.2 – Close up of the Meniscus

The curve shown above is called a *meniscus.* The fact that the liquid forms a curved rather than flat surface complicates the measurement of volume. By convention, chemists have agreed to look at the bottom-most point of the meniscus, and see where it lies. For this example, the 6 and 8 numbers printed on the cylinder represent milliliters. The lines between the 6 and 8 each correspond to 0.2 milliliters. The bottom of meniscus lies between the third and fourth lines up from 6 ml. So we know that the volume of this liquid is between 6.6 ml and 6.8 ml. For the last digit, we have to estimate. The meniscus is about a quarter of the way from the 0.6 ml line to the 0.8 ml line, so we could say that the volume of this liquid is 6.62 ml.

Unlike volume and length, mass and time have completely different tools to measure them. Mass is measured using an instrument called a balance, which most of you probably have in your bathroom. Although usually you think of a balance as measuring weight, remember we discussed that weight is a certain amount of gravity affecting a mass. If you took your bathroom scale to the moon, it would read a different weight. Thus mass is truly measured by comparing a reference mass to the mass being measured. The important thing to remember is that the weight of an object can change with location, and mass cannot. Time is a little more abstract, since the concept of time requires organisms to observe it. However, without getting into the philosophy of time, most often in chemistry time is measured with a stopwatch, or, for quicker processes, a more precise electronic device that can "clock" times shorter than are possible using the human eye and reflexes.

The idea of a "more precise electronic device" raises an interesting point. What do we mean by the word precise? Remember how we learned to read a graduated cylinder. We estimate the last digit of the measurement according to where the meniscus lies between the graduation lines. We do this for rulers and lengths, and for balances and masses. Even electronic machinery that gives us a time or mass measurement estimates the last digit of the number. This brings us to two important concepts, *uncertainty* and *precision*. You can think of uncertainty as the amount you had to estimate. For instance, when we read the graduated cylinder, we knew it was between 6.6 and 6.8ml. We could even further be fairly certain that the volume was between 6.60 and 6.70 ml. But in saying 6.62 ml, are we absolutely certain that that is the correct volume? Could it have been 6.63ml? How about 6.61ml? This leads us to think of any measured quantities with a certain level of uncertainty. And it makes us question the precision of the instrument. The uncertainty here is in the last digit, the 2, since we had to guess at that number. The precision lies in the 6.6 number, since we know without doubt that the volume is at least 6.6 mL. Uncertainty leads us to second guess our estimation, and thus prompts us to write a correct measurement reflecting this uncertainty. This is done by adding the uncertainty to the end of the measured value. In this example, we would say the volume of the liquid was 6.62 ± 0.02 since the volume could have been estimated anywhere from 6.60 to 6.64 ml. The number 0.02 is the uncertainty of the measurement. It measures how "sure" we are that the last digit of our measured value is correct. Generally, the lower this number, the more confident we are in the measurement. Also, we tend to think we have a higher precision if repeated measurements of the same object yield the same or very nearly the same numbers. For large objects this fine point does not seem to make a lot of difference. However, when measuring such small things as atoms and distances between them, this uncertainty can really add up! Even some large scale problems, such as the construction of bridges, can be seriously affected by the uncertainty in measurements.

We just said that the lower the number of uncertainty, the more certain we are of precision in its measurement. Thus, if we added decimal places, the number would decrease, and we could think of the measurement as being more precise. What if we had said "The meniscus is between the 6 ml line and the 8 ml line. It's closer to 6ml, so the volume is 6 ml." That doesn't seem very precise. We can then further increase our precision by adding another digit. "Well, I know it's between the 0.6 and 0.8ml lines, and closer to the 0.6ml line, so the volume is 6.6ml." But is that as far as we could go? No, a further digit was added to give the volume even more precision. All of these digits are known as significant figures. Every digit of a power of 10 that we add is another significant figure added. We started with one significant figure – 6 ml. We then added another – 6.6 ml has two significant figures. Our final answer, 6.62 ml has three significant figures. There are three rules for determining how many significant figures are in a number:

1. Non-zero digits are always significant.
2. Any zeros between two significant digits are significant.
3. A final zero or trailing zeros are only significant if they are marked with a decimal point.

Please remember that, in science, all numbers are based upon measurements (except for a very few that are defined). Since all measurements are uncertain, we must only use those numbers that are meaningful. We could not have measured our volume in that specific cylinder shown to be 6.625872649ml. Not all of the digits have meaning (significance) and, therefore, should not be written down. In science, only the numbers that have significance (derived from measurement) are written.

Now that we know what digits in a number are and are not significant, we need to understand that calculations involving numbers with significant digits are also governed by rules. There are two general rules for managing significant figures in calculations. In multiplication and division, we round the final result to the number of significant figures in the least well-known factor. For example:

$$\frac{(15.03)(4.87)}{1.987} = 36.8$$

The answer, 36.8, is rounded to three significant figures, because the least number of significant figures was found in the term, 4.87. The other terms, 15.03 and 1.987, each have 4 significant figures. In addition and subtraction, we round the final result to the number of decimal places in the term with the fewest digits to the right of the decimal, regardless of the significant figures of any one term. For example:

$$\begin{array}{r l} 1.003 & \\ 13.45 & \\ +\ \ 0.0057 & \\ \hline 14.4587 & \text{rounds off to} \quad 14.46 \end{array}$$

The answer, 14.4587, was rounded to two decimal places, since the least number of decimal places found in the given terms was 2 (in the term, 13.45). The term 1.003 has 3 figures after the decimal, while 0.0057 has 4. As we will see, these concepts of precision, uncertainty and significant figures are essential in laboratory work.

Section B – Atomic Society

We took a qualitative look at the atom in Chapter 2. We now need to take a closer look at some of the properties of the atom, and at how atoms combine with other atoms to make molecules. We will look at the atom both theoretically and from an experimental view, with the use of mass spectroscopy. Hopefully this section will allow us to better quantify the atom and molecule.

We spoke in the last section of Chapter 2 about the periodic table and saw that the table was arranged by two criteria, similar physical properties and atomic mass. But what is atomic mass? Atomic mass is the mass of an atom in its neutral state. This mass is measured in atomic mass units (μ) and is approximately equal to the sum of the number of protons and neutrons in the nucleus of the atom (remember we showed that the mass of the electron was so small compared to those of protons and neutrons that we need only consider the nuclear subatomic particles when estimating atomic mass). An atomic mass unit is 1.67×10^{-27} kg. Typically, however, when we speak of atomic mass, we use atomic mass units rather than kilograms because the numbers are easier to work with. For example, carbon has an atomic mass of approximately 12.011 μ, in other words, 2.01 x 10^{-26} kg. When working with either of these numbers in calculations, the atomic mass unit number is preferred over the kilogram. In addition, the kilogram value is so small, often it is rounded to zero by ordinary computer programs and calculators, so the atomic mass unit is preferred for that reason as well.

We just saw that the atomic mass in μ is approximately equal to the number of protons and neutrons. In our example above, carbon, the atomic mass is 12.011. However, when we think of "normal" carbon, we think of it as having 6 protons and 6 neutrons. That should give it an atomic number of about 12μ. Do the 6 electrons in the molecule as well really account for 0.011 u of the atomic mass? Is there something else that could be causing this mass other than the electrons or the fact that a proton and neutron don't weight exactly 1 μ? We just stated that "normal" carbon has 6 protons and 6 neutrons. If we increase or decrease the number of protons, what happens? We get a different atom, right? But what would happen if we increased or decreased the number of neutrons in the carbon nucleus? Changing the number of neutrons in an atom results in a different form of the same atom, called an *isotope.* Isotopes of a given element have the same atomic number (same number of protons in their nuclei) but different mass numbers (different number of neutrons in their nuclei). Mass number is the total number of protons and neutrons in the nucleus. Thus, carbon in its most abundant, or "normal" state is known as carbon-12. The 12 represents mass number; for carbon, 6 protons plus 6 neutrons. However, carbon can also exist as carbon-13 or carbon-14. This means that the number of neutrons in the nucleus has increased. Carbon-13 has 7 neutrons, and carbon-14 has 8 neutrons. The number of protons cannot change, since this distinguishes one atom from another. The average mass of these isotopes generate the true atomic mass of an element. The most common isotope obviously influences the atomic mass of the element the most; thus, the atomic mass rounded to the nearest whole number is the most common isotope of each element. For example, oxygen has an atomic mass of 15.9994 μ. So the most common form of oxygen is oxygen-16. Isotopes are most useful in studying radioactive decay, and the isotopes of carbon have been used to date prehistoric material on our planet in order to figure out evolutionary trends in organisms and formation and motion of the landmasses of the earth.

Now that we have taken a look at the atom itself in more depth, and can quantify atoms now, we need to be able to describe and quantify groups of atoms, molecules. As we saw in chapter 2, each atom has its own symbol; we saw that carbon was represented with the symbol C. Molecules, since they are made of groups of atoms, are given special symbols as well. Rather than giving every molecule its own symbol (that would be a LOT of symbols!), we represent molecules using *formulas*. Formulas are a representation of the elemental composition of a molecule, and use numerical subscripts to indicate the relative numbers of atoms of each kind of element present. What this means is for each molecule, we give its formula by writing down the symbols for the elements in the molecule, and add numbers to the elemental symbols if there is more than one atom of that element within the molecule. For example, we have already looked at water. A water molecule is made of one oxygen atom and two hydrogen atoms. Thus, the molecular formula for water is H_2O. Consider the formula $C_6H_{12}O_6$, which is the formula for glucose, or sugar. What can we say about the composition of this molecule? Well, we know that this molecule is made of carbon, oxygen and hydrogen, and we also know that there are 6 carbon atoms, 12 hydrogen atoms and 6 oxygen atoms in every molecule of this compound. Using the technique, we not only have a systematic method for representing a molecule, but we can also infer something about the make-up of the molecule (and sometimes even its structure) by the formula. It is important to realize that we are speaking of the chemical, or molecular formula of a specific molecule. There is another kind of formula that molecules have, and that is an *empirical formula.* These are the “simplest” formulas for a compound. They show only the ratio of the number of atoms of each element in the compound. Empirical and molecular formulas may be the same for some molecules, and different for others. For example, water has a molecular formula of H_2O, and it has the same empirical formula, since H_2O shows the simplest ratio of atoms in the molecule. However, for sugar, $C_6H_{12}O_6$, the molecular formula is not the same as the empirical formula. The empirical formula for sugar is CH_2O. This shows the ratios of the number of atoms, but it does not tell us the actual number of atoms of each element like the molecular formula does. Although empirical formulas are not often used, they are important when looking at families of molecules that exhibit similar behavior.

We now know how to properly represent a molecule, showing which atoms it contains and the number of atoms of each element within the molecule. But what else can we say about a molecule from its formula? Having learned about formulas and atomic mass allows us to derive another piece of information about the molecule, percent composition. Percent composition is an easy concept; it is simply the percentage of each element within a compound defined by the mass each element contributes to the total mass of the molecule. We find this numerical value by calculating the mass of an element in the molecule and dividing it by the total mass of the molecule. In order to calculate percent composition, we have to first be able to calculate molecular mass. For example, we found water to have a formula of H_2O, which means it has 2 hydrogen atoms and one oxygen atom. If we look to our periodic table in the front cover of this book, we see that hydrogen has an atomic mass of $1.0079\ \mu$ and oxygen has an atomic mass of $15.9994\ \mu$. Thus, the molecular mass for water is $(1.0079 + 1.0079 + 15.9994) =$

18.0152 μ. For sugar, we have carbon (atomic mass 12.011), hydrogen (mass 1.0079) and oxygen (mass 15.9994). So the molecular mass of sugar is (6 x 12.011) + (12 x 1.0079) + (6 x 15.9994), since we have 6, 12 and 6 atoms of carbon, hydrogen and oxygen, respectively. This totals 180.1578 μ, showing that it is much larger than a water molecule. Now all we have to do to find percent composition is to divide the part by the whole (we also multiply the answer by 100 to get a percentage rather than leaving it in decimal form). Thus, the percent composition of hydrogen in water is (2.0158/18.0152) x 100, or 11.19%. The percent composition of oxygen in water is (15.9994/18.0152) x 100, or 88.81%. The sum of the percentages of all the elements in a molecule must equal 100%. This way, none of the molecular mass is going unaccounted for, and it serves as a way of checking your calculations for error (if they did not add up to 100%, you may have miscalculated). For water this percent composition was fairly easy to execute. But for more complex molecules the following set-up may make it easier to find both the total molecular mass and the percent composition of each molecule. We will use our sugar example. You start by listing each type of atom vertically in a column. Then next to each atom, list the quantity of them in the molecule. Next, multiply that number by the atomic mass. The totals on each line sum up. This set up is given below.

C	6	x	12.011	=	72.0666
H	12	x	1.0079	=	12.0948
O	6	x	15.9994	=	96.9964
			Total	=	180.1578

We can then take the total and divide each line by it, giving us the percentage of each element in the molecule.

C	6	x 12.011	=	72.0666	/180.1578	=	.400 or 40.0%
H	12	x 1.0079	=	12.0948	/180.1578	=	.067 or 6.7%
O	6	x 15.9994	=	96.9964	/180.1578	=	.533 or 53.3%
		Total	=	180.1578			Total % = 100%

Using this method helps to keep the arithmetic straight for larger, more complex molecules.

We have just seen how to calculate the molecular mass of a molecule. But what did we need in order to accomplish this? We needed the formula of a molecule. What if the formula is not known? How should we find the molecular mass? There are many ways of doing this. For example, we could send the substance to a processing facility that would analyze the molecule and tell us its percent composition. We could then work our above process backwards in order to find the formula of the molecule. But this procedure is both costly and time consuming. Another way to find the molecular formula is to run the molecule through a special device called a mass spectrometer. A mass spectrometer produces charged particles, or ions, from the chemical substances that are to be analyzed. The mass spectrometer then uses electric and magnetic fields to measure the mass of the charged particles. The instrument produces a printout like the one below in figure 3.3.

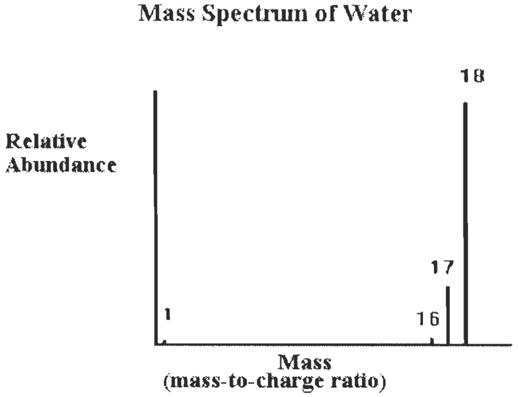


Figure 3.3 – Mass spectrum of Water

Here we see 4 signals resulting from the different fragments of the molecule. This spectrum is like a fingerprint for this molecule, since different molecules fragment in different ways. The largest peak in the mass spectrum is called the "parent peak" and is often given the symbol M+. The parent peak tells us the approximate molecular mass of the molecule sampled. In the spectrum above, we can see the parent peak for water is at 18, which coincides with what we calculated earlier. Scientists use mass spectroscopy for all sorts of purposes, ranging from environmental analysis (for example, the presence of a poison in the atmosphere) to the analysis of petroleum products, trace metals and biological materials. In the next section, we will see how scientists use this and other tools to discover properties of the world around us and solve problems in the laboratory.

Section C – Putting on the White Coat

We learned in the last two sections about how to read equipment in the laboratory, and also the concepts of uncertainty, precision and accuracy for each reading. But let us take the definitions one step further, and look at an actual problem we can solve in the laboratory, and how each term applies to the experiment. This chapter will focus on these practical applications of the concepts learned in the previous sections.

Suppose you and a friend walk into the laboratory and the instructor hands you a small piece of "pure" metal (in other words, a single element, not a mixture or compounded substance). S/he tells you that you need to find the identity of the metal and leaves. You are given an electronic balance, a beaker of water, a graduated cylinder, the periodic table and a calculator. How will you find out the identity of the metal?

Although there is more than one sequence of steps you could take to solve this problem, the overall solution is the same. We need to focus on one physical property of this metal that we can measure with the equipment available. In this case, since we can measure mass and volume, we should look at the *density*. Density is equal to the mass of the object divided by the volume of the object. You need to find the mass of the metal, and find the volume of the metal, and do a few simple calculations to find the density. Once you have the density, you can look on the periodic table to see which metal it is most likely to be. Identifying the problem, assessing how you can solve it, then taking the steps to do so is the method generally used by scientists and is described in more detail in Chapter 1.B.

Let us start with finding the mass of the metal piece. To do this, we place the metal on the balance, and read the number it gives on its digital display. The balance reads 27.32 grams. But remember that even electronic equipment estimates the last decimal place. So we really do not know if the mass is 27.30 grams or 27.34 grams. Thus we should correctly write the mass for the metal as 27.32 ± 0.02 grams. This tells us that we have an uncertainty of 0.02 for this measured value.

Now suppose we want to find the volume of the metal. Since it is an odd shape, and we do not have a ruler, we cannot take a cubic volume measurement for this object. We do, however, have the graduated cylinder and some water. Thus, we can find the volume by the amount of water displacement the object produces. Pour some water into the cylinder and find its volume. Let us say the starting volume was 14.36 ml (the graduated cylinder had 0.1 ml marks as its smallest graduation). When we drop the piece of metal into the cylinder, the new volume you read is 17.87 ml. But remember for both of these, we estimate the last decimal place. So again, we have an uncertainty of 0.02 for each volume reading. So the final volume is 3.51 ± 0.002 ml.

Now we divide the mass by the volume, making sure to observe significant figure rules and uncertainty. This means we have to find the mean density, in other words 27.32g/3.51ml, as well as the range of density (the lowest and highest possible densities taking the uncertainty into account). The mean density comes out to 7.783 ± 0.02 g/ml.

Notice that although the significant figures allow for three places after the decimal place, the uncertainty limits it to two places. We should thus write the mean density as 7.78 ± 0.02 g/ml. No we have to find the range of density. Thus one range limit is 27.30/3.53, or 7.73 ± 0.02 g/ml and the other range limit is 27.34/3.49, or 7.83 ± 0.02 g/ml. Ah ha! Look at what the possible range of densities has turned up using the uncertainty in the measured value. Although our uncertainty for the mass and volume was 0.02, the density can vary in the decimal of 0.1 g/ml. Thus, we should conclude that the density of the metal is 7.7 ± 0.1 g/ml. This is as certain as we can be that our value is correct. We can see now how uncertainty can really add up! If you scan the metals on our periodic table, you find that the closest metal to the density you calculated is 7.874 g/ml, which is the density for iron. You can now tell your instructor you know what the metal is, and have learned a way of identifying unknown substances! Since we know that our value is 7.7 ± 0.1 g/ml, and the "accepted" value is 7.874, we can find the error in our results. To do this, we subtract the experimental value from the theoretical value, then divide that by the theoretical value and multiply by 100 to find the percent error. Figure 3.4 illustrates this more clearly.

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

Figure 3.4 – Finding percent error

So in our example, [(7.874 – 7.7)/7.874] x 100 comes out to 2.2% of error in our results. This error can be attributed to human or machine error in reading the mass and volume measurements. It can also be reduced by using equipment that had smaller graduations or a balance that read to 4 places after the decimal instead of 2. That would reduce our uncertainty for the mass and volume, which would reduce the uncertainty in our final measured density, allowing us to report that density with more significant figures. We could have also tried reducing error by measuring the mass and volume several times each, thus getting a mean value for them, and finding the precision of those values. Many different methods are employed to reduce error, and generally, error of 5% or less is considered acceptable. More than 5% and either the experiment was done incorrectly, or you picked the wrong identity for the metal. It is also important to understand percent error is only considered when the theoretical values are already known. Obviously it would have been hard for Plunkett to calculate the percent error for the density of Teflon, since it was never before seen or measured.

Chapter 4 – You Are One Attractive Element!

Section A – Let's Have a Bonding Experience

A large portion of chemistry is the study of how atoms and molecules interact with each other. These interactions can come in the form of repulsions or attractions. In this section we will take a look at the attractions between atoms and see how to predict those attractions.

As we said above, atoms can interact via attraction to one another. Whether between atoms or molecules, any attraction between two substances that leads to an attachment between the two is called a *bond.* There are many different kinds of bonds, from physical to chemical. And within these main categories there are many subcategories. For now there are three categories of chemical bonds that we will look at.

ionic
covalent
metallic

These types of bonds cover nearly all of the strong attachments you could describe between atoms of any two given substances. Let us now take a look at how these differ from one another and some simple ways of determining which type of bond will form between two particular atoms.

The easiest way to determine the type of bond to expect between 2 atoms is to look at the periodic table. Many starter periodic tables have a "zig-zag" line drawn across them, like figure 4.1 below.

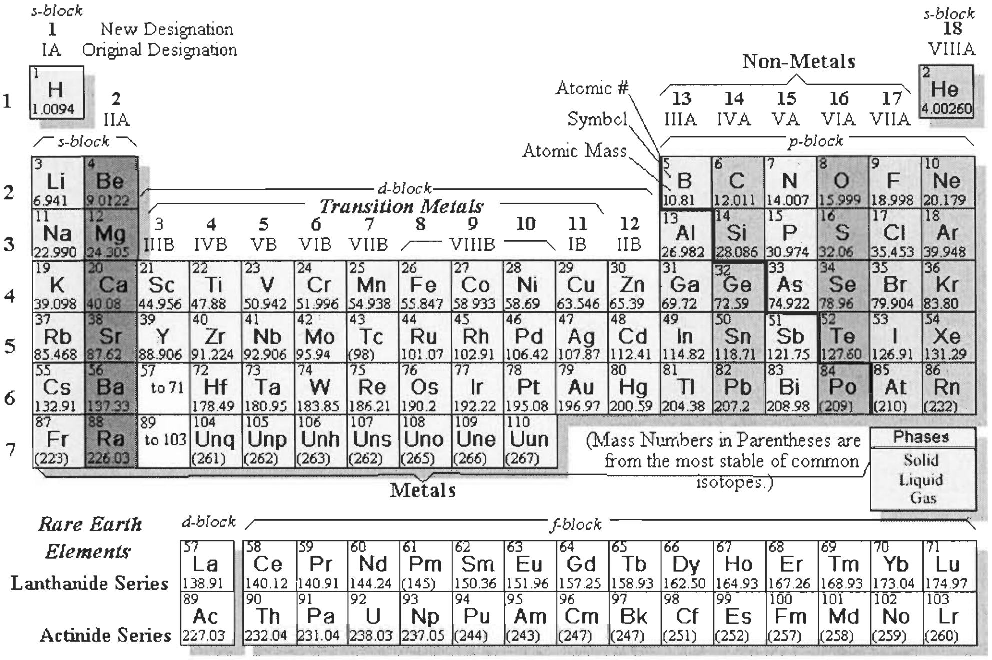


Figure 4.1 – Periodic Table

As you can see above, the "zig-zag" line starts between boron (B) and aluminum (Al) and follows a staircase pattern down the table. This is going to be the dividing line between the "left" and "right" sides of the table. As a general rule, the attachment between 2 atoms from the left side of the table is called a *metallic* bond. Any atom from the left side of the table that forms an attachment with an atom from the right side is said to make an *ionic* bond. Finally, 2 atoms from the right are said to form a *covalent* bond. While all bonding types are found in nature, let us take a closer look at only the ionic and covalent bonding types for now. We will discuss metallic bonds in chapter X.

Ionic bonds are formed when an atom from the left side of the line gives electrons to an atom on the right side of the line. We refer to this process as an electron transfer. Both atoms change their numbers of electrons in this process, and so are no longer neutral. An atom with its "normal" number of electrons carries no charge since the positive charges from the protons balance with the negative charges from the electrons. But if an atom gains or loses electrons, now the charges cannot balance. When an atom becomes charged via a gain or loss of electrons, it is said to have become an *ion*. An ion is simply an atom or molecule that has gained charge through the loss or gain of electrons. In the case of our atoms from the left and right sides of the line, the atom from the left gives up its electrons, so it becomes a positive ion, also known as a *cation*. The atom on the right side receives the electrons, so it becomes a negative ion, also known as an *anion*. The ionic bond is the strong electrical attraction between the cations and anions that result from electron transfer.

But how do we know how many electrons an atom gives up? Once again we turn to our periodic table for predicting such things. If you look back to figure 4.1, you should notice that the columns are numbered 1-18. The column numbers can be used to predict how many electrons an atom will either give or receive. For example, an atom in column 1 can donate 1 electron. An atom in column 2 donates 2 electrons. And an atom from column 13 donates 3 (for the meantime we will not look at column 3-12 since they can vary on the number of electrons they can donate). However, once we get to column 14, the column really is split by the line. Atoms in column 14 can give or receive 4 electrons. Then, we start to receive instead of donate electrons. Atoms in column 15 can receive 3 electrons, atoms in column 16 receive 2 electrons, and atoms from column 17 can receive 1 electron. Remember when we discussed the periodic table in chapter 2 we found the last column, number 18, was the inert or Noble gases. They neither donate nor receive electrons readily, which makes it difficult for them to form bonds, thus they are unreactive. Knowing this, we can predict what kinds of charges atoms will have, and how they might bond. For example, how might an atom of sodium bond to an atom of chlorine? Sodium is in the first column, so it would donate 1 electron and become a +1 cation. The symbol for this cation is Na^+. Chlorine is in column 17, and receives one electron, so chlorine would become a -1 anion. The symbol for this anion is Cl^-. Since you have a +1 and -1 we put the two together and get a neutral molecule, sodium chloride, or NaCl (which is the formula for everyday table salt). Let us try another example. Say we take an atom of calcium and an atom of bromine. Calcium is in column 2 so it donates 2 electrons and becomes a +2 cation. Bromine is in the same column as chlorine so it receives 1 electron and becomes a -1 anion. But if we put these two together, we would get a charged molecule, because the charges don't balance, and we would have an electron left over! However, if we instead put one atom of calcium and 2 atoms of bromine together, now the charges will balance and the molecule becomes neutral. We get calcium bromide, $CaBr_2$. This business of balancing charges so we can write the formula for a neutral molecule is one aspect of what is known as *stoichiometry*.

Now that we know how to predict the formulas for some ionic compounds, let us turn to covalent bonds. Covalent bonds are different from ionic bonds in one major way. In covalent compounds, atoms form attachments by sharing electrons rather than transferring them. Because these atoms to the right can sometimes share electrons in different ways, for the time being, we will just look at some common, or proposed bonding numbers for some of these elements. What we mean by that is how many bonds does an atom of each element normally form? The elements we will consider are carbon, nitrogen, oxygen, the halogens (column 17 of figure 4.1), sulfur and hydrogen. Now you may wonder, "Hydrogen? But isn't hydrogen on the left side of the line?" Yes, hydrogen is on the left side of the line. But hydrogen, as usual, is an exception to the rule. Hydrogen bonds covalently with atoms to the right of the line, not ionically. It forms ionic bonds with atoms on the left side of the line as well. Without going into depth about why these elements bond as such, let us just look at the observed bond numbers for some of these atoms in table 4.1.

Element	Bond Number
Carbon	4
Nitrogen	3
Oxygen	2
Halogen	1
Sulfur	2
Hydrogen	1

Table 4.1 – Proposed covalent bond numbers for some common elements

What these numbers mean is that for each element, the number is how many covalent bonds it needs to form to be stable (without creating any ions or having any electrons left over). For example, let us suppose we want a molecule made of one carbon and some hydrogen. Carbon needs 4 bonds to be neutral, and hydrogen only needs 1. Thus, if we had one carbon bonded to 4 hydrogens, we would get a neutral molecule (and in fact we do, CH_4 is the formula for methane, the simplest hydrocarbon). Let us try another example. Say we want a compound that has 2 nitrogens in it and some oxygen. We know that each nitrogen needs 3 bonds, and we have 2 of them, so we need 6 total bonds from oxygen. One oxygen atom cannot make 6 bonds. But what about 3 oxygen atoms? Three oxygens at 2 bonds each would make 6 bonds, what we need for nitrogen. Thus we make dinitrogen trioxide, or N_2O_3. Using these suggested bond numbers is a quick and easy method for predicting formulas for covalent compounds.

Section B – Stoichiometry from Percent Composition

In chapter 3 we looked at percent composition and how to find it from the formula of a compound. And we discussed the possibility of working the other way around. So what do we do when we discover a new compound? How do we go about finding its formula?

When a new compound is discovered, among the numerous tests it undergoes is a test of percent composition. The compound is run through a device that identifies the types and percentages of each element in the compound. So let us try an example. Suppose you think you have grown a mutant orange tree in your basement. You think that the fruit of this plant might have the cure for baldness in it. So you isolate some of the chemicals from the fruit and ship them off for testing. You receive a letter back from the testing company that says the following:

Analysis of Compound 28423-A
C = 41.0 %
H = 4.5 %
O = 54.5 %

Where do you go from here? How should you find the formula of this compound? Have you found the cure for baldness or is this a known compound?

In order to find the simplest formula for this compound, we work backwards from what we did in chapter 3.B. There, our first step in finding the composition was to multiply by the atomic mass. Here, our first step in finding the formula will be to divide each percent by the atomic mass. Thus, we have:

C 41.0 / 12.011 = 3.41
H 4.5 / 1.0079 = 4.46
O 54.5 / 15.9994 = 3.41

These numbers are the numbers of moles in a hypothetical 100-gram sample of this compound. But we do not need to worry about that. The next step is to divide each number by the smallest number, in order to find a ratio of the atoms to each other. Thus:

C 3.41 / 3.41 = 1.00
H 4.53 / 3.41 = 1.33
O 3.41 / 3.41 = 1.00

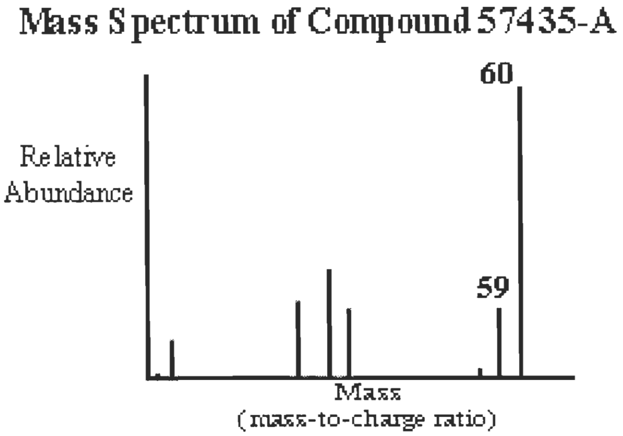
These ratios tell us that for every one carbon atom we have in this molecule, we also have one oxygen atom. It also says that for every one carbon and oxygen atom, there are 1.33 hydrogen atoms in the molecule. Obviously we do not write a formula $CH_{1.33}O$, so when decimals are obtained in the ratio, we multiply all the ratios by a factor that will provide all whole numbers. In our case above, we recognize that 1.33 can be written as 4/3. If we thus multiply all the ratios by 3, we get the formula $C_3H_4O_3$. Unfortunately all your "mutant" orange has produced is a healthy dose of vitamin C.

Section C – A Clue Sherlock! From Data to Structure

Now that we have seen how some elements bond with each other, and how to convert mass data into formulas, let us go a little more in-depth, and take an example of a real life problem.

Suppose your "mutant" orange from the last section suddenly grows a purple fruit. Now a little more confident you might have something new, you isolate a compound and send it off to be analyzed, this time requesting both mass composition and a mass spectrum of the compound in question. The testing company returns a letter with the following information:

Analysis of Compound 57435-A
C = 40.0 %
H = 13.4 %
N = 46.6 %



How should you utilize this data to get the structure of your new compound? Well, you should look at the type of data you have. You have received mass percent composition data, so you can find the empirical formula for the compound. But you have also received a mass spectrum of the compound. From that you can find the molar mass of the compound and thus its molecular formula. With the formula, you can use the preferred bonding numbers proposed in the last section to suggest some possible structures for this compound. So let us begin.

The first step is to find the empirical formula from the percent composition data. You do this very much the same way as you did in the last section. Remember that your first step in finding the formula is to divide each percent by the atomic mass. Thus, you have:

C 40.0 / 12.011 = 3.33
H 13.4 / 1.0079 = 13.29
N 46.6 / 14.0067 = 3.32

Your next step was then to divide each number by the smallest number, in order to find the number ratios of the atoms. Thus:

C 3.33 / 3.32 = 1.00
H 13.29 / 3.32 = 4.00
N 3.32 / 3.32 = 1.00

Wonderful, the ratios are all whole numbers. These ratios tell you that for every one carbon atom you have in this molecule, you also have one nitrogen atom and 4 hydrogen atoms.

Now that you have this empirical formula, you should calculate the formula mass and see if it makes sense in conjunction with the mass spectrum data. Remember from chapter 3 that we find the formula mass of a compound by multiplying the atomic mass of each element in the compound by the number of atoms of that element in the compound, and then summing those up.

C	1	x	12.011	=	12.011
H	4	x	1.0079	=	4.0316
N	1	x	14.0067	=	14.0067
			Total	=	30.0493

This tells us that the formula mass of the empirical formula is about 30.05 g/mol. But wait a minute. Take a look at the mass spectrum again. Find the parent peak. The mass spectrum of your compound says that this molecule has a molar mass of about 60, and yet you calculated a mass of about 30. Did you do the calculation wrong? Absolutely not. This displays the difference in empirical and molecular formulas as discussed in chapter 3. Since the formula mass is twice that of the calculated "empirical weight," we simply increase all the numbers of atoms in the empirical formula by 2 to find the correct molecular formula. Thus, the molecular formula for this compound is $C_2H_8N_2$. Does that work out to be the right formula mass as given by the mass spectrum? Let us do a quick check.

C	2	x	12.011	=	24.022
H	8	x	1.0079	=	8.0632
N	2	x	14.0067	=	28.0134
			Total	=	60.0986

The calculated formula mass matches that of the mass spectrum, so now you can be confident in the molecular formula, and begin proposing structures for the molecule.

Remember that carbon tends to form 4 bonds, nitrogen 3 bonds, and hydrogen 1 bond. Bearing this in mind, you can draw bonding units like the ones you see below in figure 4.2.

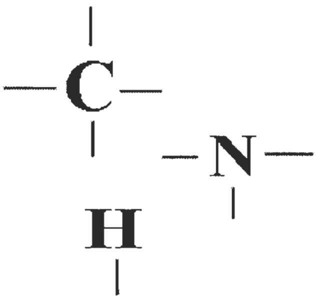


Figure 4.2 – Examples of bonding units

Taking these units, you can build structures by overlapping the lines with one another, making sure the no line is left "hanging." In other words all lines must be overlapped so that there are only lines between the atoms. Doing that, you could propose a structure like the one below in figure 4.3.

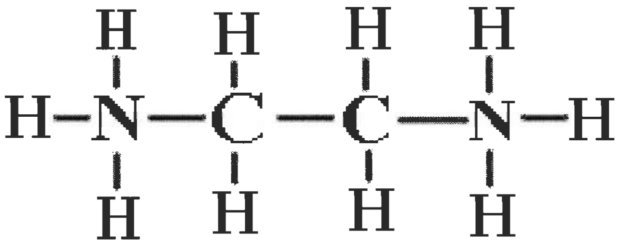


Figure 4.3 – Proposed structure

Does this structure above look like the one you drew? Maybe it does. If it does not, there is no need to feel that you are wrong, because obviously there is more than one way of overlapping these bonding units. When molecules have the same molecular formula, but different atom arrangements, the different arrangements are said to be *isomers* of each other. Figure 4.4 below shows other possible isomers that you could have drawn for your compound.

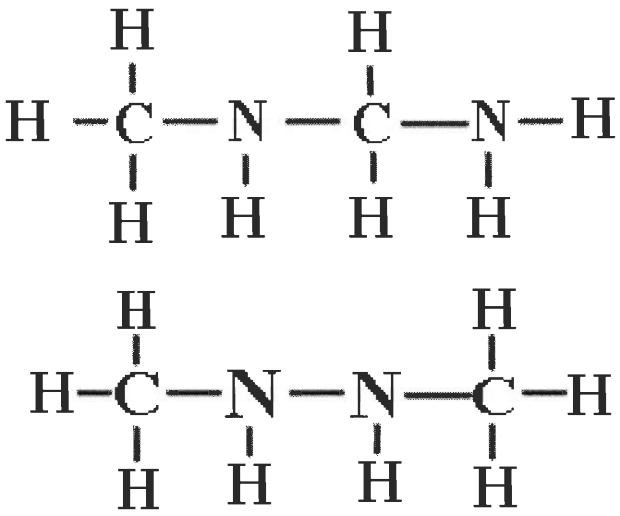


Figure 4.4 – Isomers of the proposed structure

Now we have seen how to go from simple mass data of an unknown compound to a possible structure for the compound. This method of detective work in using the “clues” you are given to solve the problem is used frequently in chemistry. This process is sometimes responsible for how discoveries are made, and is an invaluable skill to develop.

Chapter 5 – Why Don't We Get Together and React!

Section A – Did You Just Hit Me?

Now that we have taken a look at bonding and how atoms and molecules form bonds, we can now look at the process by which they get there. In this chapter we will focus on chemical reactions. Chemical reactions are processes or pathways in which atoms or molecules come together to bond. If we could think of bonding as the "what" (what atom is bonding to what other atom), then we can think of reactions as the "how" (how does that atom get to and bond with that atom).

Chemical reactions are happening constantly in our world. Everything from lighting a match or starting your car to eating dinner and breathing involve chemical reactions. With so much of life centered around reactions, it is no surprise that there is a huge diversity of reaction types that chemists have categorized. But one thing to bear in mind is that all reactions have a common thread between them. It is very easy to quickly get bogged down in learning reaction type after reaction type. But the heart of a reaction is the collision. Collisions between atoms and molecules (or even collisions between parts of the same molecule) are responsible for nearly all the chemical reactions we know of. Atoms and molecules are constantly in motion; our world is not a static one. As these molecules are bouncing around they collide with the sides of their container, but they also collide with other atoms and molecules. The reason we have to add the compounds together in the same flask is so they may interact and thus react. And obviously anything we can do in order to increase the number of collisions will also increase how much and how quickly the atoms and molecules will react with each other or themselves. Have you ever wondered why we heat our food? Or why bread has to be baked? As we will look at shortly, heating atoms and molecules makes them "move" faster. The faster they move, the more collisions take place, causing the reaction to happen faster. You could probably bake bread outside on a very hot day, but it would take you a VERY long time, longer than we have sunlight for.

Now that we know the basic concept behind what drives a chemical reaction, we should take a look at how to write a reaction. In the last chapter we spoke of an ionic bond forming between sodium metal and chlorine. When we speak of reactions, the two or more atoms or molecules that we start with we refer to as the *reactants*, and we write them towards the left. The molecule(s) we get as a result of the bonding rearrangement are called the *products*, and we write them towards the right. We separate the left side from the right by means of an arrow, which symbolizes the progress from the reactants to the products. This is illustrated more clearly below.

$$\underset{\text{Reactants}}{Na + Cl_2} \quad \rightarrow \quad \underset{\text{Products}}{NaCl}$$

So now we have written a reaction for the bond forming between sodium and chlorine to get sodium chloride. But wait a minute, there is another thing to consider when we write reactions. Remember when you write a molecular formula, that all the atoms and ion charges had to balance, in other words, we had to worry about the stoichiometry of the new molecular formula made. We also have to worry about the stoichiometry for reactions. Since the conservation law of matter states that matter cannot be created or destroyed, we must end up with the same number of atoms of each type as we started with. This conservation law mentioned in Chapter 1 is what governs stoichiometry. You cannot create something from nothing or simply have atoms and molecules disappear. The sides must always balance. Having said that, let us try to balance our sodium chloride reaction. Both sodium metal and chlorine gas written as reactants are in their "natural" forms. In other words, you would not find chlorine in nature as just Cl, it must be in the form of Cl_2, for reasons we will look at later. And we have already discovered that the formula for sodium chloride is correct as it is written above. So now we can balance our reaction. On the left we find we have 1 sodium atom and on the right we also have written one sodium atom. The sodium is balanced on both sides of the reaction so we can move on. On the left side we have 2 chlorine atoms, and on the right side we only have one. Pity, the chlorine atoms are not balanced. So, in order to balance them, we put a "2" in front of the formula for sodium chloride, like this:

$$Na + Cl_2 \quad \rightarrow \quad 2NaCl$$

Ah, now the chlorines are balanced. Does this mean that we are done? Not quite yet. In placing that "2" in front of sodium chloride, we increased the number of sodium atoms on the right side. Thus, we must also place a "2" in front of the sodium metal on the left side in order to finish balancing our equation. So our final equation for the formation of sodium chloride looks like this:

$$2Na + Cl_2 \quad \rightarrow \quad 2NaCl$$

At this point in the game this balancing process seems trivial. But as the reactions become more complex, balancing the equation can sometimes become more difficult. Let us look at one more example. Say we took the covalently bonded methane we saw in Chapter 4 and burned it (this is the reaction that occurs when you light the burner on a gas stove). When you burn a molecule made of carbon and hydrogen (called a *hydrocarbon*), you always add oxygen (O_2) and always get carbon dioxide (CO_2) and water (H_2O). Knowing this, let us try to write the balanced reaction for the combustion of methane. First, put the reactants and the products in their correct places.

$$CH_4 + O_2 \quad \rightarrow \quad CO_2 \ + \ H_2O$$

Now we must go through the reaction and make sure all the atoms are balanced. Let us start with carbon and work our way through. There is one carbon atom on the left and one on the right. Since these are balanced we can move on to hydrogen. There are 4 hydrogen atoms on the left, but only 2 on the right! Thus, we must stick a "2" in front of the water product.

$$CH_4 + O_2 \quad \rightarrow \quad CO_2 \ + \ 2H_2O$$

Now the hydrogens are balanced, we can move on to oxygen. There are 2 oxygens on the left, but 4 on the right! So we have to put a "2" in front of the oxygen in the reactants.

$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Now all the atoms on the left balance with all the atoms on the right, so our reaction is written correctly. The systematic workthrough of each element is a good method of balancing reactions.

Section B – Moles Do Not Dig in Dirt! – More Stoichiometry

Now that we have seen how to balance the atoms in a reaction, we can use the same principle for more advanced stoichiometry. Balancing the atoms as we did in the last section is important for the conservation of mass concept. But it does not give us much practical information we can use in an experiment. For example, suppose you needed to make a certain amount of product. How would you know how much of each reactant to use? We need to use the ratios we have seen to get a little more information out of our reaction.

Up to this point we have been discussing reactions in terms of numbers of atoms and molecules involved. The problem with this approach is that atoms are very small, and thus are impossible to count in practice. So instead of talking about just one molecule or atom, we talk about large groups of molecules and atoms. You can think of it like buying grapes. You don't go to the store and buy one grape - you buy a bag or bunch of grapes. Contained within that bag are the individual grapes. It is the same thing when we talk about molecules. We don't talk about single units, we talk about groups. But even 100 is too small a number to use in discussing groups of atoms. We need a much larger number for our groups of molecules. This special number in science is called the *mole*. It is important to understand that a mole is just a number. You can have a mole of atoms or a mole of grapes (that would be a LOT of grapes). This concept of the mole arose from necessity in the laboratory. Long ago before chemists could measure atomic masses or even knew there were subatomic particles they still needed some way of comparing one atom to another. They knew that hydrogen was the smallest atom, so to make things easy for themselves, they arbitrarily assigned hydrogen a mass of "1." They didn't know what the 1 was so they called it an *atomic mass unit*, or amu. Thus, carbon was about 12 amu and oxygen about 16 amu in comparison to hydrogen. Once chemists could measure atomic mass and knew of subatomic particles they realized that they needed a way of talking about amu units in terms of grams. It is impossible to measure out an amu, since it is a conceptual unit. However, chemists could measure out grams. They knew that they wanted the ratios of the amu scale to carry over into grams. Thus if a certain number of molecules of hydrogen had a mass of 1 gram, a container of oxygen with a mass of 16 grams would also have the same number of molecules. The mole was the number that would "bridge the gap" between the amu and the gram. In other words, the mole is the atomic or molecular mass in grams. Thus, one mole of hydrogen has a mass of 1 gram. A mole of oxygen has a mass of approximately 16 grams, and so on. So it is used as a standard number for all atomic masses, which we have already worked with when we calculated molecular and formula masses. When we say that $C_2H_8N_2$ has a molecular mass of about 60, we are saying that one mole of that compound would have a mass of about 60 grams. That is how the mole is used practically; it gives us a way of discussing molecules and atoms on a scale that we as humans can physically work with. But it still needs to be stressed that the mole is JUST a number! It is just a "beef up" constant that allows us to talk about a quantity of molecules or atoms on a scale we can work with on a macroscopic level. When we speak about the number of moles, we are talking *directly* about a huge number of atoms or molecules of a substance. Eventually after many years, an actual number was found for the mole, which was found to be 6.02×10^{23} (also known as Avogadro's number).

But you may be asking yourself why we have to bother with moles at all? We have the coefficients of the reactants and the products, why not just work with masses? The answer to this is simple, but important. Keep in mind that when we find our coefficients for the reactants and products, these indicate the numbers of atoms and molecules. Atoms and molecules have masses different from each other. Just like the mass of an orange differs from that of an apple, masses of atoms differ from one another. While we should not directly compare masses of atoms, we can compare the number of atoms. And since the mole is a constant number of atoms, it serves very well as a way of comparing amounts of reactants and products. Again, when we talk about a number of moles, we are directly referring to the number of atoms or molecules. Learn to equate the two, when you see the word "mole" think atom or molecule.

So how does this help us in the lab? Are there instruments that measure out moles of substances? Of course not. But there is an instrument that can determine how many moles there are in a 1 gram sample of your reactant. Your brain. We know that the periodic table gives us the mass of one mole of each element. All we have to do is a conversion like we learned in Chapter 3 and we can easily find out how many moles are in one gram of our reactant. Now we have some practical information we can use in the lab. If we know how much reactant we are starting with, then we can predict how much product we should obtain. This is useful in two ways. Not only will we know to increase or decrease the amount of reactant based upon how much product we want, but now we also have a way of monitoring how successful our reaction was. If we return little product then the reaction should either be changed or repeated. If we obtain a good percent of product compared to what we expected then the reaction is considered a success.

Before we try using the mole-to-gram calculations for reactions, let us try a few simpler examples first. For instance, let us suppose we have 2 moles of hexane, which has a formula of C_6H_{12}. How many moles of carbon do we have? Each molecule has 6 carbons, so one mole of hexane has 6 moles of carbon atoms. So in 2 moles, we have 12 moles of carbon. Now how many moles of hydrogen? Again, 12 moles of hydrogen atoms are in one mole of the molecule so 2 moles of the molecule contains 24 moles of hydrogen atoms. How many moles of hydrogen and carbon atoms would there be in half a mole of hexane? Well, there are 6 carbons and 12 hydrogens in a full mole, so in half a mole of hexane we have 3 moles of carbon and 6 moles of hydrogen. It is that simple. But this raises a very important point. We can talk about fractions of a mole. We cannot talk about fractions of an atom. You cannot say "measure out half an atom of sodium" but you can say "measure out half a mole of sodium." This makes the mole even more important in a laboratory situation.

Let us try another example. Suppose we want to do a reaction that calls for 5 moles of nitrogen dioxide, which has the formula NO_2. We obviously cannot go into the lab and measure out 5 moles, we must first convert this to grams. Then we can measure out our quantity in the lab. But how do we convert from grams to moles? First, we must find the molecular mass of nitrogen dioxide like we did in chapter 3.

N	1	x	14.0067	=	14.0067
O	2	x	15.9994	=	31.9988
			Total	=	46.0055

So we know that our molar mass is 46.0055 g/mol. But look at the units for that number. The units tell us that each mole of this molecule has a mass of 46.0055 grams. Thus, we can use this as our conversion factor, like this:

$$\frac{5 \text{ moles } NO_2}{1} \quad x \quad \frac{46.0055\text{g}}{1 \text{ mole}} \quad = \quad 230.0275 \text{ grams } NO_2$$

That is a lot of nitrogen dioxide! And while one person in the lab would never use that much, industrial chemistry often works on large scales such as this. Often in the laboratory we use fractions of a mole, since you can see how it gets to be a large quantity very quickly! One very important point to notice here, you see in our above conversion that the number 6.02×10^{23} did not show up at all. This is because this is the constant for saying how many molecules are in a mole. And since we cannot work with individual molecules, this number has no practical application for us in the laboratory. It truly is just a number, and has no physical relevance when working with and calculating conversions for reactions and experiments.

Now that we know we can use the mole, coefficients and atomic masses to obtain information from a reaction, let us try an example. We can start by looking at a simple example, like the formation of sodium chloride that we balanced in our last section.

$$2Na + Cl_2 \rightarrow 2NaCl$$

Let us suppose that we start with 2 grams of sodium and 6 grams of chlorine. To predict how much sodium chloride we will get, we will have to first convert the gram units to mole units. Thus:

$$\frac{2 \text{ grams Na}}{1} \quad x \quad \frac{1 \text{ mole}}{22.9898 \text{ g}} \quad = \quad 0.0870 \text{ moles Na}$$

and:

$$\frac{6 \text{ grams } Cl_2}{1} \quad x \quad \frac{1 \text{ mole}}{70.9060 \text{ g}} \quad = \quad 0.0846 \text{ moles } Cl_2$$

Notice that although chlorine has an atomic mass of 35.4530, we doubled it because we are reacting Cl_2, not Cl. Now that we have the number of moles of each reactant, we can compare them to the ratios in the reaction to predict the moles of product we should obtain. We have 0.846 moles of chlorine. However for every mole of chlorine that reacts we would need 2 moles of sodium, or 0.1692 moles. But we only have 0.0870 moles. This is what we refer to as the *limiting reagent* in the reaction. The limiting reagent is the reactant that limits the amount of product that can be formed. Once all the limiting reagent is used up, the reaction will have to stop. Sodium is our limiting reagent, so our molar ratios must be based on it. Thus, we will be reacting 0.0870 moles of sodium. That means we only need 0.0435 moles of chlorine for this reaction. Chlorine is said to be in *excess* in this reaction. The excess reagent is the reactant that is in surplus. Thus, if we react 0.0870 moles of sodium and 0.0435 moles of chlorine, how many moles of product should we get? The ratio is as follows:

$$\begin{array}{ccccc} \underline{2} & : & \underline{1} & : & \underline{2} \\ 0.0870 & : & 0.0435 & : & ? \end{array}$$

Obviously that molar ratio of product we can expect is 0.0870. Knowing that, and finding that the molecular mass of sodium chloride is 58.4428 (1 * 22.9898 + 1 * 35.4530), we can do another conversion to predict the mass of sodium chloride expected when we react 2 grams of sodium with 6 grams of chlorine.

$$\frac{0.0870 \text{ moles Na}}{1} \quad \text{x} \quad \frac{58.4428 \text{ g}}{1 \text{ mole}} \quad = 5.084 \text{ grams of NaCl}$$

Knowing how to predict the amount of product you expect, and what reactants you have in excess is vital to chemistry. Not only do you have to worry about your products possibly reacting with any excess reactants, but suppose you wanted to make metric tons of sodium chloride. If you used the molar ratios we just used here, all that excess chlorine would not only be an economic waste, but a serious health and environmental hazard!! Stoichiometry and knowing how to use it is a very important concept for practical use of chemical reactions.

Section C – Are You Gunna React? Or Do I Have to Make Ya?

In the last section we learned the importance of balancing the equation for a chemical reaction. We saw that stoichiometry could be used to predict the mass of the product obtained and how much of our reactants we should use. And we saw that stoichiometry was governed by the laws of conservation. But there is another property of a chemical reaction that is governed by the laws of conservation: energy.

Chemical reactions always involve a change in energy. Just as the conservation of mass law states that matter can neither be created nor destroyed, so, too, the conservation of energy law states that energy is neither created nor destroyed during any chemical process. A chemical reaction either absorbs or releases energy. Thus we can break down chemical reactions into two categories, those that absorb energy during the reaction process and those that release energy.

When a chemical reaction absorbs energy, it is said to be *endothermic*. Endothermic reactions cannot happen without the necessary amount of energy donated into the system. For example, imagine the process of baking bread. Breaking the process down we can consider just the reaction of the baking soda decomposing to release carbon dioxide (this is what makes bread rise). This reaction is endothermic because it requires heat from the oven to break the baking soda ($NaHCO_3$) down and release carbon dioxide. The energy from the oven (whether it is electrical or gas) is transferred into heat energy that the baking soda absorbs to react. Traditionally, endothermic reactions are described as reactions where the products are at a higher energy than the reactants. This is illustrated in figure 5.1 below.

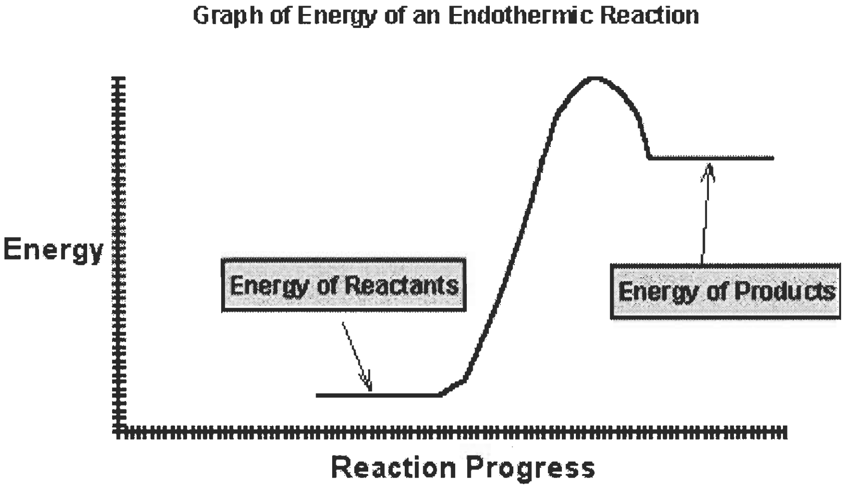


Figure 5.1 – Energy Diagram of an Endothermic Reaction

Figure 5.1 is known as an energy diagram. It plots the energy of the system as the reaction progresses. The plateau at the beginning of the curve represents the energy of the reactants, and the plateau at the end represents the energy of the products. As you can see from the figure, energy had to be added to this reaction in order to get the reactants to the products.

Similarly, a reaction that gives off energy is said to be *exothermic*. Like endothermic reactions, exothermic reactions sometimes need a small input of energy. But wait a minute; did we not just say that exothermic reactions give off energy? Yes we did, and yes they do. The secret is that over-all an exothermic reaction will release more energy than it absorbs. If you look back to figure 5.1, you will see that the reaction travels up to a large increase in energy but then decreases a little to get to the product plateau. Overall, however, the reaction absorbs energy. An exothermic reaction absorbs a small amount of energy at first, but then releases more than it absorbs. This is shown in figure 5.2.

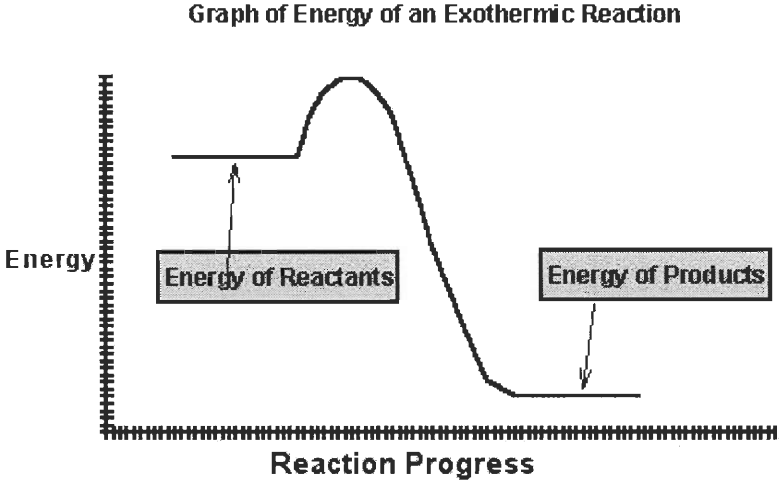


Figure 5.2 – Energy Diagram of an Exothermic Reaction

As you can see, after the initial "hump" in the energy diagram, the exothermic reaction releases energy. Traditionally an exothermic reaction is defined as a reaction in which the products are at a lower energy than the reactants, as figure 5.2 shows. A prime example of an exothermic reaction is sitting right inside your stomach. Our bodies break down the food we eat to release the energy from the chemical bonds in order to power our bodies and grow. Without exothermic reactions like this, large complex organisms like mammals would not be possible.

The "hump" in the energy diagram is what is known as the *activation energy* of the reaction. Simply put, this is the energy required to make the reaction proceed. Think of it as a "kick start" to the reaction. When starting a lawnmower, you have to pull the cord to assist the motor in starting. This is just like a chemical reaction. Think of the energy your body uses pulling the cord as the activation energy. The motor is the reaction progress. Once the activation energy is reached, or the cord pulled enough times, the reaction can proceed, the motor will run on its own. You can see that (in most cases) endothermic reactions have a much larger activation energy than exothermic reactions.

The energy change in a reaction, whether endo or exothermic, is often known as the *enthalpy* change of the reaction. We will look at enthalpy more in depth in Chapter X, but for now you should understand that enthalpy is very similar to energy. The graphs of Figure 5.1 and 5.2 show the enthalpy change for those reactions. Enthalpy has a few complexities to it, but as long as the reaction is carried out at constant pressure, than enthalpy is simply the heat of the reaction we have been discussing. Another concept often considered when looking at reactions is called *entropy*. Entropy is simply a measure of the disorder of the system. Well what does that mean and why do we care? It means that as the reaction progresses we can either have the molecules becoming more ordered, or less ordered. Consider freezing water. The molecules in the liquid phase have free range of motion, but once you freeze it, the molecules of water in the solid phase are closely packed with a much higher restriction of movement. You have decreased the entropy; there is more order in the system, so the overall disorder has decreased. However, now boil water instead of freezing it. Now the molecules have much more freedom of motion and thus are considered to have a higher state of disorder. Thus you have increased the entropy for the system.

We care about entropy and enthalpy because predicting and calculating these properties of a reaction can help us predict another property of a reaction: spontaneity. A reaction is said to be spontaneous if it occurs without having assistance from an outside source. In other words once a spontaneous reaction is kick started, it will continue to react without any further help. If the reaction cannot occur on its own, but instead must be driven from outside, then the reaction is not spontaneous. Spontaneity is measured by a quantity called the *free energy* of the system. It is dependent on enthalpy, entropy, and also temperature. Obviously a glass of water sitting at room temperature on your desk will not freeze because the reaction is not spontaneous at that temperature. But put the glass outside on a cold winter day and freezing becomes a spontaneous process. The concepts mentioned in this section are a branch of chemistry known as thermodynamics. We will take a closer look at thermodynamics and how to calculate the properties introduced here later in Chapter X.

Musings on a Revolutionary Text

I believe it is important for all authors to review and comment on their work, for the benefit of both creator and reader. Reflecting on this project, I feel not only pleased but proud of the results. The objective of this project was to create a text book that would help anyone learn chemistry; I wanted to break down the long held belief that chemistry was impossibly hard to learn. All too often I find myself dismayed at people who are initially excited to find out I will soon graduate then wrinkle their nose and glaze over when I tell them I am in the field of chemistry. "Oh that was my worst subject in high school; it's so hard to understand." Their words echoed in my ears constantly as I was writing. This is for them. And this is for the younger generations, because I feel that more young scientists would enter the field of chemistry if the negative image that has snowballed around the subject was melted away.

The approach to the text sounded simple to me: write a text anyone could understand. This proved to be easier said than done. Many chapter drafts were generated and exchanged back and forth between my advisor and myself. After the cores of the chapters were laid down, the actual writing could begin. Flushing out chapters was typically done once, then "tweaked" to make sure the writing flowed, the information was correct, and that it was easy to understand. I found many "test subjects" in my friends and family, from both scientific and non-scientific backgrounds. I made sure that my engineering friends could understand it, and asked them to bring up anything they did not understand to me. Often when they did, I would verbally explain it to them, and then run to the computer to put my understood verbal explanation into the chapter. The biggest aid in making sure the chapters ran smoothly was my advisor, however. Being a professor he has a greater knowledge as to what students needed to understand and grasp the concepts I introduced. My advisor was also crucial in explaining to me that you have to build from the bottom up. I had a tendency to want to take the chapters at too quick a pace, or to introduce a concept that needed other concepts explained first in order to understand it. This text was like a ship to me, I was the engine, my advisor was the navigation and my friends and family were the passengers.

To be honest, the project did not go as I initially expected. Walking into the project I thought I would be writing *some* chapters, perhaps towards the middle or the end of the book. It did not even occur to me I would be starting from the beginning and continuing from there. This shifted the focus of what I originally would have liked to write about, but made the project much more interesting and challenging. I also found that my advisor and I differed in our opinion of word choice slightly, but for such a minor point I was willing to submit to his suggestions. I understood that just because it made sense in my mind did not mean it would make sense to everyone else. I think the order of topics introduced is slightly different as well than they would be had I written this text on my own. Nevertheless, the project accomplished the goals we had set for it, and overall I'm extremely happy with the way it turned out. I view it as sailing to a tropical island. You have it in your sights because it looks beautiful and untouched, only to get blown off course to a different island you never even saw before. And reflecting back at the other one you notice a tourist city that was hidden from your view before. In other words, the book as it is now is not what I originally planned, but I'm more pleased with how it turned out than I ever was with my original plans. I think this current layout reaches the objectives better than my plans could have.

This project has been a tremendous learning experience for me. Not only do I have a deeper appreciation for all my professors who have taught from the "bad" books and still made the subject clear and interesting, I also have a better appreciation of how difficult writing a book can be. I also not only gained a firmer grasp of fundamental chemical concepts myself, but I have a better understanding of how to communicate chemistry to non-scientifically minded individuals. Breaking down concepts and being able to teach someone something new has led to another feeling I will take with me for the rest of my life, and that is teaching. This project has inflamed an already growing desire within me to teach, and not only do I hope to eventually become a professor, but I sincerely plan on continuing this project and hope to one day see the text in classrooms. It also tested my faith in my field, since I found rather early on that concepts I accept as fact are either not fully explained or they are only believed to be true because they have never been disproved. To have all my "advanced knowledge" hinging on the fact that the electron stays in an orbit around the nucleus and not being able to explain why is something I lost sleep over! But learning how to get over those hurdles and get past writer's block and communicate a subject I'm passionate about to another generation is why I took on this task in the first place. Loving a subject is one thing. Involving and teaching it to someone else is what human interaction and progress is all about.