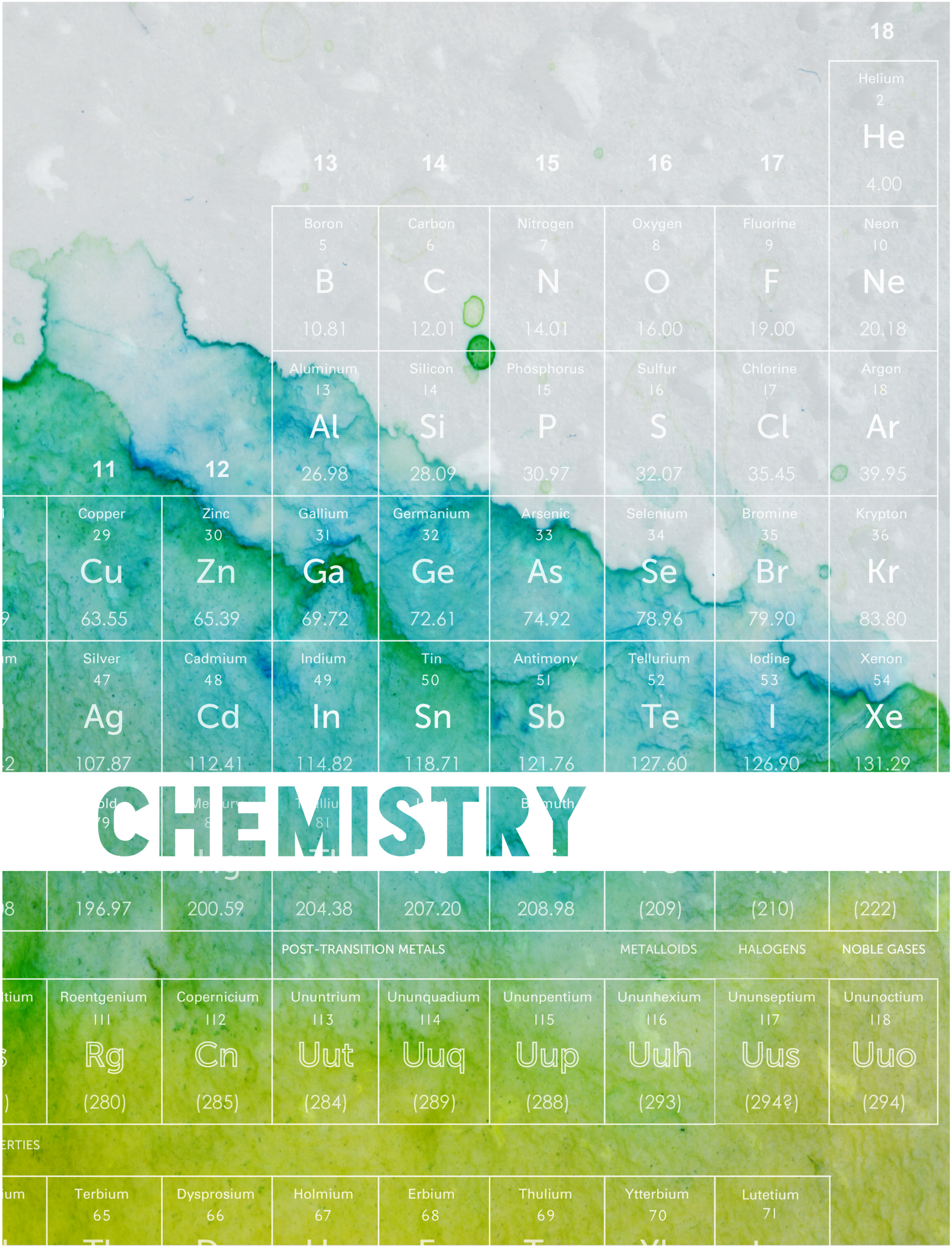
CHEMISTRY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CREDITS

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We also thank the amazing Utah science teachers whose collaborative efforts made the book possible. Thank you for your commitment to science education and Utah students!

Textbook design, periodic table and cover by Corrine Beaumont, PhD.

THIS BOOK IS YOURS TO KEEP, SO WRITE ON ALL THE PAGES YOU LIKE. THERE'S EVEN SPACE ON THE BACK COVER FOR YOUR NAME.

IT'S YOUR LEARNING, IT'S YOUR BOOK. ENJOY.

This space on the edge of each page is your place for drawing, taking notes and making the book your own.

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UTAH SCIENCE CORE CURRICULUM ALIGNMENT

STANDARD 1: STUDENTS WILL UNDERSTAND THAT ALL MATTER IN THE UNIVERSE HAS A COMMON ORIGIN AND IS MADE OF ATOMS, WHICH HAVE STRUCTURE AND CAN BE SYSTEMATICALLY ARRANGED ON THE PERIODIC TABLE.

Objective 1: Recognize the origin and distribution of elements in the universe. [P. 40]

a) Identify evidence supporting the assumption that matter in the universe has a common origin.
b) Recognize that all matter in the universe and on earth is composed of the same elements.
c) Identify the distribution of elements in the universe.
d) Compare the occurrence of heavier elements on earth and the universe.

Objective 2: Relate the structure, behavior, and scale of an atom to the particles that compose it. [P. 44]

a) Summarize the major experimental evidence that led to the development of various atomic models, both historical and current.
b) Evaluate the limitations of using models to describe atoms.
c) Discriminate between the relative size, charge, and position of protons, neutrons, and electrons in the atom.
d) Generalize the relationship of proton number to the element's identity.
e) Relate the mass and number of atoms to the gram-sized quantities of matter in a mole.

Objective 3: Correlate atomic structure and the physical and chemical properties of an element to the position of the element on the periodic table. [P. 71]

a) Use the periodic table to correlate the number of protons, neutrons, and electrons in an atom.
b) Compare the number of protons and neutrons in isotopes of the same element.
c) Identify similarities in chemical behavior of elements within a group.
d) Generalize trends in reactivity of elements within a group to trends in other groups.
e) Compare the properties of elements (e.g., metal, nonmetallic, metalloid) based on their position in the periodic table.

STANDARD 2: STUDENTS WILL UNDERSTAND THE RELATIONSHIP BETWEEN ENERGY CHANGES IN THE ATOM SPECIFIC TO THE MOVEMENT OF ELECTRONS BETWEEN ENERGY LEVELS IN AN ATOM RESULTING IN THE EMISSION OR ABSORPTION OF QUANTUM ENERGY. THEY WILL ALSO UNDERSTAND THAT THE EMISSION OF HIGH-ENERGY PARTICLES RESULTS FROM NUCLEAR CHANGES AND THAT MATTER CAN BE CONVERTED TO ENERGY DURING NUCLEAR REACTIONS.

Objective 1: Evaluate quantum energy changes in the atom in terms of the energy contained in light emissions. [P. 86]

a) Identify the relationship between wavelength and light energy.
b) Examine evidence from the lab indicating that energy is absorbed or released in discrete units when electrons move from one energy level to another.
c) Correlate the energy in a photon to the color of light emitted.
d) After observing spectral emissions in the lab (e.g., flame test, spectrum tubes), identify unknown elements by comparison to known emission spectra.

Objective 2: Evaluate how changes in the nucleus of an atom result in emission of radioactivity. [P. 95]

a) Recognize that radioactive particles and wavelike radiations are products of the decay of an unstable nucleus.
b) Interpret graphical data relating half-life and age of a radioactive substance.
c) Compare the mass, energy, and penetrating power of alpha, beta, and gamma radiation.
d) Compare the strong nuclear force to the amount of energy released in a nuclear reaction and contrast it to the amount of energy released in a chemical reaction.
e) After researching, evaluate and report the effects of nuclear radiation on humans or other organisms.

STANDARD 3: STUDENTS WILL UNDERSTAND CHEMICAL BONDING AND THE RELATIONSHIP OF THE TYPE OF BONDING TO THE CHEMICAL AND PHYSICAL PROPERTIES OF SUBSTANCES.

Objective 1: Analyze the relationship between the valence (outermost) electrons of an atom and the type of bond formed between atoms. [P. 116]

a) Determine the number of valence electrons in atoms using the periodic table.
b) Predict the charge an atom will acquire when it forms an ion by gaining or losing electrons.
c) Predict bond types based on the behavior of valence (outermost) electrons.
d) Compare covalent, ionic, and metallic bonds with respect to electron behavior and relative bond strengths.

Objective 2: Explain that the properties of a compound may be different from those of the elements or compounds from which it is formed. [P. 130]

a) Use a chemical formula to represent the names of elements and numbers of atoms in a compound and recognize that the formula is unique to the specific compound.
b) Compare the physical properties of a compound to the elements that form it.
c) Compare the chemical properties of a compound to the elements that form it.
d) Explain that combining elements in different proportions results in the formation of different compounds with different properties.

Objective 3: Relate the properties of simple compounds to the type of bonding, shape of molecules, and intermolecular forces. [P. 152]

a) Generalize, from investigations, the physical properties (e.g., malleability, conductivity, solubility) of substances with different bond types.
b) Given a model, describe the shape and resulting polarity of water, ammonia, and methane molecules.
c) Identify how intermolecular forces of hydrogen bonds in water affect a variety of physical, chemical, and biological phenomena (e.g., surface tension, capillary action, boiling point).

STANDARD 4: STUDENTS WILL UNDERSTAND THAT IN CHEMICAL REACTIONS MATTER AND ENERGY CHANGE FORMS, BUT THE AMOUNTS OF MATTER AND ENERGY DO NOT CHANGE.

Objective 1: Identify evidence of chemical reactions and demonstrate how chemical equations are used to describe them. [P. 168]

a) Generalize evidences of chemical reactions.
b) Compare the properties of reactants to the properties of products in a chemical reaction.
c) Use a chemical equation to describe a simple chemical reaction.
d) Recognize that the number of atoms in a chemical reaction does not change.
e) Determine the molar proportions of the reactants and products in a balanced chemical reaction.
f) Investigate everyday chemical reactions that occur in a student's home (e.g., baking, rusting, bleaching, cleaning).

Objective 2: Analyze evidence for the laws of conservation of mass and conservation of energy in chemical reactions. [P. 178]

a) Using data from quantitative analysis, identify evidence that supports the conservation of mass in a chemical reaction.
b) Use molar relationships in a balanced chemical reaction to predict the mass of product produced in a simple chemical reaction that goes to completion.
c) Report evidence of energy transformations in a chemical reaction.
d) After observing or measuring, classify evidence of temperature change in a chemical reaction as endothermic or exothermic.
e) Using either a constructed or a diagrammed electrochemical cell, describe how electrical energy can be produced in a chemical reaction (e.g., half reaction, electron transfer)

STANDARD 5: STUDENTS WILL UNDERSTAND THAT MANY FACTORS INFLUENCE CHEMICAL REACTIONS AND SOME REACTIONS CAN ACHIEVE A STATE OF DYNAMIC EQUILIBRIUM.

Objective 1: Evaluate factors specific to collisions (e.g., temperature, particle size, concentration, and catalysts) that affect the rate of chemical reaction. [P. 218]

a) Design and conduct an investigation of the factors affecting reaction rate and use the findings to generalize the results to other reactions.

b) Use information from graphs to draw warranted conclusions about reaction rates.
c) Correlate frequency and energy of collisions to reaction rate.
d) Identify that catalysts are effective in increasing reaction rates.

Objective 2: Recognize that certain reactions do not convert all reactants to products, but achieve a state of dynamic equilibrium that can be changed. [P. 232]

a) Explain the concept of dynamic equilibrium.
b) Given an equation, identify the effect of adding either product or reactant to a shift in equilibrium.
c) Indicate the effect of a temperature change on the equilibrium, using an equation showing a heat term.

STANDARD 6: STUDENTS WILL UNDERSTAND THE PROPERTIES THAT DESCRIBE SOLUTIONS IN TERMS OF CONCENTRATION, SOLUTES, SOLVENTS, AND THE BEHAVIOR OF ACIDS AND BASES.

Objective 1: Describe factors affecting the process of dissolving and evaluate the effects that changes in concentration have on solutions. [P. 254]

a) Use the terms solute and solvent in describing a solution.
b) Sketch a solution at the particle level.
c) Describe the relative amount of solute particles in concentrated and dilute solutions and express concentration in terms of molarity and molality.
d) Design and conduct an experiment to determine the factors (e.g., agitation, particle size, temperature) affecting the relative rate of dissolution.
e) Relate the concept of parts per million (PPM) to relevant environmental issues found through research.

Objective 2: Summarize the quantitative and qualitative effects of colligative properties on a solution when a solute is added. [P. 265]

a) Identify the colligative properties of a solution.
b) Measure change in boiling and/or freezing point of a solvent when a solute is added.
c) Describe how colligative properties affect the behavior of solutions in everyday applications (e.g., road salt, cold packs, antifreeze).

Objective 3: Differentiate between acids and bases in terms of hydrogen ion concentration. [P. 268]

a) Relate hydrogen ion concentration to pH values and to the terms acidic, basic or neutral.
b) Using an indicator, measure the pH of common household solutions and standard laboratory solutions, and identify them as acids or bases.
c) Determine the concentration of an acid or a base using a simple acid-base titration.
d) Research and report on the uses of acids and bases in industry, agriculture, medicine, mining, manufacturing, or construction.
e) Evaluate mechanisms by which pollutants modify the pH of various environments (e.g., aquatic, atmospheric, soil).

INTRODUCTION

CHAPTER 0

Vocabulary

- chemistry

LESSON OBJECTIVES

The student will:

- give a brief history of how chemistry began.
- list some new materials produced by chemists.

INTRODUCTION

During medieval times, a group of people known as alchemists began looking for ways to transform common metals, such as lead, copper, and iron, into gold. Can you imagine how much money you would make if you could go to the store, buy some iron nails, and turn them into gold? You'd be rich in no time!

The Origin and Evolution of Chemistry

Alchemists experimented with many different kinds of chemicals, searching for what they termed the "philosopher's stone" – a legendary substance necessary for transforming common metals into gold (see Figure below). We now know that there is no such thing as a philosopher's stone, nor is there any chemical reaction that creates gold from another metal. We know this because we now have a much better understanding of how the matter in our universe behaves. Nevertheless, those early alchemists kindled interest in chemical transformations and inspired the development of modern chemistry.

This painting shows an alchemist in search of the philosopher's stone.

Chemistry is the scientific study of matter and the changes that it undergoes. It's no coincidence that the word "chemistry" looks a lot like the word "alchemy." Early alchemists were commonly known as "chemists," and over time, people started referring to their work, particularly the more legitimate forms of it, as chemistry. While it may seem strange, in many ways it is appropriate that our word for the present-day study of matter comes from the early practice of alchemy. A lot of the techniques and equipment fundamental to modern chemistry were actually developed by early alchemists.

Early chemistry, or alchemy, was not very systematic or well-reviewed. In fact, in many areas alchemy was considered to be form of magic or sorcery. It wasn't until the late 17th century that European chemists began applying a scientific methodic process. Robert Boyle (1627 – 1691) was the first European chemist to do so, using quantitative experiments to measure the relationship between the pressure and the volume of a gas. His use of scientific methods paved the way for other European chemists and helped to establish the modern science of chemistry.

The man who would greatly advance the development of modern chemistry was Antoine Lavoisier (1743 - 1794). Considered the father of modern chemistry, Lavoisier (seen in Figure below) discovered that although matter may change its shape or form, its mass always remains the same. As a result, he would state the first version of the law of conservation of mass. Lavoisier also wrote the first extensive list of elements, including oxygen and hydrogen, and helped to reform chemical nomenclature. While Lavoisier was extremely important to the advancement of chemistry, there were many important figures that helped the science of chemistry move forward and improve our understanding of the natural world.

Antoine Lavoisier is considered to be the father of modern chemistry due to his many contributions to chemistry.

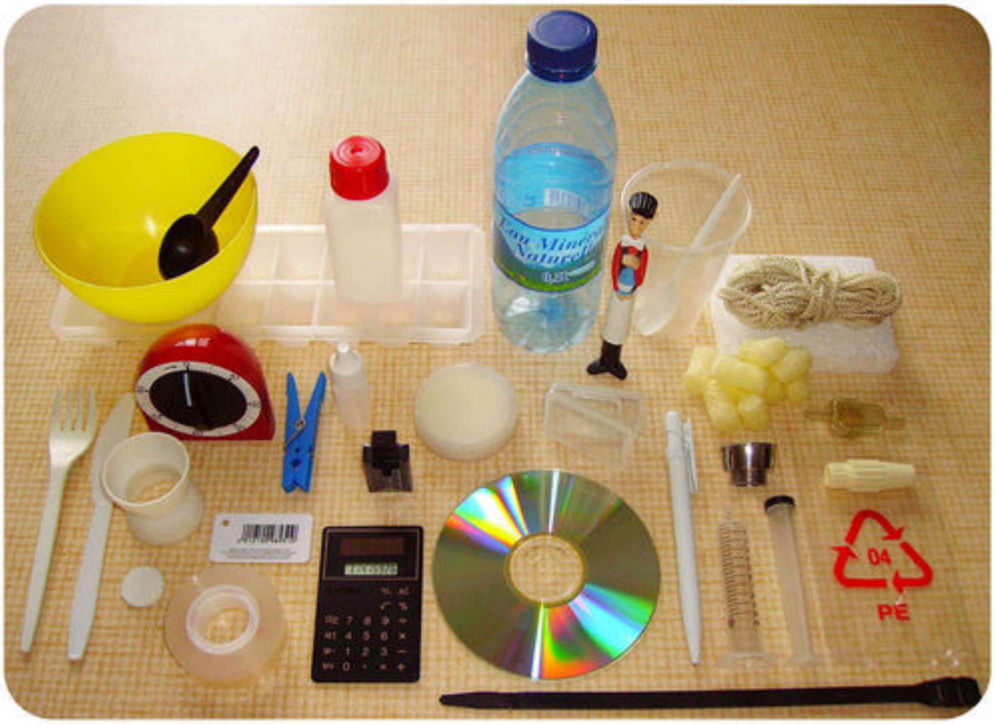
What Chemists Do

You might wonder why the study of chemistry is so important if you can't use it to turn iron into gold or to develop a potion that will make you immortal. Why didn't chemistry die when scientists like Boyle and Lavoisier proved alchemy was nothing but a hoax? Although we can't use chemistry to make gold or to live forever, modern chemistry is still very powerful. There may be no such thing as a potion that cures all diseases, but many chemists today are working to develop cures for specific diseases, including HIV/AIDS and various forms of cancer.

Chemists apply information about matter and the changes it undergoes to improve our lives in many different ways. Modern chemists study not only chemicals that can help us, but also chemicals that can hurt us. For example, environmental chemists test the air, soil, and water in our neighborhoods to make sure that we aren't exposed to heavy metals (such as mercury or lead) or chemical pesticides. Moreover, when environmental chemists do find dangerous substances, they use their knowledge of chemistry to clean up the contamination. Similarly, every time you buy packaged food from the grocery store, you can be sure that many tests have been done by chemists to ensure those foods don't contain any toxins or carcinogens (cancer-causing chemicals).

Chemists are also responsible for creating many important materials that we use today. In addition, many technologies rely on chemistry as well. In fact, flat-screen LCD televisions, cubic zirconium rings, and energy-efficient LED lights are all thanks to our improved understanding of chemistry.

One of the huge breakthroughs in recent history has been the discovery of plastics (see Figure below). Initially, plastic was made by chemically modifying cellulose, a naturally occurring chemical found in plants. As chemical knowledge developed, however, scientists began to realize that plastics had special properties. On a microscopic scale, plastics are composed of thousands of tiny chains of molecules all tangled up together. Scientists reasoned that if they altered the chemicals in these chains but still managed to keep the chains intact, they could make new plastics with new properties. Thus began the plastic revolution!



Some common objects made of plastic.

Semiconductors are another class of "new" materials whose development is largely based on our improved understanding of chemistry. Because scientists know how matter is put together, they can predict how to fine-tune the chemical composition of a semiconductor in order to make it absorb light and act as a solar cell, or to emit light and act as a light source.

We've come a long way from our early days of producing bronze and steel. Nevertheless, as our understanding of chemistry improves, we will be able to create even more useful materials than the ones we have today.

Lesson Summary

- Chemistry began as the study of alchemy. Most alchemists were searching for the philosopher's stone, a fabled substance that could turn common metals into gold.
- Chemistry is the scientific study of matter and the changes that it undergoes.
- The word "chemistry" comes from the Arabic word "al-kimia," meaning "the art of transformation."
- Chemists apply information about matter and the changes it undergoes in many different ways to improve our lives.

Review Questions

1. What was the origin of the word "chemistry"?
2. Name at least two new materials created by chemists in the last 40 years.

Glossary

Chemistry is the scientific study of matter and the changes that it undergoes

THE SCIENTIFIC METHOD

Lesson Objectives
The student will:

- describe the scientific method of problem solving.
- list some values of the scientific method of problem solving.
- describe the difference between hypothesis, theory, and scientific law as scientific terms.
- explain the necessity for experimental controls.
- identify components in an experiment that represent experimental controls.
- explain the concept of a model and why scientists use models.
- explain the limitations of models as scientific representations of reality.

Vocabulary

- experiment
- hypothesis
- model
- scientific law
- scientific method
- theory

The Scientific Method of Problem Solving

In the 16th and 17th centuries, innovative thinkers were developing a new way to understand the nature of the world around them. They were developing a method that relied upon making observations of phenomena and drawing conclusions that corresponded to their observations.

The **scientific method** is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. Scientists frequently list the scientific method as a series of steps. Some scientists oppose this listing of steps because not all steps occur in every case, nor do they always occur in the same order. The scientific method is listed here as a series of steps, but you should remember that you are not required to rigidly follow this list. Instead, the scientific method is a valuable tool that provides a basic and adaptable strategy for tackling scientific questions.

1. Identify the problem or phenomenon that needs to be investigated. This is sometimes referred to as "defining the problem."
2. Gather and organize data on the problem. This step is also known as "making observations."
3. Suggest a possible solution or explanation. A suggested solution is called a **hypothesis**.
4. Test the hypothesis by making new observations.
5. If the new observations support the hypothesis, you accept the hypothesis for further testing. If the new observations do not agree with your hypothesis, add the new observations to your observation list and revisit your problem or experiment. This step is called a conclusion.

Scientific Hypotheses, Theories, and Laws

Hypotheses that have passed many supportive tests are often called theories. A **theory** is an explanation that summarizes a hypothesis or a set of hypotheses and has been supported with repeated testing by many different people. Theories have a great deal more supportive testing behind them than do hypotheses.

In science, theories can either be descriptive (qualitative) or mathematical (quantitative). However, a scientific theory must be falsifiable, or capable of being proved false, in order to be accepted as a theory. A theory is never proven true and is never a "fact." As long as a theory is consistent with all observations, scientists will continue to use it. When a theory is contradicted by observations, it is discarded, replaced or modified.

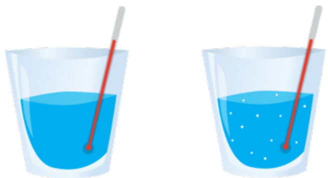
A theory is also a possible explanation for a law. A scientific law is a statement that summarizes the results of many observations and experiments. A law describes an observed pattern in data that occurs without any known exception. A law that has withstood the test of time is incorporated into the field of knowledge. Because they explain the patterns described in laws, theories can be used to predict future events. In this video a teacher discusses the difference between a theory and a law: **http://www.youtube.com/watch?v=eDED5fCY86s (4:18).**

Experimentation

The scientific method requires that observations be made. Sometimes the phenomenon we wish to observe does not occur in nature or is inconvenient for us to observe. If we can arrange for the phenomenon to occur at our convenience, we can control other variables and have all of our measuring instruments on hand to help us make observations. An experiment is a controlled method of testing a hypothesis under the conditions we want at a time and place of our choosing. When scientists conduct experiments, they are usually seeking new information or trying to verify someone else's data. In comparison, classroom experiments often demonstrate and verify information that is already known to scientists but may be new to students.

Suppose a scientist observed two pools of water in bowl-shaped rocks that are located near each other while walking along the beach on a very cold day following a rainstorm. One of the pools was partially covered with ice, while the other pool had no ice on it. The unfrozen pool seemed to contain seawater that splashed up on the rock from the surf, but the other pool was too high up for seawater to splash in and was most likely filled with only rainwater.

Since both pools were at the same temperature, the scientist wondered why one pool was partially frozen and the other was not. By tasting the water (not a good idea), the scientist determined that the unfrozen pool tasted saltier than the partially frozen one. The scientist thought perhaps salt water had a lower freezing point than fresh water, so she decided to go home to test her hypothesis. In order to test this hypothesis, the scientist will conduct an experiment during which she can make accurate observations. So far, the scientist has identified a question, gathered a small amount of data, and suggested a hypothesis.



For the experiment, the scientist prepared two identical containers of fresh water and added some salt to one of them. A thermometer was placed in each container, and both containers were placed in a freezer. The scientist then observed the conditions and temperatures of the two liquids at regular intervals (see the tables below).

Temperature and Conditions of Fresh Water

time (min)	temperature (°C)	condition
0	25	liquid
5	20	liquid
10	15	liquid
15	10	liquid
20	5	liquid
25	0	frozen
30	-5	frozen

Temperature and Conditions of Salt Water

time (min)	temperature (°C)	condition
0	25	liquid
5	20	liquid
10	15	liquid
15	10	liquid
20	5	liquid
25	0	liquid
30	-5	frozen

The scientist found support for her hypothesis from this experiment: fresh water freezes at a higher temperature than salt water. Much more support would be needed before the scientist is confident in this hypothesis. She would perhaps ask other scientists to verify the work.

In the scientist's experiment, it was necessary that she freeze the salt water and fresh water under exactly the same conditions. Why? The scientist was testing whether or not the presence of salt in water would alter its freezing point. It is known that even changing the air pressure will alter the freezing point of water. In order to conclude that the presence of the salt was what caused the change in freezing point, all other conditions had to be identical. When doing an experiment, it is important to set up the experiment so that relationships can be seen clearly.

Scientific Models

Chemists rely on both careful observation and well-known physical laws. By putting observations and laws together, chemists develop models. A **model** is a descriptive, graphic, or three-dimensional representation of a hypothesis or theory used to help enhance understanding. Scientists often use models when they need a way to communicate their understanding of what might be very small (such as an atom or molecule) or very large (such as the universe).

A model is any simulation, substitute, or stand-in for what you are actually studying and provide a way of predicting what will happen given a certain set of circumstances. A good model contains the essential variables that you are concerned with in the real system, explains all the observations on the real system, and is as simple as possible. A model may be as uncomplicated as a sphere representing the earth or billiard balls representing gaseous molecules, but it may also be as complex as mathematical equations representing light.

If you were asked to determine the contents of a box that cannot be opened, you could do a variety of experiments in order to develop an idea (or a model) of what the box contains. You would probably shake the box, perhaps put magnets near it, and possibly determine its mass. When you completed your experiments, you would develop an idea of what is inside; that is, you would propose a model of what is inside the box that cannot be opened. With your model, you could predict how the unopened box would behave under a different set of conditions.

However, even though your model may be capable of accurately predicting some behavior of the unopened box, you would find that the model does not always agree with new experimental results and observations. The model is only be as good as the data you have collected. Because you would never be able to open the box to see what is inside, you also would never be able to create a perfectly accurate model of the box. The model can only be modified and refined with further experimentation.

Chemists have created models about what happens when different chemicals are mixed together, heated up, cooled down, or compressed by using many observations from past experiments. They use these models to predict what might happen during future experiments. Once chemists have models that predict the outcome of experiments reasonably well, those working models can be applied for practical purposes, such as producing an especially strong plastic or detecting potential toxins in your food.

A good example of how a model is useful to scientists is to examine how models were used to develop the atomic theory. As you will learn in the chapter “The Atomic Theory,” the concept of an atom has changed over many years. In order to understand the different theories of atomic structure proposed by various scientists, models were drawn to make the concepts easier to understand.

LESSON SUMMARY

- The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions.
- Parts included in the scientific method are:
    - Identify the problem.
    - Gather data (make observations).
    - Suggest a hypothesis.
    - Test the hypothesis (experiment).
    - Accept the hypothesis for further testing, or reject the hypothesis and make a new one.
- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is an explanation that summarizes a hypothesis or a set of hypotheses and has been supported with repeated testing.
- A scientific law is a statement that summarizes the results of many observations and experiments.
- An experiment is a controlled method of testing a hypothesis.
- A model is a descriptive, graphic, or three-dimensional representation of a hypothesis or theory used to help enhance understanding.
- Scientists often use models when they need a way to communicate their understanding of what might be very small (such as an atom or molecule) or very large (such as the universe).

Vocabulary

- base unit
- metric system

LESSON OBJECTIVES

The student will:

- state an advantage of using the metric system over the United States customary system.
- state the different prefixes used in the metric system.

Introduction

Even in ancient times, humans needed measurement systems for commerce. Land ownership required measurements of length, and the sale of food and other commodities required measurements of mass. The first elementary efforts in measurement required convenient objects to be used as standards, such as the human body. Inch and foot are examples of measurement units that are based on parts of the human body. The inch is based on the width of a man's thumb, and the foot speaks for itself. The grain is a unit of mass measurement that is based upon the mass of a single grain of wheat. Because grains of wheat are fairly consistent in mass, the quantity of meat purchased could be balanced against some number of grains of wheat on a merchant's balance.

It should be apparent that measuring the foot of two different people would lead to different results. One way to achieve greater consistency was for everyone to use the foot of one person, such as the king, as the standard. The length of the king's foot could be marked on pieces of wood, and everyone who needed to measure length could have a copy. Of course, this standard would change when a new king was crowned.

What were needed were objects that could be safely stored without changing over time to serve as standards of measurement. Copies of these objects could then be made and distributed so that everyone was using the exact same units of measure. This was especially important when the requirements of science necessitated accurate, reproducible measurements.

The Metric System

The **metric system** is an international decimal-based system of measurement. Because the metric system is a decimal system, making conversions between different units of the metric system are always done with factors of ten. To understand why this makes the metric system so useful and easy to manipulate, let's consider the United States customary system – that is, the measurement system commonly used in the US. For instance, if you need to know how many inches are in a foot, you need to remember:

$$12 \text{ inches} = 1 \text{ foot}$$

Now imagine that you now need to know how many feet are in a mile. What happens if you have never memorized this fact before? Of course, you can find this conversion online or elsewhere, but the point is that this information must be given to you, as there is no way for you to derive it by yourself. This is true about all parts of the United States customary system: you have to memorize all the facts that are needed for different measurements.

Metric Prefixes and Equivalents

The metric system uses a number of prefixes along with the base units. A **base unit** is one that cannot be expressed in terms of other units. The base unit of mass is the gram (g), that of length is the meter (m), and that of volume is the liter (L). Each base unit can be combined with different prefixes to define smaller and larger quantities. When the prefix “centi-” is placed in front of gram, as in centigram, the unit is now $\frac{1}{100}$ of a gram. When “milli-” is placed in front of meter, as in millimeter, the unit is now $\frac{1}{1,000}$ of a meter. Common prefixes are shown in **Table** below.

Common Prefixes

Prefix	Meaning	Symbol
pico-	10^{-12}	p
nano-	10^{-9}	n
micro-	10^{-6}	μ (pronounced *mu*)
milli-	10^{-3}	m
centi-	10^{-2}	c
deci-	10^{-1}	d
kilo-	10^{3}	k

Common metric units, their symbols, and their relationships to a base unit are shown below:

1.00 picogram = 1.00 pg = 1.00×10^{-12}

1.00 nanosecond = 1.00 ns = 1.00×10^{-9}

1.00 micrometer = 1.00 μm = 1.00×10^{-6}

1.00 centimeter = 1.00 cm = 1.00×10^{-12}

1.00 deciliter = 1.00 dL = 1.00×10^{-2}

1.00 kilogram = 1.00 kg = 1.00×10^{3}

You can express a given measurement in more than one unit. If you express a measured quantity in two different metric units, then the two measurements are metric equivalents. Common metric equivalents are shown below.

- Length:

  1,000 millimeters = 1 meter
  100 centimeters = 1 meter
  10 millimeters = 1 centimeter

- Mass:

  1,000 milligrams = 1 gram

  1,000 grams = 1 kilogram

- Volume:

$$1 \text{ liter} = 1,000 \text{ milliliters}$$

LESSON SUMMARY

- The metric system is an international decimal-based system of measurement.
- The metric system uses a number of prefixes along with the base units.
- The prefixes in the metric system are multiples of 10.
- A base unit is one that cannot be expressed in terms of other units
- If you express a measured quantity in two different metric units, then the two measurements are metric equivalents.

LESSON OBJECTIVES

The student will:

- explain the difference between mass and weight.
- identify SI units of mass, distance (length), volume, temperature, and time.
- define derived unit.
- describe absolute zero.

Vocabulary

- International System of Units
- Kelvin temperature scale
- kilogram
- length
- mass
- meter
- second
- temperature

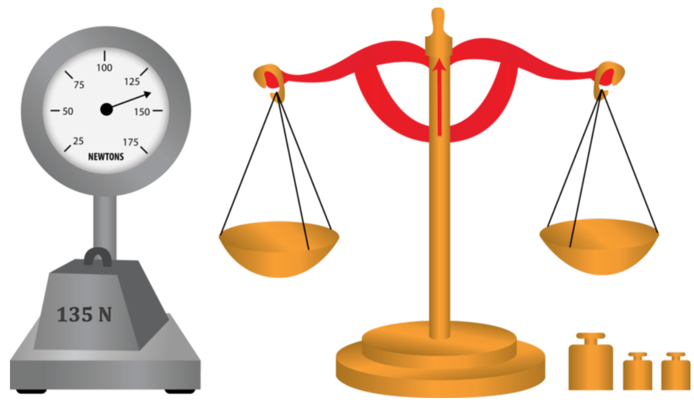
Introduction

The **International System of Units**, abbreviated SI from the French *Le Système International d' Unites*, is the main system of measurement units used in science. Since the 1960s, the International System of Units has been agreed upon internationally as the standard metric system. The SI base units are based on physical standards. The definitions of the SI base units have been and continue to be modified, and new base units are added as advancements in science are made. Each SI base unit, except the kilogram, is described by stable properties of the universe.

Mass and Its SI Unit

Mass and weight are not the same thing. Although we often use these terms interchangeably, each one has a specific definition and usage. The **mass** of an object is a measure of the amount of matter in it and remains the same regardless of where the object is placed. For example, moving a brick to the moon does not cause matter in the brick to disappear or to be removed. The weight of an object is the force of attraction between the object and the Earth (or whatever large, gravity-producing body the object is located on). This attraction is due to the force of gravity. Since the force of gravity is not the same at every point on the Earth's surface, the weight of an object is not constant. The gravitational pull on the object varies and depends on where the object is with respect to the Earth or other gravity-producing object. For example, a man who weighs 180 pounds on Earth would weigh only 45 pounds if he were in a stationary position 4,000 miles above the Earth's surface. This same man would weigh only 30 pounds on the moon, because the moon's gravitational pull is one-sixth that of Earth's. The mass of this man, however, would be the same in each situation because the amount of matter in the man is constant.

We measure weight with a scale, which contains a spring that compresses when an object is placed on it. An illustration of a scale is depicted on the left in the diagram below. If the gravitational pull is less, the spring compresses less and the scale shows less weight. We measure mass with a balance, depicted on the right in the diagram below. A balance compares the unknown mass to known masses by balancing them on a lever. If we take our balance and known masses to the moon, an object will have the same measured mass that it had on the Earth. The weight, of course, would be different on the moon. Consistency requires that scientists use mass and not weight when measuring the amount of matter.



The basic unit of mass in the International System of Units is the kilogram. A kilogram is equal to 1,000 grams. A gram is a relatively small amount of mass, so larger masses are often expressed in kilograms. When very tiny amounts of matter are measured, we often use milligrams, with one milligram equal to 0.001 gram. Other larger, smaller, or intermediate mass units may also be appropriate.

At the end of the 18th century, a kilogram was the mass of a cubic decimeter of water. In 1889, a new international prototype of the kilogram was made from a platinum-iridium alloy. The kilogram is equal to the mass of this international prototype, which is held in Paris, France. A copy of the standard kilogram is shown in **Figure** below.

This image shows a copy of the standard kilogram stored and used in Denmark.

Length and Its SI Unit

Length is the measurement of anything from end to end. In science, length usually refers to how long an object is. There are many units and sets of standards used in the world for measuring length. The ones familiar to you are probably inches, feet, yards, and miles. Most of the world, however, measure distances in meters and kilometers for longer distances, and in centimeters and millimeters for shorter distances. For consistency and ease of communication, scientists around the world have agreed to use the SI standards, regardless of the length standards used by the general public.

This image shows the standard meter used in France in the 18th century.

The SI unit of length is the **meter**. In 1889, the definition of the meter was the length of a bar made of platinum-iridium alloy stored under conditions specified by the International Bureau of Standards. In 1960, this definition of the standard meter was replaced by a definition based on a wavelength of krypton-86 radiation. In 1983, that definition was replaced by the following: the meter is the length of the path traveled by light in a vacuum during a time interval of $\frac{1}{299,792,458}$ of a second.

Measuring Temperature

Temperature represents the average kinetic energy of the particles making up a material. A thermometer is a device that measures temperature. The name is made up of *thermo*, which means heat, and *meter*, which means to measure. One of the earliest inventors of a thermometer was Galileo. He is said to have used a device called a thermoscope around the year 1600. The thermometers we typically use today, however, are different from the one Galileo used.

Daniel Fahrenheit established the Fahrenheit scale. On his temperature scale, Fahrenheit designated the freezing point of water as $32°$Fand the boiling point of water as $212°$F. Therefore, the distance between these two points would be divided into 180 degrees. The Fahrenheit temperature scale is used in the United States for most daily expressions of temperature. In another temperature scale established by Anders Celsius, Celsius designated the freezing point of water as $0°C$ and the boiling point of water as 100ºC. Therefore, the temperatures between these two points on the Celsius scale are divided into 100 degrees. Clearly, the size of a Celsius degree and the size of a Fahrenheit degree are not the same.

A third temperature scale was established to address this issue. This temperature scale was designed by Lord Kelvin. Lord Kelvin stated that there is no upper limit to how hot things can get, but there is a limit as to how cold things can get. Kelvin developed the idea of absolute zero, which is the temperature that molecules stop moving and have zero kinetic energy. The **Kelvin temperature scale** has its zero at absolute zero (determined to be $-273.15°C$) and uses the same degree size as a degree on the Celsius scale. As a result, the mathematical relationship between the Celsius scale and the Kelvin scale is: K = ºC +273.15. On the Kelvin scale, water freezes at 273.15 K and boils at 373.15 K. In the case of the Kelvin scale, the degree sign is not used. Temperatures are expressed, for example, simply as 450 K.

Time and Its SI Unit

The SI unit for time is the second. The second was originally defined as a tiny fraction of the time required for the Earth to orbit the Sun. It has since been redefined several times. The definition of a **second** (established in 1967 and reaffirmed in 1997) is: the duration of 9,192,631,770 periods of radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom.

LESSON SUMMARY

- The International System of Units, abbreviated SI from the French *Le Système International d' Unites*, is internationally agreed upon since the 1960s as the standard metric system.

Mass and weight:
- The mass of an object is a measure of the amount of matter in it.
- The mass of an object remains the same regardless of where the object is placed.
- The basic unit of mass in the International System of Units is the kilogram.
- The weight of an object is the force of attraction between the object and the earth (or whatever large, gravity-producing body the object is located on).

Length:
- Length is the measurement of anything from end to end.
- The SI unit of length is the meter.

Volume:
- The volume of an object is the amount of space it takes up.
- The cubic meter is the SI unit of volume.

Temperature and heat:
- Temperature represents the average kinetic energy of the particles that make up a material.
- The Kelvin temperature scale has its zero at absolute zero (determined to be –273.15°C) and uses the same degree size as the Celsius scale.
- The mathematical relationship between the Celsius scale and the Kelvin scale is $K = °C + 273.15$.

Time:
- The SI unit for time is the second.

REVIEW QUESTIONS

1. What is the basic unit of measurement in the metric system for length?
2. What is the basic unit of measurement in the metric system for mass?
3. Give the temperatures in Celsius for the freezing and boiling points of water.
4. Give the temperatures in Kelvin for the freezing and boiling points of water.
5. Would it be comfortable to swim in a swimming pool whose water temperature is 275 K? Why or why not?

LESSON OBJECTIVES

The student will:

- construct conversion factors from equivalent measurements.
- apply the techniques of dimensional analysis to solving problems.
- perform metric conversions using dimensional analysis.

Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another unit. A simple conversion factor can be used to convert meters into centimeters, or a more complex one can be used to convert miles per hour into meters per second. Since most calculations require measurements to be in certain units, you will find many uses for conversion factors. What must always be remembered is that a conversion factor has to represent a fact; because the conversion factor is a fact and not a measurement, the numbers in a conversion factor are exact. This fact can either be simple or complex. For instance, you probably already know the fact that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^5 miles/second. Either one of these can be used as a conversion factor, depending on the type of calculation you might be working with.

Dimensional Analysis

Frequently, it is necessary to convert units measuring the same quantity from one form to another. For example, it may be necessary to convert a length measurement in meters to millimeters. This process is quite simple if you follow a standard procedure called dimensional analysis (also known as unit analysis or the factor-label method). **Dimensional analysis** is a technique that involves the study of the dimensions (units) of physical quantities. It is a convenient way to check mathematical equations. Dimensional analysis involves considering the units you presently have and the units you wish to end up with, as well as designing conversion factors that will cancel units you don't want and produce units you do want. The conversion factors are created from the equivalency relationships between the units. For example, one unit of work is a newton meter (abbreviated N $\cdot$m). If you have measurements in newtons (a unit for force, *F*) and meters (a unit for distance, *d*), how would you calculate work? An analysis of the units will tell you that you should multiply force times distance to get work:

$$W = F \times d$$

Suppose you want to convert 0.0856 meters into millimeters. In this case, you need only one conversion factor that will cancel the meters unit and create the millimeters unit. The conversion factor will be created from the relationship:

$$1000\text{mL} = 1\text{ m}$$

In the above expression, the meter units will cancel and only the millimeter unit will remain.

Vocabulary

- conversion factor
- dimensional analysis

Example:

Convert 1.53 g to cg.

The equivalency relationship is 1.00 g = 100cg, so the conversion factor is constructed from this equivalency in order to cancel grams and produce centigrams.

$$(1.53 \text{ g}) \cdot (\frac{100 \text{ cg}}{1 \text{ g}}) = 153 \text{ cg}$$

Example:

Convert 1000 in. to ft.

The equivalency between inches and feet is 12 in. = 1 ft. The conversion factor is designed to cancel inches and produce feet.

$$(1000. \text{ in.}) \cdot (\frac{1 \text{ ft}}{12 \text{ in.}}) = 83.33 \text{ ft}$$

Each conversion factor is designed specifically for the problem. In the case of the conversion above, we need to cancel inches, so we know that the inches component in the conversion factor needs to be in the denominator.

Example:

Convert 425 klums to piks given the equivalency relationship

$$10 \; klums = 1 \; pik$$

$$(425 \; klums) \cdot (\frac{1 \; pik}{10 \; klums}) = 42.5 \; piks$$

Sometimes, it is necessary to insert a series of conversion factors. Suppose we need to convert miles to kilometers, and the only equivalencies we know are 1 mi = 5,280 ft, 12 in = 1 ft, 2.54 cm = 1 in., 100 cm = 1 m, and 1000 m = 1 km. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence.

Example:

Convert 12 mi to km.
In each step, the previous unit is canceled and the next unit in the sequence is produced. Conversion factors for area and volume can also be produced by this method.

Example:

Convert 1500 cm^2 *to* m^2.

Example:

Convert 12 in^3 *to* cm^3.

LESSON SUMMARY

- Conversion factors are used to convert one unit of measurement into another unit.
- Dimensional analysis involves considering both the units you presently have and the units you wish to end up with, as well as designing conversion factors that will cancel units you don't want and produce units you do want.

FURTHER READING / SUPPLEMENTAL LINKS

This website provides a math skills review on dimensional analysis.
http://www.chem.tamu.edu/class/fyp/mathrev/mr-da.html

REVIEW QUESTIONS

1. The density of a certain solid is measured and found to be $12.68\ \mathrm{g/mL}$. Convert this measurement into kg/L .
2. In a nuclear chemistry experiment, an alpha particle is found to have a velocity of 14,285 m/s. Convert this measurement into miles/hour (mi/h).

SCIENTIFIC NOTATION

Objectives:

- use scientific notation to express large and small numbers.

Vocabulary

- scientific notation

Introduction

Work in science frequently involves very large and very small numbers. The speed of light, for example, is 300,000,000 m/s; the mass of the earth is 6,000,000,000,000,000,000,000,000 kg; and the mass of an electron is 0.0000000000000000000000000000009 kg. It is very inconvenient to write out such numbers and even more inconvenient to attempt to carry out mathematical operations with them. Scientists and mathematicians have designed an easier method to deal with such long numbers. This more convenient system is called exponential notation by mathematicians and scientific notation by scientists.

What is Scientific Notation?

In **scientific notation**, very large and very small numbers are expressed as the product of a number between 1 and 10 multiplied by some power of 10. For example, the number 9,000,000 can be written as the product of 9 times 1,000,000. In turn, 1,000,000 can be written as 10^6. Therefore, 9,000,000 can be written as 9×10^6. In a similar manner, 0.00000004 can be written as 4 times $\frac{1}{10^8}$, or 4×10^{-8}.

This is called the "coefficient." → $\boxed{6.5} \times 10^{\boxed{4}}$ ← This is called the "exponent."

Examples of Scientific Notation	
Decimal Notation	Scientific Notation
$95,672$	9.5672×10^4
$8,340$	8.34×10^3
100	1×10^2
7.21	7.21×10^0
0.014	1.4×10^{-2}
0.0000000080	8.0×10^{-9}
0.00000000000975	9.75×10^{-12}

As you can see from the examples in **Table** above, to convert a number from decimal form into scientific notation, you count the number of spaces needed to move the decimal, and that number becomes the exponent of 10. If you are moving the decimal to the left, the exponent is positive, and if you are moving the decimal to the right, the exponent is negative. You should note that *all significant figures are maintained in scientific notation.* You will probably realize that the greatest advantage of using scientific notation occurs when there are many non-significant figures.

LESSON SUMMARY

- Very large and very small numbers in science are expressed in scientific notation.

REVIEW QUESTIONS

1. Write the following numbers in scientific notation.

    a. 0.0000479

    b. $251,000,000$

    c. $4,260$

    d. 0.00206

PERIODIC TABLE

CHAPTER 1

Standard 1: Students will understand that all matter in the universe has a common origin and is made of atoms, which have structure and can be systematically arranged on the period table.

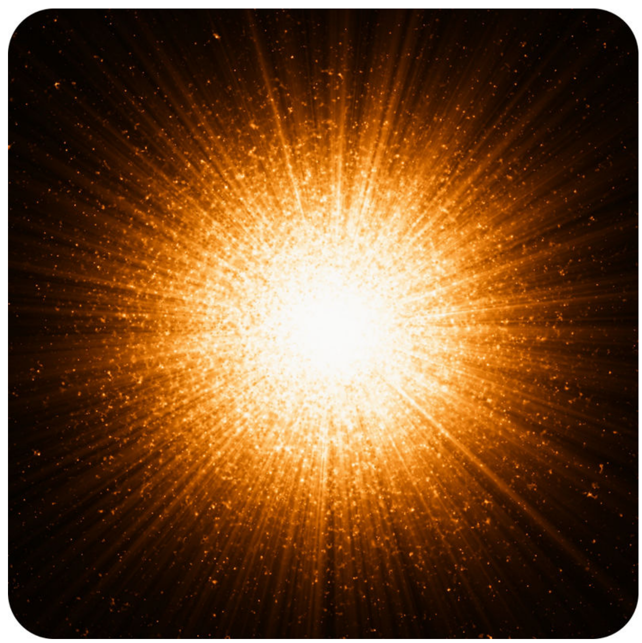
Standard 1, Objective 1: Recognize the origin and distribution of elements in the universe.

Objectives
- Identify evidence supporting the Big Bang theory
- Recognize that all matter in the universe and on earth is composed of the same elements.
- Identify the distribution of elements in the universe and compare the occurrence of heavier elements on earth and in the universe.

Introduction

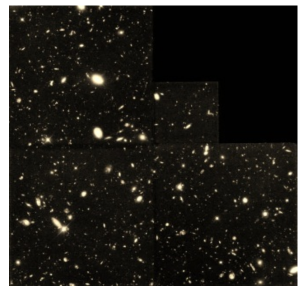
The Big Bang is the currently accepted theory of the early development of the universe. Cosmologists study the origin of the universe and use the term Big Bang to illustrate the idea that the universe was originally an extremely hot and dense point in the space at some finite time in the past and has since cooled by expanding to the present state. The universe continues to expand today. The theory is supported by the most comprehensive and accurate explanations from current scientific evidence and observation.

According to the theory, the Universe would have cooled sufficiently to allow energy to be converted into subatomic particles (protons, neutrons, and electrons and many other particles). While protons and neutrons would have formed the first atomic nuclei only a few minutes after the Big Bang, it would then have taken thousands of years for electrons to lose enough energy to form neutral atoms. The first element produced would be hydrogen. Giant clouds of hydrogen would then form stars and galaxies. Other elements were formed by fusion (combining small atoms together to make larger atoms) within the stars.



Evidence for the Big Bang Theory

Many scientists have contributed to gathering evidence and developing theories to contribute to our understanding of the origin of the universe and the Big Bang Theory. Georges Lemaître, a Belgian priest, physicist, and astronomer was the first person to propose the hypothesis of the "primeval atom" which later became known as the Big Bang Theory. Lemaître's hypothesis used the work of earlier astronomers and proposed that the inferred moving away of galaxies were due to the expansion of the Universe.
More evidence of the expanding universe was provided by Alexander Friedmann, Russian cosmologist and mathematician. He derived the "Friedmann" equations from Albert Einstein's equations of general relativity, showing that the Universe might be expanding in contrast to the static universe model advocated by Einstein at that time. Albert Einstein had found that his newly developed theory of general relativity indicated that the universe must be either expanding or contracting. Unable to believe what his own equations were telling him, Einstein introduced a "fudge factor" to the equations to avoid this "problem". When Einstein heard of Hubble's discovery, he said that changing his equations was "the biggest blunder of his life."



Edwin Hubble is regarded as the leading observational cosmologist of the 1900s. He is credited with the discovery of galaxies other than the Milky Way. In 1929 Hubble presented evidence that galaxies were moving away from each other and that galaxies that are further away are moving faster, as was first suggested by Lemaître in 1927. Hubble's evidence is now known as red shift. This discovery was the first observational support for the Big Bang Theory. If the distance between galaxies is increasing today, then galaxies and everything else in the universe must have been closer together in the past. Think about the expanding universe, and then reverse it. If we start at the present and go back into the past, the universe gets smaller. What is the end result of a contracting universe? A point. In the very distant past, the universe must have indeed been extremely small and had extreme densities and temperatures.

The opponents to Big Bang Theory argued that if the universe had existed as a point in space, large amounts of radiation would have been produced as the subatomic particles formed from the cooling and expanding energy. After cosmic microwave background radiation was discovered in 1964 and the analysis matched the amount of missing radiation from the Big Band, most scientists were fairly convinced by the evidence that some Big Bang scenario must have occurred.

In the last quarter century, large particle accelerators have been built to provide significant confirmation of the Big Bang Theory. Several particles have been discovered which support the idea that energy can be converted to particles which combine to form protons. Although these accelerators have limited capabilities when probing into such high energy regimes, significant evidence continues to support the Big Bang Theory.

Elements and Big Bang Theory

If the Big Bang theory was correct, scientists predicted that they should still find most of the universe to be still composed of the hydrogen that was formed in the first few minutes after the big bang as the universe cooled and expanded. The observed abundances of hydrogen and other very light elements throughout the universe closely match the calculated predictions for the formation of these elements from the rapid expansion and cooling in the first minutes of the universe. Over 90% of the entire universe is composed of the lightest of the elements, hydrogen and helium. The heavier elements, from helium to iron were formed from fusion within stars. The same elements that the earth is made of are found throughout the universe. Fred Hoyle, who originally criticized Big Bang Theory, provided an explanation of nuclear fusion in stars as that later helped considerably in the effort to describe how heavier elements were formed from the initial hydrogen.

The earth consists of much heavier elements. The most abundant elements in the earth's crust include oxygen, silicon, and aluminum. These elements were formed by fusion of the earliest (and heaviest stars) formed. The core of the earth is primarily iron. This iron was also formed is these very early, heavy stars. The radioactive elements found on the earth were most probably formed as these heavy stars died the violent death known as supernovae. The iron (and other elements near it on the periodic table) were thrown into the void of space with very high speeds allowing them to form still heavier elements by a similar process to which elements heavier than uranium (artificial or man-made) elements have been formed during the 20th century.

By analyzing the light given off by stars throughout the universe, scientists have been able to determine the elements that make up various objects in the universe. When we look at the light from the sun and other stars in the universe, we see the same types of spectral lines (unique patterns of light given off by an element's electrons) present throughout the universe. From this, we can conclude that matter exists in all stars and in our galaxy in the same general amounts. The current universe contains about 80% Hydrogen, then about 10% Helium, and 2% Oxygen. The remainder of the elements makes up the other 8%. The earth, however, is composed of a greater abundance of the heavier elements. There is about 49% Oxygen, 25% Silicon, and about 7% Aluminum, with the remainder of the elements making up about 19%.

(Source: http://en.wikipedia.org/wiki/Abundance_of_the_chemical_elements and http://www.chem.wisc.edu/deptfiles/genchem/sstutorial/Text3/Tx33/tx33.html)

Lesson Summary

- The Big Bang Theory proposes that all matter in the universe was once contained in a small point, but has since expanded and cooled.
- The theory is supported by scientists as it provides a satisfactory explanation for the observations that the universe is expanding today, that the universe is composed mostly of hydrogen and oxygen, cosmic background radiation, etc.
- The theory also provides as explanation for where elements heavier than hydrogen were formed, through fusion into heavier elements

Vocabulary

- Big Bang Theory: the idea that the universe was originally extremely hot and dense at some finite time in the past and has since cooled by expanding to the present state and continues to expand today
- Cosmic background radiation: energy in the form of radiation leftover from the early big bang
- Atom: smallest part of an element that can't be broken down by ordinary means
- Element: same type of atom that behaves in a certain way, such as Oxygen.
- Nucleus: the center of an atom, containing one or more protons and neutrons
- Proton: positively charge particle in the nucleus of an atom that has a relative mass of 1 atomic mass unit.
- Neutron: particle with no charge in the nucleus of an atom that has a relative mass of 1 atomic mass unit.
- Electron: negatively charged particle that orbits the nucleus of an atom at high speed.
- It also has a mass of 1/1840th atomic mass unit. It doesn't weigh very much!
- Fusion: joining of two nuclei under high pressure and temperature.

Review Questions

1. Why is the abundance of hydrogen and helium so important in accepting Big Bang Theory?
2. What evidence exists that the Big Bang did occur? How do these evidences support the theory?
3. Earth (and the other inner planets) contains large amounts of elements heavier than carbon. Where did these elements come from?
4. The Big Bang is considered a theory. Lemaître's work is considered a hypothesis. Hubble is known for the law of comic expansion. Compare and contrast these three concepts. Why is one considered a hypothesis, one a theory, and still another a law?

Further Reading / Supplementary Links

To see a video documenting the early history of the concept of the atom, go to http://www.uen.org/dms/ Go to the k-12 library. Search for "Stephen Hawking". Watch program 2: In The Beginning.

(you can get the username and password from your teacher)

STANDARD 1, OBJECTIVE 2: RELATE THE STRUCTURE, BEHAVIOR, AND SCALE OF AN ATOM TO THE PARTICLES THAT COMPOSE IT.

Objectives

- Give a short history of the concept of the atom.
- Describe the contributions of Democritus and Dalton to atomic theory.
- Summarize Dalton's atomic theory and explain its historical development.

EARLY IDEAS OF ATOMS INTRODUCTION

All matter in the universe is made out of tiny building blocks called atoms. All modern scientists accept the concept of the atom, but when the concept of the atom was first proposed about 2,500 years ago, ancient philosophers laughed at the idea. It has always been difficult to convince people of the existence of things that are too small to see. We will spend some time considering the evidence (observations) that convince scientists of the existence of atoms.

Democritus was known as "The Laughing Philosopher." It's a good thing he liked to laugh, because most other philosophers were laughing at his theories.

Democritus and the Greek Philosophers

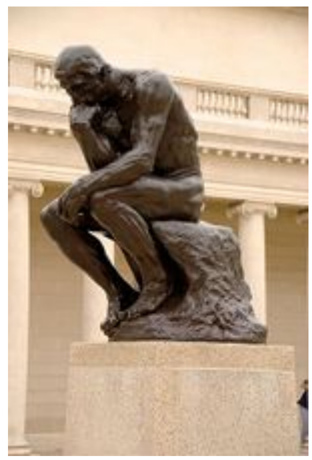
Before we discuss the experiments and evidence that have, over the years, convinced scientists that matter is made up of atoms, it's only fair to give credit to the man who proposed "atoms" in the first place. About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge, entity. In other words, "everything was one." They believed that all objects, all matter, and all substances were connected as a single, big, unchangeable "thing."

One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that atomos, or atomon—tiny, indivisible, solid objects - make up all matter in the universe.

Democritus then reasoned that changes occur when the many atomos in an object were reconnected or recombined in different ways. Democritus even extended his theory, suggesting that there were different varieties of atomos with different shapes, sizes, and masses. He thought, however, that shape, size and mass were the only properties differentiating the different types of atomos. According to Democritus, other characteristics, like color and taste, did not reflect properties of the atomos themselves, but rather, resulted from the different ways in which the atomos were combined and connected to one another.

Greek philosophers truly believed that, above all else, our understanding of the world should rely on "logic." In fact, they argued that the world couldn't be understood using our senses at all, because our senses could deceive us. Therefore, instead of relying on observation, Greek philosophers tried to understand the world using their minds and, more specifically, the power of reason.

So how could the Greek philosophers have known that Democritus had a good idea with his theory of "atomos?" It would have taken some careful observation and a few simple experiments. Now you might wonder why Greek philosophers didn't perform any experiments to actually test Democritus' theory. The problem, of course, was that Greek philosophers didn't believe in experiments at all. Remember, Greek philosophers didn't trust their senses, they only trusted the reasoning power of the mind.



Greek philosophers tried to understand the nature of the world through reason and logic but not through experiment and observation.

The early Greek philosophers tried to understand the nature of the world through reason and logic, but not through experiment and observation. As a result, they had some very interesting ideas, but they felt no need to justify their ideas based on life experiences. In a lot of ways, you can think of the Greek philosophers as being "all thought and no action." It's truly amazing how much they achieved using their minds, but because they never performed any experiments, they missed or rejected a lot of discoveries that they could have made otherwise. Greek philosophers dismissed Democritus' theory entirely. Sadly, it took over two millennia before the theory of atomos (or "atoms," as they're known today) was fully appreciated.

Dalton's Atomic Theory

Although the concept of atoms is now widely accepted, this wasn't always the case. Scientists didn't always believe that everything was composed of small particles called atoms. The work of several scientists and their experimental data gave evidence for what is now called the atomic theory.

In the late 1700s, Antoine Lavoisier, a French scientist, experimented with the reactions of many metals. He carefully measured the mass of a substance before reacting and again measured the mass after a reaction had occurred in a closed system (meaning that nothing could enter or leave the container). He found that no matter what reaction he looked at, the mass of the starting materials was always equal to the mass of the ending materials. This is now called the law of conservation of mass. This went contrary to what many scientists at the time thought. For example, when a piece of iron rusts, it appears to gain mass. When a log is burned, it appears to lose mass. In these examples, though, the reaction does not take place in a closed container and substances, such as the gases in the air, are able to enter or leave. When iron rusts, it is combining with oxygen in the air, which is why it seems to gain mass. What Lavoisier found was that no mass was actually being gained or lost. It was coming from the air. This was a very important first step in giving evidence for the idea that everything is made of atoms. The atoms (and mass) are not being created or destroyed. The atoms are simply reacting with other atoms that are already present.

In the late 1700s and early 1800s, scientists began noticing that when certain substances, like hydrogen and oxygen, were combined to produce a new substance, like water, the reactants (hydrogen and oxygen) always reacted in the same proportions by mass. In other words, if 1 gram of hydrogen reacted with 8 grams of oxygen, then 2 grams of hydrogen would react with 16 grams of oxygen, and 3 grams of hydrogen would react with 24 grams of oxygen. Strangely, the observation that hydrogen and oxygen always reacted in the "same proportions by mass" wasn't special. In fact, it turned out that the reactants in every chemical reaction reacted in the same proportions by mass. This observation is summarized in the law of definite proportions. Take, for example, nitrogen and hydrogen, which react to produce ammonia. In chemical reactions, 1 gram of hydrogen will react with 4.7 grams of nitrogen, and 2 grams of hydrogen will react with 9.4 grams of nitrogen. Can you guess how much nitrogen would react with 3 grams of hydrogen? Scientists studied reaction after reaction, but every time the result was the same. The reactants always reacted in the same proportions.

Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

At the same time that scientists were finding this pattern out, a man named John Dalton was experimenting with several reactions in which the reactant elements formed more than one type of product, depending on the experimental conditions he used. One common reaction that he studied was the reaction between carbon and oxygen. When carbon and oxygen react, they produce two different substances – we'll call these substances "A" and "B." It turned out that, given the same amount of carbon, forming B always required exactly twice as much oxygen as forming A. In other words, if you can make A with 3 grams of carbon and 4 grams of oxygen, B can be made with the same 3 grams of carbon, but with 8 grams oxygen. Dalton asked himself – why does B require 2 times as much oxygen as A? Why not 1.21 times as much oxygen, or 0.95 times as much oxygen? Why a whole number like 2?

The situation became even stranger when Dalton tried similar experiments with different substances. For example, when he reacted nitrogen and oxygen, Dalton discovered that he could make three different substances – we'll call them "C," "D," and "E." As it turned out, for the same amount of nitrogen, D always required twice as much oxygen as C. Similarly, E always required exactly four times as much oxygen as C. Once again, Dalton noticed that small whole numbers (2 and 4) seemed to be the rule. This observation came to be known as the law of multiple proportions.

Dalton thought about his results and tried to find some theory that would explain it, as well as a theory that would explain the Law of Conservation of Mass (mass is neither created nor destroyed, or the mass you have at the beginning is equal to the mass at the end of a change). One way to explain the relationships that Dalton and others had observed was to suggest that materials like nitrogen, carbon and oxygen were composed of small, indivisible quantities which Dalton called "atoms" (in reference to Democritus' original idea). Dalton used this idea to generate what is now known as Dalton's Atomic Theory which stated the following:

- Matter is made of tiny particles called atoms.
- Atoms are indivisible (can't be broken into smaller particles). During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed.
- All atoms of a given element are identical in mass and other properties.
- The atoms of different elements differ in mass and other properties.
- Atoms of one element can combine with atoms of another element to form "compounds" – new, complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers.

LESSON SUMMARY

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called "atomos."
- Other Greek philosophers disliked Democritus' "atomos" theory because they felt it was illogical.
- Dalton used observations about the ratios in which elements will react to combine and The Law of Conservation of Mass to propose his Atomic Theory.
- Dalton's Atomic Theory states:

    1. Matter is made of tiny particles called atoms.
    2. Atoms are indivisible. During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed.
    3. All atoms of a given element are identical in mass and other properties.
    4. The atoms of different elements differ in mass and other properties.
    5. Atoms of one element can combine with atoms of another element to form "compounds" – new complex particles. In a given compound, however, the different types of atoms are always present in the same relative numbers.

VOCABULARY

- Atom: Democritus' word for the tiny, indivisible, solid objects that he believed made up all matter in the universe
- Dalton's Atomic Theory: the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom

Further Reading / Supplemental Links
To see a video documenting the early history of the concept of the atom, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "history of the atom". Watch part 01. (you can get the username and password from your teacher)
Vision Learning: From Democritus to Dalton: http://visionlearning.com/library/module_viewer.php?c3=&mid=49&l=

REVIEW QUESTIONS

1. (Multiple choice) Which of the following is not part of Dalton's Atomic Theory?
    a) matter is made of tiny particles called atoms.
    b) during a chemical reaction, atoms are rearranged.
    c) during a nuclear reaction, atoms are split apart.
    d) all atoms of a specific element are the same.

2. Democritus and Dalton both suggested that all matter was composed of small particles, called atoms. What is the greatest advantage Dalton's Atomic Theory had over Democritus'?
3. It turns out that a few of the ideas in Dalton's Atomic Theory aren't entirely correct. Are inaccurate theories an indication that science is a waste of time?

Further Understanding of the Atom

Objectives

- Explain the observations that led to Thomson's discovery of the electron.
- Describe Thomson's "plum pudding" mode of the atom and the evidence for it
- Draw a diagram of Thomson's "plum pudding" model of the atom and explain why it has this name.
- Describe Rutherford's gold foil experiment and explain how this experiment altered the "plum pudding" model.
- Draw a diagram of the Rutherford model of the atom and label the nucleus and the electron cloud.

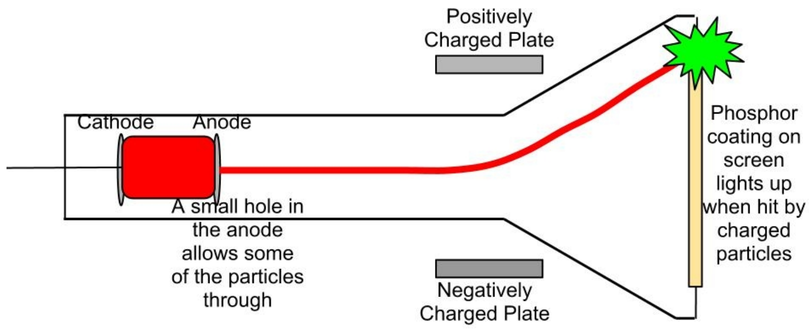
Introduction

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory wasn't the entire story. As it turns out, Dalton had a lot right. He was right in saying matter is made up of atoms; he was right in saying there are different kinds of atoms with different mass and other properties; he was "almost" right in saying atoms of a given element are identical; he was right in saying during a chemical reaction, atoms are merely rearranged; he was right in saying a given compound always has atoms present in the same relative numbers. But he was WRONG in saying atoms were indivisible or indestructible. As it turns out, atoms are divisible. In fact, atoms are composed of even smaller, more fundamental particles. These particles, called subatomic particles, are particles that are smaller than the atom. We'll talk about the discoveries of these subatomic particles next.

J.J. Thomson conducted experiments that suggested that Dalton's atomic theory wasn't telling the entire story.

Thomson's Plum Pudding Model

In the mid-1800s, scientists were beginning to realize that the study of chemistry and the study of electricity were actually related. First, a man named Michael Faraday showed how passing electricity through mixtures of different chemicals could cause chemical reactions. Shortly after that, scientists found that by forcing electricity through a tube filled with gas, the electricity made the gas glow! Scientists didn't, however, understand the relationship between chemicals and electricity until a British physicist named J. J. Thomson began experimenting with what is known as a cathode ray tube.



Thomson's experiment with cathode rays found that the ray moved away from negatively charged plates and toward positively charges plates. What does this say about the charge of the ray? Image by Tracy Poulsen

The figure shows a basic diagram of a cathode ray tube like the one J. J. Thomson would have used. A cathode ray tube is a small glass tube with a cathode (a negatively charged metal plate) and an anode (a positively charged metal plate) at opposite ends. By separating the cathode and anode by a short distance, the cathode ray tube can generate what are known as cathode rays – rays of electricity that flow from the cathode to the anode. J. J. Thomson wanted to know what cathode rays were, where cathode rays came from, and whether cathode rays had any mass or charge. The techniques that J. J. Thomson used to answer these questions were very clever and earned him a Nobel Prize in physics. First, by cutting a small hole in the anode, J. J. Thomson found that he could get some of the cathode rays to flow through the hole in the anode and into the other end of the glass cathode ray tube. Next, J. J. Thomson figured out that if he painted a substance known as "phosphor" onto the far end of the cathode ray tube, he could see exactly where the cathode rays hit because the cathode rays made the phosphor glow.

J. J. Thomson must have suspected that cathode rays were charged, because his next step was to place a positively charged metal plate on one side of the cathode ray tube and a negatively charged metal plate on the other side of the cathode ray tube, as shown in Figure 3. The metal plates didn't actually touch the cathode ray tube, but they were close enough that a remarkable thing happened! The flow of the cathode rays passing through the hole in the anode was bent upwards towards the positive metal plate and away from the negative metal plate. Using the "opposite charges attract, like charges repel" rule, J. J. Thomson argued that if the cathode rays were attracted to the positively charged metal plate and repelled from the negatively charged metal plate, they themselves must have a negative charge!

J. J. Thomson then did some rather complex experiments with magnets, and used his results to prove that cathode rays were not only negatively charged, but also had mass.

Remember that anything with mass is part of what we call matter. In other words, these cathode rays must be the result of negatively charged "matter" flowing from the cathode to the anode. But there was a problem. According to J. J. Thomson's measurements, either these cathode rays had a ridiculously high charge, or else had very, very little mass – much less mass than the smallest known atom. How was this possible? How could the matter making up cathode rays be smaller than an atom if atoms were indivisible? J. J. Thomson made a radical proposal: maybe atoms are divisible. J. J. Thomson suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as electrons. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.

Now imagine what would happen if atoms were made entirely of electrons. First of all, electrons are very, very small; in fact, electrons are about 2,000 times smaller than the smallest known atom, so every atom would have to contain a whole lot of electrons. But there's another, even bigger problem: electrons are negatively charged. Therefore, if atoms were made entirely out of electrons, atoms would be negatively charged themselves… and that would mean all matter was negatively charged as well. Of course, matter isn't negatively charged. In fact, most matter is what we call neutral – it has no charge at all. If matter is composed of atoms, and atoms are composed of negative electrons, how can matter be neutral? The only possible explanation is that atoms consist of more than just electrons. Atoms must also contain some type of positively charged material that balances the negative charge on the electrons. Negative and positive charges of equal size cancel each other out, just like negative and positive numbers of equal size. What do you get if you add +1 and -1? You get 0, or nothing. That's true of numbers, and that's also true of charges. If an atom contains an electron with a -1 charge, but also some form of material with a +1 charge, overall the atom must have a (+1) + (-1) = 0 charge – in other words, the atom must be neutral, or have no charge at all.

Thomson's plum pudding model was much like a chocolate chip cookie. Notice how the chocolate chips are the negatively charged electrons, while the positive charge is spread throughout the entire batter.

Based on the fact that atoms are neutral, and based on J. J. Thomson's discovery that atoms contain negative subatomic particles called "electrons," scientists assumed that atoms must also contain a positive substance. It turned out that this positive substance was another kind of subatomic particle, known as the proton. Although scientists knew that atoms had to contain positive material, protons weren't actually discovered, or understood, until quite a bit later.

When Thomson discovered the negative electron, he realized that atoms had to contain positive material as well – otherwise they wouldn't be neutral overall. As a result, Thomson formulated what's known as the "plum pudding" model for the atom. According to the "plum pudding" model, the negative electrons were like pieces of fruit and the positive material was like the batter or the pudding. This made a lot of sense given Thomson's experiments and observations. Thomson had been able to isolate electrons using a cathode ray tube; however he had never managed to isolate positive particles. As a result, Thomson theorized that the positive material in the atom must form something like the "batter" in a plum pudding, while the negative electrons must be scattered through this "batter." (If you've never seen or tasted a plum pudding, you can think of a chocolate chip cookie instead. In that case, the positive material in the atom would be the "batter" in the chocolate chip cookie, while the negative electrons would be scattered through the batter like chocolate chips.)

Notice how easy it would be to pick the pieces of fruit out of a plum pudding. On the other hand, it would be a lot harder to pick the batter out of the plum pudding, because the batter is everywhere. If an atom were similar to a plum pudding in which the electrons are scattered throughout the "batter" of positive material, then you'd expect it would be easy to pick out the electrons, but a lot harder to pick out the positive material.

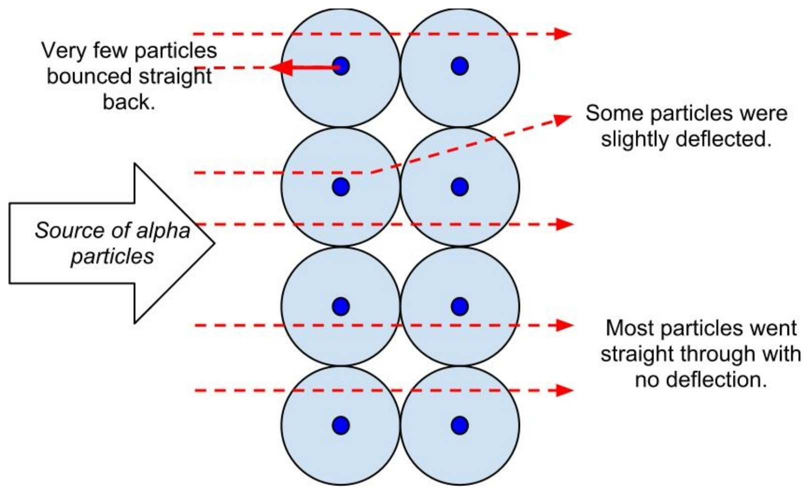
J.J. Thomson had measured the charge to mass ratio of the electron, but had been unable to accurately measure the charge on the electron. With his oil drop experiment, Robert Millikan was able to accurately measure the charge of the electron. When combined with the charge to mass ratio, he was able to calculate the mass of the electron. What Millikan did was to put a charge on tiny droplets of oil and measured their rate of descent. By varying the charge on different drops, he noticed that the electric charges on the drops were all multiples of 1.6x10-19C, the charge on a single electron.

Ernest Rutherford

Rutherford's Nuclear Model

Everything about Thomson's experiments suggested the "plum pudding" model was correct – but according to the scientific method, any new theory or model should be tested by further experimentation and observation. In the case of the "plum pudding" model, it would take a man named Ernest Rutherford to prove it inaccurate. Rutherford and his experiments will be the topic of the next section.

Disproving Thomson's "plum pudding" model began with the discovery that an element known as uranium emits positively charged particles called alpha particles as it undergoes radioactive decay. Radioactive decay occurs when one element decomposes into another element. It only happens with a few very unstable elements. Alpha particles themselves didn't prove anything about the structure of the atom, they were, however, used to conduct some very interesting experiments.



Ernest Rutherford's Gold Foil Experiment in which alpha particles were shot at a piece of gold foil. Most of the particles went straight through, but some bounced straight back, indicating they were hitting a very small, very dense particle in the atom.

Image by Tracy Poulsen

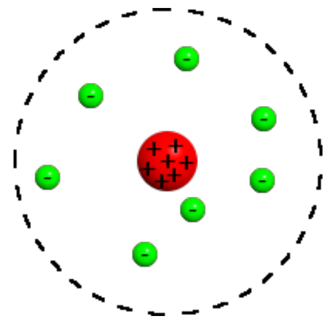
Ernest Rutherford was fascinated by all aspects of alpha particles. For the most part, though, he seemed to view alpha particles as tiny bullets that he could use to fire at all kinds of different materials. One experiment in particular, however, surprised Rutherford, and everyone else.

Rutherford found that when he fired alpha particles at a very thin piece of gold foil, an interesting thing happened. Almost all of the alpha particles went straight through the foil as if they'd hit nothing at all. This was what he expected to happen. If Thomson's model was accurate, there was nothing hard enough for these small particles to hit that would cause any change in their motion.

Every so often, though, one of the alpha particles would be deflected slightly as if it had bounced off of something hard. Even less often, Rutherford observed alpha particles bouncing straight back at the "gun" from which they had been fired! It was as if these alpha particles had hit a wall "head-on" and had ricocheted right back in the direction that they had come from.

Rutherford thought that these experimental results were rather odd. Rutherford described firing alpha particles at gold foil like shooting a high-powered rifle at tissue paper. Would you ever expect the bullets to hit the tissue paper and bounce back at you? Of course not! The bullets would break through the tissue paper and keep on going, almost as if they'd hit nothing at all. That's what Rutherford had expected would happen when he fired alpha particles at the gold foil. Therefore, the fact that most alpha particles passed through didn't shock him. On the other hand, how could he explain the alpha particles that got deflected? Furthermore, how could he explain the alpha particles that bounced right back as if they'd hit a wall?

Rutherford concluded that the only way to explain his results was to assume that the positive matter forming the gold atoms was not, in fact, distributed like the batter in plum pudding, but rather, was concentrated in one spot, forming a small positively charged particle somewhere in the center of the gold atom. We now call this clump of positively charged mass the nucleus. According to Rutherford, the presence of a nucleus explained his experiments, because it implied that most alpha particles passed through the gold foil without hitting anything at all. Once in a while, though, the alpha particles would actually collide with a gold nucleus, causing the alpha particles to be deflected, or even to bounce right back in the direction they came from.



Rutherford suggested that electrons surround a central nucleus.

(Image from http://upload.wikimedia.org/wikipedia/commons/7/7d/Rutherfordsches_Atommodell.png)

While Rutherford's discovery of the positively charged atomic nucleus offered insight into the structure of the atom, it also led to some questions. According to the "plum pudding" model, electrons were like plums embedded in the positive "batter" of the atom. Rutherford's model, though, suggested that the positive charge wasn't distributed like batter, but rather, was concentrated into a tiny particle at the center of the atom, while most of the rest of the atom was empty space. What did that mean for the electrons? If they weren't embedded in the positive material, exactly what were they doing? And how were they held in the atom? Rutherford suggested that the electrons might be circling or "orbiting" the positively charged nucleus as some type of negatively charged cloud, but at the time, there wasn't much evidence to suggest exactly how the electrons were held in the atom.

Despite the problems and questions associated with Rutherford's experiments, his work with alpha particles definitely seemed to point to the existence of an atomic "nucleus." Between J. J. Thomson, who discovered the electron, and Rutherford, who suggested that the positive charges in an atom were concentrated at the atom's center, the 1890s and early 1900s saw huge steps in understanding the atom at the "subatomic" (or smaller than the size of an atom) level. Although there was still some uncertainty with respect to exactly how subatomic particles were organized in the atom, it was becoming more and more obvious that atoms were indeed divisible. Moreover, it was clear that an atom contains negatively charged electrons and a nucleus containing positive charges. In the next section, we'll look more carefully at the structure of the nucleus, and we'll learn that while the atom is made up of positive and negative particles, it also contains neutral particles that neither Thomson, nor Rutherford, were able to detect with their experiments.

LESSON SUMMARY

- Dalton's Atomic Theory wasn't entirely correct. It turns out that atoms can be divided into smaller subatomic particles.
- According to Thomson's "plum pudding" model, the negatively charged electrons in an atom are like the pieces of fruit in a plum pudding, while the positively charged material is like the batter.
- When Ernest Rutherford fired alpha particles at a thin gold foil, most alpha particles went straight through; however, a few were scattered at different angles, and some even bounced straight back.
- In order to explain the results of his Gold Foil experiment, Rutherford suggested that the positive matter in the gold atoms was concentrated at the center of the gold atom in what we now call the nucleus of the atom.

VOCABULARY

- Subatomic particles: particles that are smaller than the atom
- Electron: a negatively charged subatomic particle
- Proton: a positively charged subatomic particle
- Nucleus: the small, dense center of the atom

2.2: REVIEW QUESTIONS

Decide whether each of the following statements is true or false.

1. Electrons (cathode rays) are positively charged.
2. Electrons (cathode rays) can be repelled by a negatively charged metal plate.
3. J.J. Thomson is credited with the discovery of the electron.
4. The plum pudding model is the currently accepted model of the atom

#5-11: Match each conclusion regarding subatomic particles and atoms with the observation/data that supports it.

Conclusion	Observations
All atoms have electrons	a. Most alpha particles shot at gold foil go straight through, without any change in their direction.
Atoms are mostly empty space.	b. A few alpha particles shot at gold foil bounce in the opposite direction.
Electrons have a negative charge	c. Some alpha particles (with positive charges) when shot through gold foil bend away from the gold.
The nucleus is positively charged	d. No matter which element Thomson put in a cathode ray tube, the same negative particles with the same properties (such as charge & mass) were ejected.
Atoms have a small, dense nucleus	e. The particles ejected in Thomson's experiment bent away from negatively charged plates, but toward positively charged plates.

12. What is the name given to the tiny clump of positive material at the center of an atom?

13. Electrons are ______ negatively charged metals plates and ______ positively charged metal plates.

FURTHER READING / SUPPLEMENTAL MATERIAL

A short history of the changes in our model of the atom, an image of the plum pudding model, and an animation of Rutherford's experiment can be viewed at Plum Pudding and Rutherford Page (http://www.newcastle-schools.org.uk/nsn/chemistry/Radioactivity/Plub%20Pudding%20and%20Rutherford%20Page.htm).

To see a video documenting the early history of the concept of the atom, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "history of the atom". Watch part 02. (you can get the username and password from your teacher)

Vision Learning: The Early Days (Thomson, etc)
http://visionlearning.com/library/module_viewer.php?mid=50&l=&c3=

Discovery of Electron (YouTube):
http://www.youtube.com/watch%3Fv%3DIdTxGJjA4Jw

Thomson's Experiment: http://www.aip.org/history/electron/jjthomson.htm

Discovery of Atomic Nucleus (YouTube):
http://www.youtube.com/watch%3Fv%3DwzALbzTdnc8

Rutherford's Experiment:
http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/ruther14.swf

PROTONS, NEUTRONS AND ELECTRONS IN ATOMS

Objectives

- Describe the locations, charges, and masses and the three main subatomic particles.
- Define atomic number.
- Describe the size of the nucleus in relation to the size of the atom.
- Define mass number.
- Explain what isotopes are and how isotopes affect an element's atomic mass.
- Determine the number of protons, neutrons, and electrons in an atom.

Introduction

Dalton's Atomic Theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it wasn't entirely accurate, because contrary to what Dalton believed, atoms can, in fact, be broken apart into smaller subunits or subatomic particles. We have been talking about the electron in great detail, but there are two other particles of interest to use: protons and neutrons. In this section, we'll look at the atom a little more closely.

Protons, Electrons, and Neutrons

We already learned that J.J. Thomson discovered a negatively charged particle, called the electron. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus which is called a proton. There is a third subatomic particle, known as a neutron. Ernest Rutherford proposed the existence of a neutral particle, with the approximate mass of a proton. Years later, James Chadwick proved that the nucleus of the atom contains this neutral particle that had been proposed by Ernest Rutherford. Chadwick observed that when beryllium is bombarded with alpha particles, it emits an unknown radiation that has approximately the same mass as a proton, but no electrical charge. Chadwick was able to prove that the beryllium emissions contained a neutral particle—Rutherford's neutron.

Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

As you might have already guessed from its name, the neutron is neutral. In other words, it has no charge whatsoever, and is therefore neither attracted to nor repelled from other objects. Neutrons are in every atom (with one exception), and they're bound together with other neutrons and protons in the atomic nucleus.

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to, nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).
Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Sub-Atomic Particles, Properties and Location			
Particle	**Relative Mass (amu)**	**Electric Charge**	**Location**
electron	$\frac{1}{1840}$	-1	outside the nucleus
proton	1	+1	nucleus
neutron	1	0	nucleus

The table shown gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in grams. The second column, however, shows the masses of the three subatomic particles in "atomic mass units". An atomic mass unit (amu) is defined as one-twelfth the mass of a carbon-12 atom. Atomic mass units (amu) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1.0 in this unit system.

It is difficult to find qualities that are different from each element and distinguish on element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons.

In addition to mass, another important property of subatomic particles is their charge. You already know that neutrons are neutral, and thus have no charge at all. Therefore, we say that neutrons have a charge of zero. What about electrons and protons? You know that electrons are negatively charged and protons are positively charged, but what's amazing is that the positive charge on a proton is exactly equal in magnitude (magnitude means "absolute value" or "size when you ignore positive and negative signs") to the negative charge on an electron. The third column in the table shows the charges of the three subatomic particles. Notice that the charge on the proton and the charge on the electron have the same magnitude.

Negative and positive charges of equal magnitude cancel each other out. This means that the negative charge on an electron perfectly balances the positive charge on the proton. In other words, a neutral atom must have exactly one electron for every proton. If a neutral atom has 1 proton, it must have 1 electron. If a neutral atom has 2 protons, it must have 2 electrons. If a neutral atom has 10 protons, it must have 10 electrons. You get the idea. In order to be neutral, an atom must have the same number of electrons and protons.

Atomic Number and Mass Number

Scientists can distinguish between different elements by counting the number of protons. If an atom has only one proton, we know it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom… the list goes on. Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. Therefore, scientists give this number a special name. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms. The periodic table gives the atomic number of each element. The atomic number is a whole number usually written above the chemical symbol of each element. The atomic number for hydrogen is 1, because every hydrogen atom has 1 proton. The atomic number for helium is 2 because every helium atom has 2 protons. What is the atomic number of carbon?

Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one electron as well. Gold, on the other hand, has an atomic number of 79, which means that an atom of gold has 79 protons, and, if it's neutral, and 79 electrons as well.

The mass number of an atom is the total number of protons and neutrons in its nucleus. Why do you think that the "mass number" includes protons and neutrons, but not electrons? You know that most of the mass of an atom is concentrated in its nucleus. The mass of an atom depends on the number of protons and neutrons. You have already learned that the mass of an electron is very, very small compared to the mass of either a proton or a neutron (like the mass of a penny compared to the mass of a bowling ball).

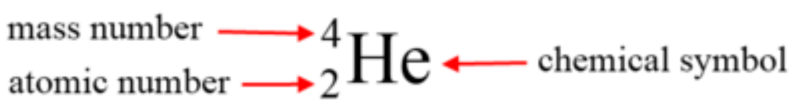
Counting the number of protons and neutrons tells scientists about the total mass of an atom.

mass number A = (number of protons) + (number of neutrons)

An atom's mass number is a very easy to calculate provided you know the number of protons and neutrons in an atom.

Example: What is the mass number of an atom of helium that contains 2 neutrons?
Solution: (number of protons) = 2 (Remember that an atom of helium always has 2 protons.) (number of neutrons) = 2 mass number = (number of protons) + (number of neutrons) mass number = 2 + 2 = 4

There are two main ways in which scientists frequently show the mass number of an atom they are interested in. It is important to note that the mass number is not given on the periodic table. These two ways include writing a nuclear symbol or by giving the name of the element with the mass number written.



To write a nuclear symbol, the mass number is placed at the upper left (superscript) of the chemical symbol and the atomic number is placed at the lower left (subscript) of the symbol. The complete nuclear symbol for helium-4 is drawn below.

The following nuclear symbols are for a nickel nucleus with 31 neutrons and a uranium nucleus with 146 neutrons.

$$^{59}_{28}\mathrm{Ni} \qquad ^{238}_{92}\mathrm{U}$$

In the nickel nucleus represented above, the atomic number 28 indicates the nucleus contains 28 protons, and therefore, it must contain 31neutrons in order to have a mass number of 59. The uranium nucleus has 92 protons as do all uranium nuclei and this particular uranium nucleus has 146 neutrons.

The other way of representing these nuclei would be Nickel-59 and Uranium-238, where 59 and 238 are the mass numbers of the two atoms, respectively. Note that the mass numbers (not the number of neutrons) is given to the side of the name.

Isotopes

Unlike the number of protons, which is always the same in atoms of the same element, the number of neutrons can be different, even in atoms of the same element. Atoms of the same element, containing the same number of protons, but different numbers of neutrons are known as isotopes. Since the isotopes of any given element all contain the same number of protons, they have the same atomic number (for example, the atomic number of helium is always 2). However, since the isotopes of a given element contain different numbers of neutrons, different isotopes have different mass numbers. The following two examples should help to clarify this point.

Example: a) What is the atomic number and the mass number of an isotope of lithium containing 3 neutrons. A lithium atom contains 3 protons in its nucleus. b) What is the atomic number and the mass number of an isotope of lithium containing 4 neutrons. A lithium atom contains 3 protons in its nucleus.
Solution: a) atomic number = (number of protons) = 3 (number of neutrons) = 3 mass number = (number of protons) + (number of neutrons) mass number = 3 + 3 = 6 b) atomic number = (number of protons) = 3 (number of neutrons) = 4 mass number = (number of protons) + (number of neutrons) mass number = 3 + 4 = 7

Notice that because the lithium atom always has 3 protons, the atomic number for lithium is always 3. The mass number, however, is 6 in the isotope with 3 neutrons, and 7 in the isotope with 4 neutrons. In nature, only certain isotopes exist. For instance, lithium exists as an isotope with 3 neutrons, and as an isotope with 4 neutrons, but it doesn't exists as an isotope with 2 neutrons, or as an isotope with 5 neutrons.

This whole discussion of isotopes brings us back to Dalton's Atomic Theory. According to Dalton, atoms of a given element are identical. But if atoms of a given element can have different numbers of neutrons, then they can have different masses as well! How did Dalton miss this? It turns out that elements found in nature exist as constant uniform mixtures of their naturally occurring isotopes. In other words, a piece of lithium always contains both types of naturally occurring lithium (the type with 3 neutrons and the type with 4 neutrons). Moreover, it always contains the two in the same relative amounts (or "relative abundances"). In a chunk of lithium, 93% will always be lithium with 4 neutrons, while the remaining 7% will always be lithium with 3 neutrons.

Dalton always experimented with large chunks of an element – chunks that contained all of the naturally occurring isotopes of that element. As a result, when he performed his measurements, he was actually observing the averaged properties of all the different isotopes in the sample. For most of our purposes in chemistry, we will do the same thing and deal with the average mass of the atoms. Luckily, aside from having different masses, most other properties of different isotopes are similar.

We can use what we know about atomic number and mass number to find the number of protons, neutrons, and electrons in any given atom or isotope. Consider the following examples:

Example: How many protons, electrons, and neutrons are in an atom of $^{40}_{19}K$?
Solution: Finding the number of protons is simple. The atomic number, # of protons, is listed in the bottom right corner. # protons = 19. For all atoms with no charge, the number of electrons is equal to the number of protons. # electrons =19. The mass number, 40, is the sum of the protons and the neutrons. To find the # of neutron, subtract the number of protons from the mass number. # neutrons = 40 – 19 = 21.

Example: How many protons, electrons, and neutrons in an atom of zinc-65?
Solution: Finding the number of protons is simple. The atomic number, # of protons, is found on the periodic table. All zinc atoms have # protons = 30. For all atoms with no charge, the number of electrons is equal to the number of protons. # electrons =30. The mass number, 65, is the sum of the protons and the neutrons. To find the # of neutron, subtract the number of protons from the mass number. # neutrons = 65 – 30 = 35.

LESSON SUMMARY

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they're neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- The mass number of an atom is the sum of the protons and neutrons in the atom
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

VOCABULARY

- Neutron: a subatomic particle with no charge
- Atomic mass unit (amu): a unit of mass equal to one-twelfth the mass of a carbon-twelve atom
- Atomic number: the number of protons in the nucleus of an atom
- Mass number: the total number of protons and neutrons in the nucleus of an atom
- Isotopes: atoms of the same element that have the same number of protons but different numbers of neutrons

Further Reading / Supplemental Material
Jeopardy Game: http://www.quia.com/cb/36842.html
For a Bill Nye video on atoms, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "Bill Nye atoms". (you can get the username and password from your teacher)

REVIEW QUESTIONS

1. Label each of the following statements as true or false.
   a) The nucleus of an atom contains all of the protons in the atom.
   b) The nucleus of an atom contains all of the electrons in the atom.
   c) Neutral atoms must contain the same number of neutrons as protons.
   d) Neutral atoms must contain the same number of electrons as protons.

2. Match the subatomic property with its description:

Sub-Atomic Particle	Characteristics
electron	a. has a charge of +1
neutron	b. has a mass of approximately 1/1840 amu
proton	c. is neither attracted to, nor repelled from charged objects

3. Indicate whether each statement is true or false.

   a) An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
   b) A neutral atom with 4 protons must have 4 electrons.
   c) An atom with 7 protons and 7 neutrons will have a mass number of 14.
   d) An atom with 7 protons and 7 neutrons will have an atomic number of 14.
   e) A neutral atom with 7 electrons and 7 neutrons will have an atomic number of 14.

4. Use the periodic table to find the symbol for the element with:
   - 44 electrons in a neutral atom
   - 30 protons
   - An atomic number of 36

5. In the table below, Column 1 contains data for 5 different elements. Column 2 contains data for the same 5 elements, however different isotopes of those elements. Match the atom in the first column to its isotope in the second column.

Original element	Isotope of the same element
an atom with 2 protons and 1 neutron	a. a C (carbon) atom with 6 neutrons
a Be (beryllium) atom with 5 neutrons	b. an atom with 2 protons and 2 neutrons
an atom with an atomic number of 6 and mass number of 13	c. an atom with an atomic number of 7 and a mass number of 15
an atom with 1 proton and a mass number of 1	d. an atom with an atomic number of 1 and 1 neutron
an atom with an atomic number of 7 and 7 neutrons	e. an atom with an atomic number of 4 and 6 neutrons

6. Write the nuclear symbol for each element described:
   a) 32 neutrons in an atom with mass number of 58

   b) An atom with 10 neutrons and 9 protons.

Indicate the number of protons, neutrons, and electrons in each of the following atoms:

$^{4}_{2}He$

Sodium-23

$^{1}_{1}H$

Iron-55

$^{37}_{17}Cl$

Boron-11

$^{238}_{92}U$

Uranium-235

THE MOLE

Objectives

- **Use Avogadro's number to convert to moles and vice versa given the number of particles of a substance.**
- **Use the molar mass to convert to grams and vice versa given the number of moles of a substance.**

Introduction

When objects are very small, it is often inconvenient or inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross."

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number

In chemistry, it is impossible to deal with a single atom or molecule because we can't see them or count them or weigh them. Chemists have selected a number of particles with which to work that is convenient. Since molecules are extremely small, you may suspect that this number is going to be very large and you are right. The number of particles in this group is 6.02 x 1023 particles and the name of this group is the mole (the abbreviation for mole is mol). One mole of any object is 6.02 x 1023 of those objects. There is a very particular reason that this number was chosen and we hope to make that reason clear to you.

When chemists are carrying out chemical reactions, it is important that the relationship between the numbers of particles of each reactant is known. Chemists looked at the atomic masses on the periodic table and understood that the mass ratio of one carbon atom to one sulfur atom was 12 amu to 32 amu. They realized that if they massed out 12 grams of carbon and 32 grams of sulfur, they would have the same number of atoms of each element. They didn't know how many atoms were in each pile but they knew the number in each pile had to be the same. This is the same logic as knowing that if a basketball has twice the mass of a soccer ball and you massed out 100 lbs of basketballs and 50 lbs of soccer balls, you would have the same number of each ball. Many years later, when it became possible to count particles using electrochemical reactions, the number of atoms turned out to be 6.02x1023 particles. Eventually chemists decided to call that number of particles a mole.

The number 6.02 x 1023 is called Avogadro's number. Avogadro, of course, had no hand in determining this number, rather it was named in honor of Avogadro.

Converting Between Molecules to Moles

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given a number of molecules of a substance, we can convert it into moles by dividing by Avogadro's number and vice versa.

Example: How many moles are present in 1 billion (1x109) molecules of water?
Solution: $1{\times}10^9 molecules\ H_2O \cdot \frac{1\ mol\ H_2O}{6.02{\times}10^{23} molecules\ H_2O} = \mathbf{1.7{\times}10^{-15} mol\ H_2O}$

You should note that this amount of water is too small for even our most delicate balances to determine the mass. A very large number of molecules must be present before the mass is large enough to detect with our balances.

Example: How many molecules are present in 0.00100 mol?
Solution: $0.00100\ mol \cdot \frac{6.02{\times}10^{23} molecules}{1\ mol} = \mathbf{6.02{\times}10^{20}\ molecules}$

Converting Grams to Moles and Vice Versa

1.00 mol of carbon-12 atoms has a mass of 12.0 g and contains 6.02x1023 atoms. Likewise, 1.00 mol of water has a mass of 18.0 grams and contains 6.02x1023 molecules. 1.00 mole of any element or compound has a mass equal to its molecular mass in grams and contains 6.02x1023 particles. The mass, in grams, of 1 mole of particles of a substance is now called the molar mass (mass of 1.00 mole).

To quickly find the molar mass of a substance, you need to look up the masses on the periodic table and add them together. For example, water has the formula H2O. Hydrogen has a mass of 1.0084 g/mol (see periodic table) and oxygen has a mass of 15.9994 g/mol. The molar mass of H2O=2(1.0084g/mol) + 15.9994g/mol = 18.0162g/mol. This means that 1 mole of water has a mass of 18.0162 grams.

We can also convert back and forth between grams of substance and moles. The conversion factor for this is the molar mass of the substance. The molar mass is the ratio giving the number of grams for each one mole of a substance. This ratio is easily found by adding up the atomic masses of the elements within a compound using the periodic table. This ratio has units of grams per mole or g/mol.

Example: Find the molar mass of each of the following: a) S b) H_2O c) F_2 d) H_2SO_4 e) $Al_2(SO_4)_3$
Solution: You will need a periodic table to solve these problems. Look for each element's mass. a) Look for sulfur on the periodic table. Its molar mass is 32.065 g/mol. That means that one mole of sulfur has a mass of 32.065 grams. b) This compound contains two hydrogen atoms and one oxygen atoms. To find the molar mass of H_2O, we need to add the mass of two hydrogen atoms plus the mass of one oxygen atom. We get: 2(1.008) + 16.00 = 18.016 g/mol. That means that one mole of water has a mass of just over 18 grams. c) This compound contains two fluorine atoms. To find the molar mass of F_2, we need to add the mass of two fluorine atoms. We get: 2(19.00) = 38.00 g/mol d) This compound contains two hydrogen atoms, one sulfur atom, and four oxygen atoms. To find the molar mass of H_2SO_4, we need to add the mass of two hydrogen atoms plus the mass of one sulfur atom plus the mass of four oxygen atoms. We get: 2(1.008) + 32.065 + 4(16.00) = 100.097 g/mol e) This compound contains two aluminum atoms, three sulfur atoms, and twelve oxygen atoms. To find the molar mass of $Al_2(SO_4)_3$, we need to add the mass of all of these atoms. We get: 2(26.98) + 3(32.065) + 12(16.00) = 342.155 g/mol

To convert the grams of a substance into moles, we use the ratio molar mass. We divide by the molar mass and to convert the moles of a substance into grams, we multiply by the molar mass.

Example: How many moles are present in 108 grams of water?
Solution: $108\ g\ H_2O \cdot \frac{1\ mol\ H_2O}{18.02\ g\ H_2O} = 5.99\ mol\ H_2O$ To get the ratio 1 mol H_2O=18.02 g, we added up the molar mass of H_2O using the masses on a periodic table.

Example: What is the mass of 7.50 mol of CaO?
Solution: $7.50\ mol\ CaO \cdot \frac{56.0\ g\ CaO}{1\ mol\ CaO} = 420.\ g\ CaO$ To get the ratio 1 mol CaO=56.0 g, we added up the molar mass of CaO using the masses on a periodic table.

We will be using these ratios again to solve more complex problems in the next chapters. Being able to use these ratios is a very important skill for later math problems.

LESSON SUMMARY

There are 6.02x1023 particles in 1.00 mole. This number is called Avogadro's number. The molar mass of a substance can be found by adding up the masses on a periodic table. Using the factor-label method, it is possible to convert between grams, moles, and the number of atoms or molecules.

VOCABULARY

Avogadro's number: The number of objects in a mole; equal to 6.02x1023.
Mole: An Avogadro's number of objects.
Molar Mass: The mass, in grams, of 1 mole of a substance. This can be found by adding up the masses on the periodic table.

Further Reading / Supplemental Links
http://learner.org/resources/series61.html The learner.org website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called The Mole.
Using Avogadro's law, the mass of a substance can be related to the number of particles contained in that mass. The Mole:
(http://www.learner.org/vod/vod_window.html?pid=803)
Vision Learning tutorial: The Mole
http://visionlearning.com/library/module_viewer.php?mid=53&l=&c3=

REVIEW QUESTIONS

1. **How many molecules are present in the following quantities?**

0.250 mol H2O	0.0045 mol Al2(CO3)3

2. **How many moles are present in the following quantities?**

1.0x1020 molecules H2O	5 billion atoms of carbon

3. **What is the molar mass of each of the following substances? Include units with your answer.**

H2O	NaOH	NH4Cl
H2SO4	Al2(CO3)3	PbO2

4. Convert the following to moles.

60.0 g NaOH	5.70 g H2SO4
2.73 g NH4Cl	10.0g PbO2

5. Convert the following to grams.

0.100 mol CO2	0.500 mol (NH4)2CO3
0.437 mol NaOH	3.00 mol H2O

6. How many molecules are present in the following masses?

1.00 g Na2CO3	1000. g H2O

7. Convert the following to grams.

2.0x1023 molecules H2	1.75x1024 molecules NaCl
8.6x1022 molecules NaOH	

STANDARD 1, OBJECTIVE 3: CORRELATE ATOMIC STRUCTURE AND THE PHYSICAL AND CHEMICAL PROPERTIES OF AN ELEMENT TO THE POSITION OF THE ELEMENT ON THE PERIODIC TABLE.

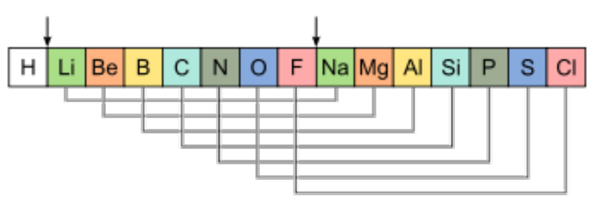
MENDELEEV'S PERIODIC TABLE

Objectives

- Describe the method Mendeleev used to make his periodic table.
- List the advantages and disadvantages Mendeleev's table had over other methods of organizing the elements.
- Explain how our current periodic table differs from Mendeleev's original table.

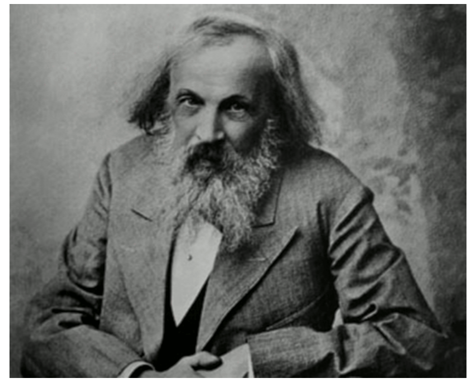
Introduction

During the 1800s, when most of the elements were being discovered, many chemists tried to classify the elements according to their similarities. In 1829, Johann Döbereiner noted chemical similarities in several groups of three elements and placed these elements into what he called triads. His groupings included the triads of 1) chlorine, bromine, and iodine, 2) sulfur, selenium, and tellurium, 3) calcium, strontium, and barium, and 4) lithium, sodium, and potassium. In all of the triads, the atomic weight of the second element was almost exactly the average of the atomic weights of the first and third element.



John Newlands' law of octaves suggested that, if elements are aligned in order of increasing mass, every eighth element would have similar properties.

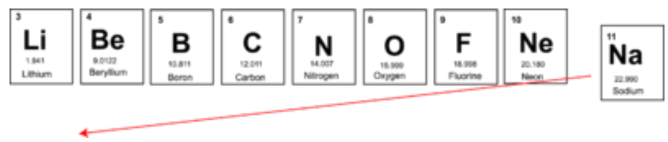
In 1864, John Newlands saw a connection between the chemical properties of elements and their atomic masses. He stated that if the known elements, beginning with lithium are arranged in order of increasing mass, the eighth element will have properties similar to the first, the ninth similar to the second, the tenth similar to the third, and so on. Newlands called his relationship the law of octaves, comparing the elements to the notes in a musical scale. Newlands tried to force all the known elements to fit into his octaves but many of the heavier elements, when discovered, did not fit into his patterns.



Dmitri Mendeleev created the first periodic table in 1869.

Mendeleev Organized His Table According to Chemical Behavior

By 1869, a total of 63 elements had been discovered. As the number of known elements grew, scientists began to recognize patterns in the way chemicals reacted and began to devise ways to classify the elements. Dmitri Mendeleev, a Siberian-born Russian chemist, was the first scientist to make a periodic table much like the one we use today.



Mendeleev's table listed the elements in order of increasing atomic mass. Then he placed elements underneath other elements with similar chemical behavior. For example, lithium is a shiny metal, soft enough to be cut with a spoon. It reacts readily with oxygen and reacts violently with water.

When it reacts with water, it produces hydrogen gas and lithium hydroxide. As we proceed through the elements with increasing mass, we will come to the element sodium. Sodium is a shiny metal, soft enough to be cut with a spoon. It reacts readily with oxygen and reacts violently with water. When it reacts with water, it produces hydrogen gas and sodium hydroxide. You should note that the description of the chemical behavior of sodium is very similar to the chemical description of lithium. When Mendeleev found an element whose chemistry was very similar to a previous element, he placed it below the similar element.

Changes to our Modern Periodic Table

The periodic table we use today is similar to the one developed by Mendeleev, but is not exactly the same. There are some important distinctions: Mendeleev's table did not include any of the noble gases, which were discovered later. These were added by Sir William Ramsay as Group 0, without any disturbance to the basic concept of the periodic table. (These elements were later moved to form group 18 or 8A.) Other elements were also discovered and put into their places on the periodic table.

As previously noted, Mendeleev organized elements in order of increasing atomic mass, with some problems in the order of masses. In 1914 Henry Moseley found a relationship between an element's X-ray wavelength and its atomic number, and therefore organized the table by nuclear charge (or atomic number) rather than atomic weight. Thus Moseley placed argon (atomic number 18) before potassium (atomic number 19) based on their X-ray wavelengths, despite the fact that argon has a greater atomic weight (39.9) than potassium (39.1). The new order agrees with the chemical properties of these elements, since argon is a noble gas and potassium an alkali metal. Similarly, Moseley placed cobalt before nickel, and was able to explain that tellurium should be placed before iodine, not because of an error in measuring the mass of the elements (as Mendeleev suggested), but because tellurium had a lower atomic number than iodine.

Moseley's research also showed that there were gaps in his table at atomic numbers 43 and 61 which are now known to be technetium and promethium, respectively, both radioactive and not naturally occurring. Following in the footsteps of Dmitri Mendeleev, Henry Moseley also predicted new elements.

You already saw that the elements in vertical columns are related to each other by their electron configuration, but remember that Mendeleev did not know anything about electron configuration. He placed the elements in their positions according to their chemical behavior. Thus, the vertical columns in Mendeleev's table were composed of elements with similar chemistry. These vertical columns are called groups or families of elements.

LESSON SUMMARY

The periodic table in its present form was organized by Dmitri Mendeleev.
Mendeleev organized the elements in order of increasing atomic mass and in groups of similar chemical behavior. He also left holes for missing elements and used the patterns of his table to make predictions of properties of these undiscovered elements.
The modern periodic table now arranges elements in order of increasing atomic number. Additionally, more groups and elements have been added as they have been discovered.

VOCABULARY

- Periodic table: a tabular arrangement of the chemical elements according to atomic number.
- Mendeleev: the Russian chemist credited with organizing the periodic table in the form we use today.
- Moseley: the chemist credited with finding that each element has a unique atomic number

Further Reading / Supplemental Links

Tutorial: Vision Learning: The Periodic Table
http://visionlearning.com/library/module_viewer.php?mid=52&l=&c3=

How the Periodic Table Was Organized (YouTube):
http://www.youtube.com/watch%3Fv%3DCdkpoQk2LDE

For several videos and video clips describing the periodic table, go to http://www.uen.org/dms/. Go to the k-12 library. Search for “periodic table”. (you can get the username and password from your teacher)

Review Questions

1. What general organization did Mendeleev use when he constructed his table?
2. How did Mendeleev’s system differ from Newlands’s system?
3. Did all elements discovered at the time of Mendeleev fit into this organization system? How would the discovery of new elements have affecting Mendeleev’s arrangement of the elements?
4. Look at Mendeleev’s predictions for Germanium (ekasilicon). How was Mendeleev able to make such accurate predictions?
5. What problems did Mendeleev have when arranging the elements according to his criteria? What did he do to fix his problems?
6. What discovery did Henry Moseley make that changed how we currently recognize the order of the elements on the periodic table?
7. List three ways in which our current periodic table differs from the one originally made by Mendeleev.

METALS, NONMETALS, AND METALLOIDS

Objectives

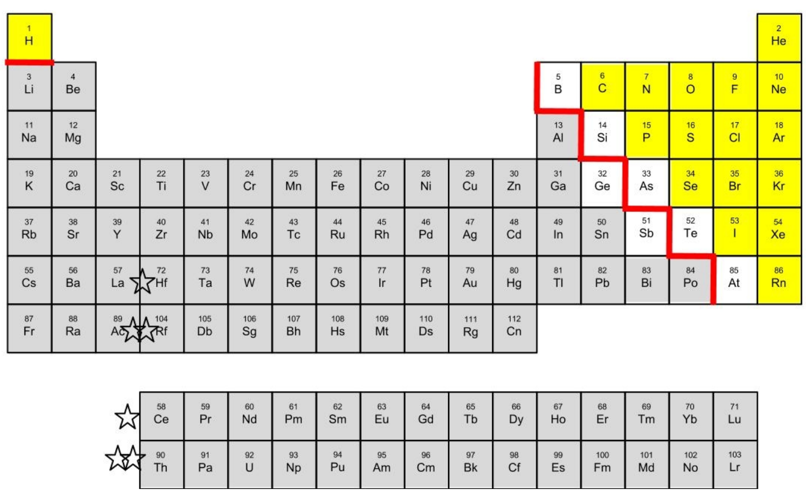
- Describe the differences among metals, nonmetals, and metalloids.
- Identify an element as a metal, nonmetal, or metalloid given a periodic table or its properties.

Introduction

In the periodic table, the elements are arranged according to similarities in their properties. The elements are listed in order of increasing atomic number as you read from left to right across a period and from top to bottom down a group. In this section you will learn the general behavior and trends within the periodic table that result from this arrangement in order to predict the properties of the elements.

Metals, Non-metals, and Metalloids

There is a progression from metals to non-metals across each row of elements in the periodic table. The diagonal line at the right side of the table separates the elements into two groups: the metals and the non-metals. The elements that are on the left of this line tend to be metals, while those to the right tend to be non-metals (with the exception of hydrogen which is a nonmetal). The elements that are directly on the diagonal line are metalloids, with some exceptions. Aluminum touches the line, but is considered a metal. Metallic character generally increases from top to bottom down a group and right to left across a period, meaning that francium (Fr) has the most metallic character of all of the discovered elements.



The division of the periodic table into metals and non-metals. The metalloids are most of the elements along the line drawn. Additionally, the element hydrogen is a NONMETAL, even though it is on the left side of the periodic table.

Image by Tracy Poulsen

Most of the chemical elements are metals. Most metals have the common properties of being shiny, very dense, and having high melting points. Metals tend to be ductile (can be drawn out into thin wires) and malleable (can be hammered into thin sheets). Metals are good conductors of heat and electricity. All metals are solids at room temperature except for mercury. In chemical reactions, metals easily lose electrons to form positive ions. Examples of metals are silver, gold, and zinc.

Nonmetals are generally brittle, dull, have low melting points, and they are generally poor conductors of head heat and electricity. In chemical reactions, they ted to gain electrons to form negative ions. Examples of non-metals are hydrogen, carbon, and nitrogen. Metalloids have properties of both metals and nonmetals. Metalloids can be shiny or dull. Electricity and heat can travel through metalloids, although not as easily as they can through metals. They are also called semimetals. They are typically semi-conductors, which means that they are elements that conduct electricity better than insulators, but not as well as conductors. They are valuable in the computer chip industry. Examples of metalloids are silicon and boron.

LESSON SUMMARY

- There is a progression from metals to non-metals across each period of elements in the periodic table.
- Metallic character generally increases from top to bottom down a group and right to left across a period.

VOCABULARY

- periodic law: states that the properties of the elements recur periodically as their atomic numbers increase
- ductile: can be drawn out into thin wires
- malleable: can be hammered into thin sheets

REVIEW QUESTIONS

1. Label each of the following elements as a metal, nonmetal, or metalloid.
    a. Carbon
    b. Bromine
    c. Oxygen
    d. Plutonium
    e. Potassium
    f. Helium

2. Given each of the following properties, label the property of as that of a metal, nonmetal, or metalloid.
    a. Lustrous
    b. Semiconductors
    c. Brittle
    d. Malleable
    e. Insulators
    f. Conductors

3. The elements mercury and bromine are both liquids at room temperature, but mercury is considered a metal and bromine is considered a nonmetal. How can that be? What properties do metals and nonmetals have?

Families and Periods of the Periodic Table

Objectives

- **Give the name and location of specific groups on the periodic table, including alkali metals, alkaline earth metals, noble gases, halogens, and transition metals.**
- **Explain the relationship between the chemical behavior of families in the periodic table and their electron configuration.**
- **Identify elements that will have the most similar properties to a given element.**

Introduction

Since the families of elements were organized by their chemical behavior, it is predictable that the individual members of each chemical family will have similar electron configurations.

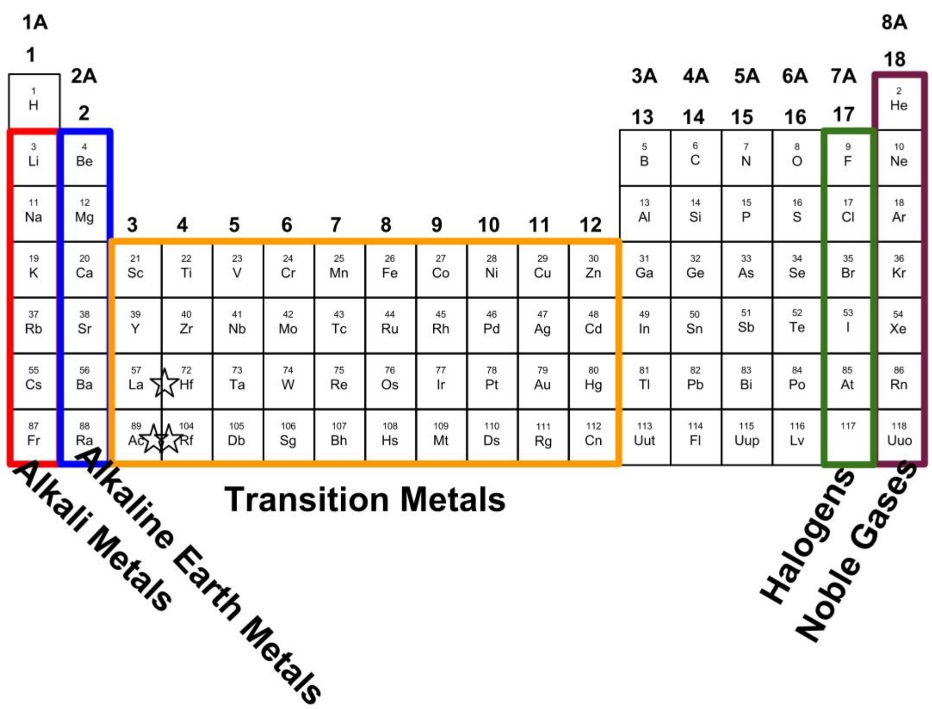
Families of the Periodic Table

Remember that Mendeleev arranged the periodic table so that elements with the most similar properties were placed in the same group. A group is a vertical column of the periodic table. All of the 1A elements have one valence electron. This is what causes these elements to react in the same ways as the other members of the family. The elements in 1A are all very reactive and form compounds in the same ratios with similar properties with other elements. Because of their similarities in their chemical properties, Mendeleev put these elements into the same group. Group 1A is also known as the alkali metals. Although most metals tend to be very hard, these metals are actually soft and can be easily cut.

Group 2A is also called the alkaline earth metals. Once again, because of their similarities in electron configurations, these elements have similar properties to each other. The same pattern is true of other groups on the periodic table. Remember, Mendeleev arranged the table so that elements with the most similar properties were in the same group on the periodic table.

It is important to recognize a couple of other important groups on the periodic table by their group name. Group 7A (or 17) elements are also called halogens. This group contains very reactive nonmetals elements.
The noble gases are in group 8A. These elements also have similar properties to each other, the most significant property being that they are extremely unreactive rarely forming compounds. We will learn the reason for this later, when we discuss how compounds form. The elements in this group are also gases at room temperature.

An alternate numbering system numbers all of the groups from 1-18. In this numbering system, group 1A is group 1; group 2A is group 2; the halogens (7A) are group 17; and the noble gases (8A) are group 18. You will come across periodic table with both numbering systems. It is important to recognize which numbering system is being used and to be able to find the number of valence electrons in the main block elements regardless of which numbering systems is being used.



Families of the periodic table. Image by Tracy Poulsen

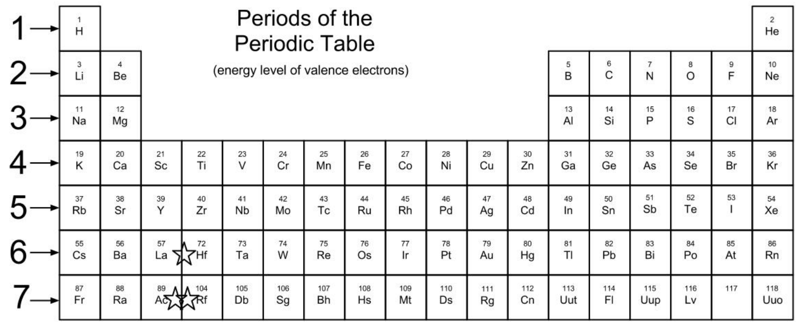


Image by Tracy Poulsen

Periods of the Periodic Table

If you can locate an element on the Periodic Table, you can use the element's position to figure out the energy level of the element's valence electrons. A period is a horizontal row of elements on the periodic table. The period also happens to be the energy level of the element's valence electrons. For example, the elements sodium (Na) and magnesium (Mg) are both in period 3. The elements astatine (At) and radon (Rn) are both in period 6.

LESSON SUMMARY

- The vertical columns on the periodic table are called groups or families because of their similar chemical behavior.
- All the members of a family of elements have the same number of valence electrons and similar chemical properties
- The horizontal rows on the periodic table are called periods.

VOCABULARY

- Group (family): a vertical column in the periodic table
- Alkali metals: group 1A of the periodic table
- Alkali earth metals: group 2A of the periodic table
- Halogens: group 7A of the periodic table
- Noble gases: group 8A of the periodic table
- Transition elements: groups 3 to 12 of the periodic table

Further Reading / Supplemental Links
http://www.wou.edu/las/physci/ch412/perhist.htm
http://www.aip.org/history/curie/periodic.htm
http://web.buddyproject.org/web017/web017/history.html
http://www.dayah.com/periodic
http://www.chemtutor.com/perich.htm

REVIEW QUESTIONS

Multiple Choice

1. Which of the following elements is in the same family as fluorine?
   a) silicon
   b) antimony
   c) iodine
   d) arsenic
   e) None of these.

2. Elements in a ______________ have similar chemical properties.
   a) period
   b) family
   c) both A and B
   d) neither A nor B

3. Which of the following elements would you expect to be most similar to carbon?
   a) Nitrogen
   b) Boron
   c) Silicon

4. Give the name of the family in which each of the following elements is located:
    a) astatine
    b) krypton
    c) barium
    d) francium

5. Which family is characterized by each of the following descriptions?
    a) A very reactive family of nonmetals
    b) Have 7 valence electrons
    c) A nonreactive family of nonmetals
    d) Forms colorful compounds
    e) Have 2 valence electrons
    f) A very reactive family of metals

ATOMIC ENERGY

CHAPTER 2

Standard 2: Students will understand the relationship between energy changes in the atom specific to the movement of electrons between energy levels in an atom resulting in the emission or absorption of quantum energy. They will also understand that the emission of high-energy particles results from nuclear changes and that matter can be converted to energy during nuclear reactions.

STANDARD 2, OBJECTIVE 1: EVALUATE QUANTUM ENERGY CHANGES IN THE ATOM IN TERMS OF THE ENERGY CONTAINED IN LIGHT EMISSIONS.

ELECTROMAGNETIC RADIATION AND WAVES

Introduction

When you walk outside on a sunny day, the only kinds of radiation you can detect are visible light, which you can detect with your eyes, and infrared light, which you feel as warmth on your skin.

Sunlight consists of all the different kinds of electromagnetic radiation, from harmless radio waves to deadly gamma rays.
Fortunately, Earth's atmosphere prevents most of the harmful radiation from reaching Earth's surface. You can read about the different kinds of electromagnetic radiation in this section.

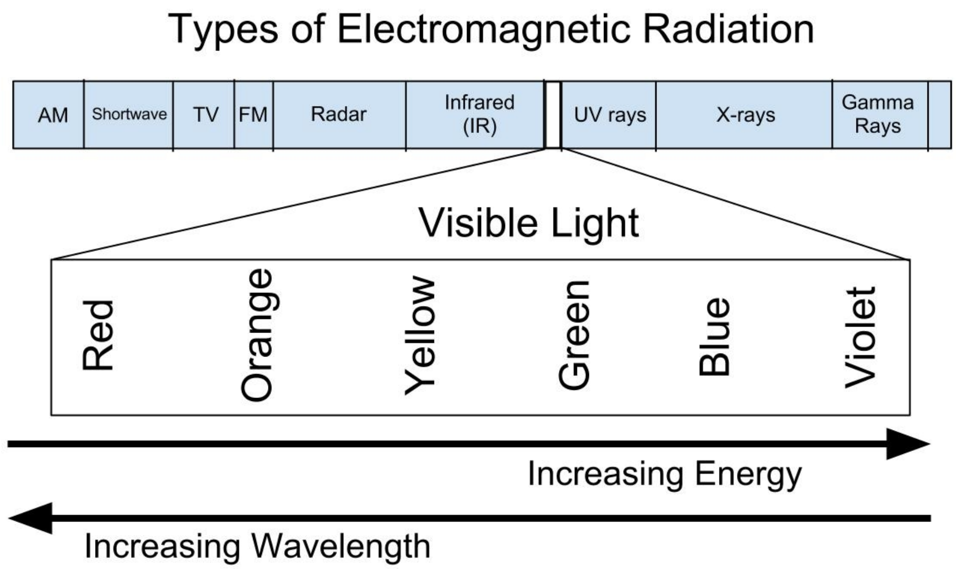
Electromagnetic Radiation

Electromagnetic radiation is energy that travels in waves across space as well as through matter. Most of the electromagnetic radiation on Earth comes from the sun. Like other waves, electromagnetic waves are characterized by certain wavelengths and wave frequencies. **Wavelength** is the distance between two corresponding points on adjacent waves. Wave **frequency** is the number of waves that pass a fixed point in a given amount of time. Electromagnetic waves with shorter wavelengths have higher frequencies and more energy.

A Spectrum of Electromagnetic Waves

Visible light and infrared light are just a small part of the full range of electromagnetic radiation, which is called the electromagnetic spectrum. You can see the waves of the electromagnetic spectrum in the Figure below. At the top of the diagram, the wavelengths of the waves are given. Also included are objects that are about the same size as the corresponding wavelengths. The frequencies and energy levels of the waves are shown at the bottom of the diagram. Some sources of the waves are also given.

For a video introduction to the electromagnetic spectrum, go to this URL:
http://www.youtube.com/watch?NR=1&feature=endscreen&v=cfXzwh3KadE



- On the left side of the electromagnetic spectrum diagram are radio waves and microwaves. Radio waves have the longest wavelengths and lowest frequencies of all electromagnetic waves. They also have the least amount of energy.
- On the right side of the diagram are X rays and gamma rays. They have the shortest wavelengths and highest frequencies of all electromagnetic waves. They also have the most energy.
- Between these two extremes are waves that are commonly called light. Light includes infrared light, visible light, and ultraviolet light. The wavelengths, frequencies, and energy levels of light fall in between those of radio waves on the left and X rays and gamma rays on the right.
- (visible light spectrum & Energy = ROYGBIV)

Summary

- Electromagnetic radiation travels in waves through space or matter. Electromagnetic waves with shorter wavelengths have higher frequencies and more energy.
- The full range of electromagnetic radiation is called the electromagnetic spectrum. From longest to shortest wavelengths, it includes radio waves, microwaves, infrared light, visible light, ultraviolet light, X rays, and gamma rays.

Practice

At the first URL below, read about electromagnetic waves with different frequencies. Then use the information to complete the table at the second URL.
http://www.darvill.clara.net/emag/index.htm and http://www.darvill.clara.net/emag/gcseemag.pdf

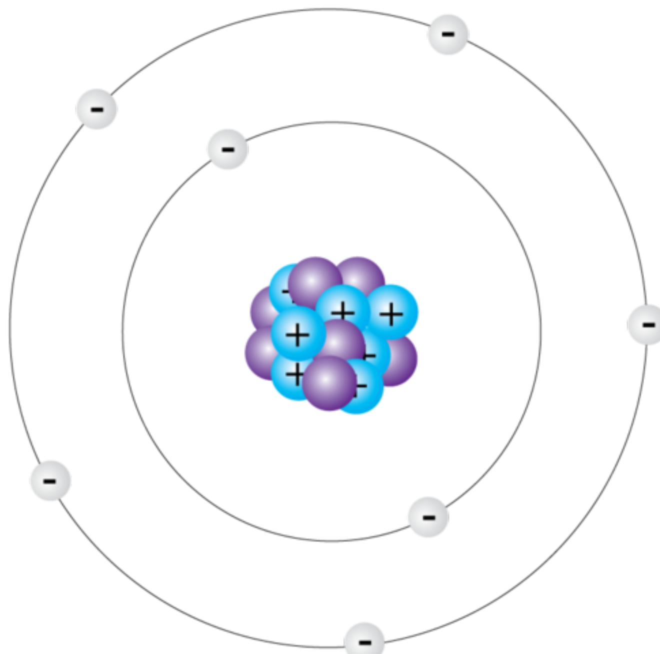
Review

1. Describe the relationship between the wavelength and frequency of electromagnetic waves.
2. What is the electromagnetic spectrum?
3. Which electromagnetic waves have the longest wavelengths?
4. Which type of light has the highest frequencies?
5. Explain why gamma rays are the most dangerous of all electromagnetic waves.

ELECTRONS

Where Are the Electrons?

To help us understand the way in which electrons move energy levels we will look at Bohr's model of the atom. In this model, negative electrons circle the positive nucleus at fixed distances from the nucleus, called **energy levels**. While more modern and accurate models of the atom demonstrate that electrons do not orbit the nucleus in fixed paths, they actually occupy regions of space called orbitals. You can see the model in Figure below for an atom of the element nitrogen using Bohr's model. Bohr's model is useful for understanding properties of elements and their chemical interactions.



Electrons absorb energy and get pushed up energy levels. When they drop energy levels they emit energy in the form of light. The visible light spectrum contains violet light, indigo, blue, green, yellow, orange, and red. The violet light has the highest amount of energy, or frequency, and the lowest wavelength. Red light has the lowest amount of energy or frequency, and has the longest wavelength.

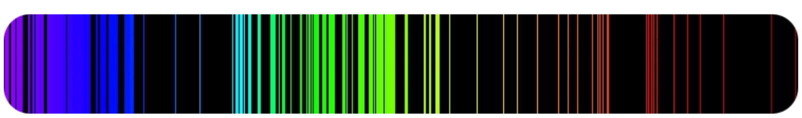
The light emitted by the sign containing neon gas (on the left) is different from the light emitted by the sign containing argon gas (on the right).

Each Element Has a Unique Spectrum

The light frequencies emitted by atoms are mixed together by our eyes so that we see a blended color. Several physicists, including Anders J. Angstrom in 1868 and Johann J. Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out so they could see the individual frequencies that made up the light.

In the figure above, we see the **emission spectrum**, or atomic spectrum, for hydrogen gas. The emission spectrum of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity.

When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the light is separated into individual colors, we see that the hydrogen spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors. Every atom has its own characteristic spectrum; no two atomic spectra are alike. The image below shows the emission spectrum of iron. Because each element has a unique emission spectrum, elements can be identified using them.



You may have heard or read about scientists discussing what elements are present in the sun or some more distant star and wondered how scientists could know what elements are present in a place no one has ever been. Scientists determine what elements are present in distant stars by analyzing the light that comes from those stars and using the atomic spectrum to identify the elements emitting that light.

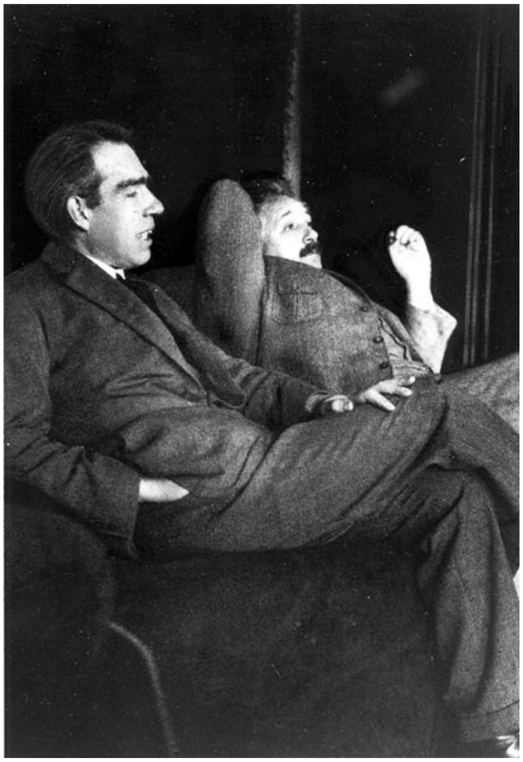
In the same way that we can identify elements by their line emission spectrum, elements can also be identified using a flame test. In a flame test an element is burned in a flame and the results flame color is indicative and unique to the element.

The results of some flame tests are as follows:

- Lithium burns red
- Copper burns green/blue
- Magnesium burns white
- Strontium burns red
- Barium burns yellow/green
- Potassium burns light purple
- Sodium burns yellow/orange.

Go ahead and draw the colors of those elements below:

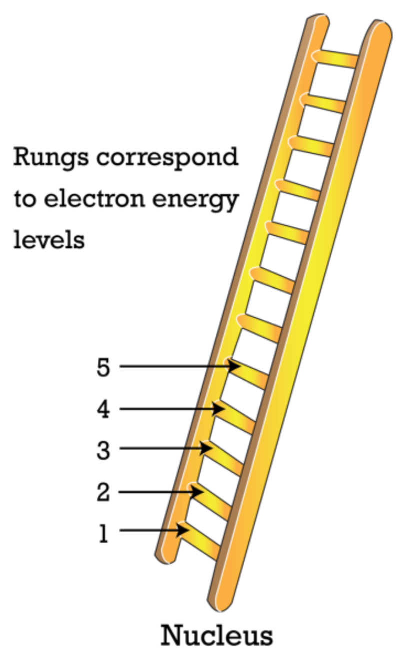
Energy Levels



Niels Bohr and Albert Einstein in 1925. Bohr received the Nobel prize for physics in 1922.

The key idea in Bohr's model of the atom is that electrons occupy definite orbits that require the electron to have a specific amount of energy. In order for an electron to be in the electron cloud of an atom, it must be in one of the allowable orbits and it must have the precise energy required for that orbit. Orbits closer to the nucleus would require smaller amounts of energy for an electron and orbits farther from the nucleus would require the electrons to have a greater amount of energy. The possible orbits are known as **energy levels**. One of the weaknesses of Bohr's model was that he could not offer a reason why only certain energy levels or orbits were allowed.

Bohr hypothesized that the only way electrons could gain or lose energy would be to move from one energy level to another, thus gaining or losing precise amounts of energy. The energy levels are **quantized**, meaning that only specific amounts are possible. It would be like a ladder that had rungs only at certain heights. The only way you can be on that ladder is to be on one of the rungs and the only way you could move up or down would be to move to one of the other rungs. Suppose we had such a ladder with 10 rungs. Other rules for the ladder are that only one person can be on a rung and in normal state, the ladder occupants must be on the lowest rung available. If the ladder had five people on it, they would be on the lowest five rungs. In this situation, no person could move down because all the lower rungs are full. Bohr worked out rules for the maximum number of electrons that could be in each energy level in his model and required that an



atom is in its normal state (ground state) had all electrons in the lowest energy levels available. Under these circumstances, no electron could lose energy because no electron could move down to a lower energy level. In this way, Bohr's model explained why electrons circling the nucleus did not emit energy and spiral into the nucleus.

Bohr's Model and Atomic Spectra

The evidence used to support Bohr's model came from the atomic spectra. He suggested that an atomic spectrum is created when the electrons in an atom move between energy levels. The electrons typically have the lowest energy possible, called **ground state**. If the electrons are given energy (through heat, electricity, light, etc.), the electrons in an atom could absorb energy by jumping to a higher energy level or an **excited state**. When the electrons fall back to a lower energy level, the electrons then give off the absorbed energy in the form of a light particle called a **photon**. The energy emitted by electrons dropping back to lower energy levels would always be in precise amounts of energy because the differences in energy levels are precise. This explains why you see specific lines of light when looking at an atomic spectrum - each line of light matches a specific "step down" that an electron can take in that atom.

LESSON SUMMARY

- Bohr's model suggests each atom has a set of unchangeable energy levels and electrons in the electron cloud of that atom must be in one of those energy levels.
- Bohr's model suggests that the atomic spectra of atoms is produced by electrons gaining energy from some source, jumping up to a higher energy level, then dropping back to a lower energy level and emitting the energy difference between the two energy levels.

Standard 2. Objective 2: Evaluate how changes in the nucleus of an atom result in emission of radioactivity.

Radiation and Radioactive Particles

Lesson objectives

- What is a radioactive particle
- Describe what half life is how it is used.
- Compare the mass, energy and penetrating power of alpha, beta and gamma radiation.

Elements that have a larger number of neutrons than protons in the nucleus of an atom tend to break down and release energy. Oxygen isotopes all have 8 protons, but Oxygen 16 has 8 neutrons and Oxygen-17 has 9 neutrons. Oxygen-17 will be unstable, and release radiation. Carbon-12 and Carbon-14 are isotopes. Carbon-14 has more neutrons than Carbon-12 so is unstable. Radiation is the release of energy and particles from the breakdown or decay of an unstable nucleus.

Radioactive decay is the process in which the nuclei of radioactive atoms emit charged particles and energy, which are called by the general term radiation. Radioactive atoms have unstable nuclei, and when the nuclei emit radiation, they become more stable. Radioactive decay is a nuclear—rather than chemical—reaction because it involves only the nuclei of atoms. In a nuclear reaction, one element may change into another.

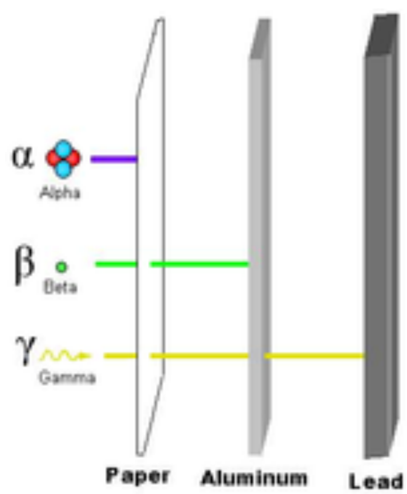
There are several types of radioactive decay, including alpha, beta, and gamma decay. In all three types, nuclei emit radiation, but the nature of the radiation differs. The Table below shows the radiation emitted in each type of decay.

Type	Radiation Emitted
Alpha decay	alpha particle (2 protons and 2 neutrons) + energy (in the form of electromagnetic radiation)
Beta decay	beta particle (1 electron) + energy (in the form of electromagnetic radiation)
Gamma decay	energy (gamma ray)

Both alpha and beta decay change the number of protons in an atom's nucleus, thereby changing the atom to a different element. In alpha decay, the nucleus loses two protons. In beta decay, the nucleus gains a proton as a neutron becomes a proton. In gamma decay, no change in proton number occurs, so the atom does not become a different element.

Penetrating Power of Radioactive Decay

The charged particles and energy emitted during radioactive decay can harm living things, but the three types of radioactive decay aren't equally dangerous. That's because they differ in how far they can travel and what they can penetrate. You can see this in the **Figure** below and in the animation at the following URL: http://www.cna.ca/curriculum/cna_radiation/gamma_rays-eng.asp?bc=Gamma_Rays&pid=Gamma_Rays



Alpha particles can travel only a few centimeters through air. They can burn the skin but not penetrate it. Beta particles can travel up to a meter through air. They can penetrate and damage skin. Gamma rays can travel thousands of meters through air. They can penetrate and damage cells deep inside the body.

SUMMARY

- Radioactive decay is the process in which unstable nuclei of radioactive atoms become stable by emitting charged particles and energy.
- There are three types of radioactive decay: alpha decay, beta decay, and gamma decay. Alpha and beta decay change one element into another. Gamma decay does not.
- Radioactive decay can damage living things. Alpha decay is the least damaging, and gamma decay is the most damaging.
- The most common emissions of radioactive elements were called alpha (α), beta (β), and gamma (γ)radiation.

REVIEW

1. What is radioactive decay?
2. Compare and contrast alpha, beta, and gamma decay.

3. Assume that you cut a sheet of paper down the center to get two halves. Then you cut each half down the center to get four pieces. If you keep cutting the pieces of paper in half, you would soon a reach a point where the pieces are too small to cut again. A radioactive isotope is a little like that sheet of paper.

HALF LIFE

What Is a Radioactive Isotope?

A radioactive isotope, or radioisotope, has atoms with unstable nuclei. The unstable nuclei naturally decay, or break down, by losing energy and particles of matter to become more stable. If they gain or lose protons as they decay, they become different elements. Over time, as the nuclei continue to decay, less and less of the original radioisotope remains.

Rate of Radioactive Decay

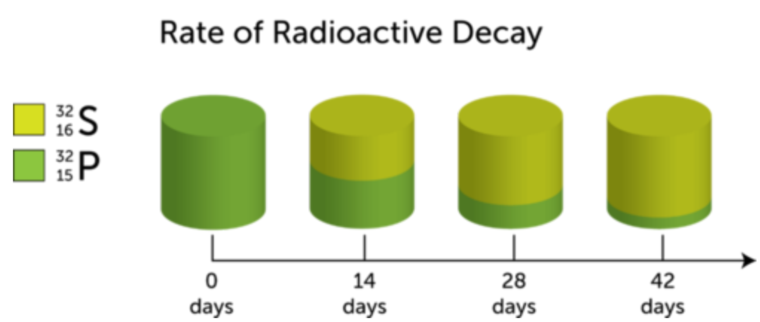
A radioisotope decays and changes to a different element at a constant rate. The rate is measured in a unit called the **half-life**. This is the length of time it takes for half of a given amount of the radioisotope to decay. This rate is always the same for a given radioisotope, regardless of temperature, pressure, or other conditions outside the nuclei of its atoms.

Q: How is repeatedly cutting paper in half like the decay of a radioisotope?

A: As a radioisotope decays, the amount of the radioisotope decreases by half during each half-life, just as a piece of paper decreases in size by half each time you cut it down the center. You can see a video of this half-life analogy at the following URL.
http://blip.tv/chemteam/an-analogy-for-half-life-4507204

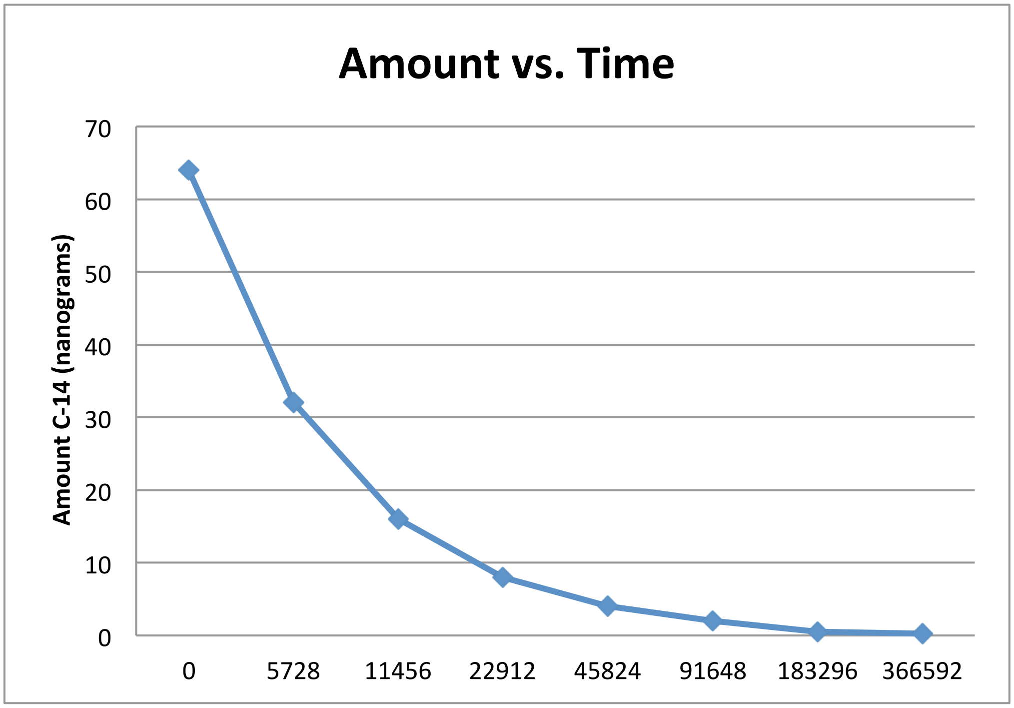
Half-Life Example

The concept of half-life is illustrated in the Figure below for the decay of phosphorus-32 to sulfur-32. The half-life of phosphorus-32 is 14 days. After 14 days, half of the original amount of phosphorus-32 has decayed, so only half remains. After another 14 days, half of the remaining amount (or a quarter of the original amount) is still left, and so on.



Q: What fraction of the original amount of phosphorus-32 remains after three half-lives?

A: After three half-lives, or 42 days, 1/8 (1/2 x 1/2 x 1/2) of the original amount of phosphorus-32 remains.



In the graph above, a nanogram sample of Carbon-14 decays over time. After 1 half life, 32 nanograms remains. After two half-lives about 15 nanograms remains at 11,456 years.

You can simulate radioactive decay of radioisotopes with different half-lives at the URL below.
http://www.colorado.edu/physics/2000/isotopes/radioactive_decay3.html

The following have a radioactive half-life: Uranium-238 has a half life of 4.47 billion years. Potassium-40 has a half-life of 1.28 billion years. Carbon-14 is 5,700 years, Hydrogen-3 is 12.3 years.

Q: If you had 1 gram of carbon-14, how many years would it take for radioactive decay to reduce it to 1/4 gram?

A: 1 gram would decay to ¼ gram in 2 half-lives. One half-life is 5,700 years, so two half-lives are 11,400 years.

SUMMARY

- A radioisotope decays and changes to a different element at a certain constant rate called the half-life. This is the length of time it takes for half of a given amount of the radioisotope to decay.
- Different radioisotopes may vary greatly in their rate of decay. The more unstable their nuclei are, the faster they decay.

PRACTICE

Complete the radioactivity worksheet at this URL:
http://zimearth.pbworks.com/f/Radioactivity+Worksheet.pdf

REVIEW

1. Define half-life.
2. Why do radioisotopes differ in the length of their half-lives?
3. What fraction of a given amount of hydrogen-3 would be left after 36.9 years of decay? (*Hint*: Find the half-life of hydrogen-3 in the table above.)

RADIOACTIVE DATING

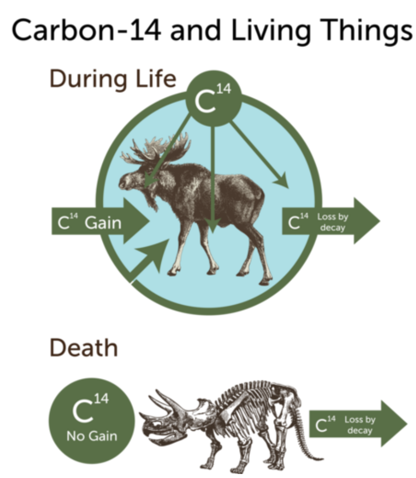
Different Isotopes, Different Half-Lives

Different radioisotopes decay at different rates. You can see some examples in the Table below. Radioisotopes with longer half-lives are used to date older rocks or other specimens, and those with shorter half-lives are used to date younger ones. For example, the oldest rocks at the bottom of the Grand Canyon were dated by measuring the amounts of potassium-40 in the rocks. Carbon-14 dating, in contrast, is used to date specimens that are much younger than the rocks in the Grand Canyon.

Parent Isotope	Daughter Isotope	Half-Life
potassium-40	argon-40	1.3 billion years
uranium-235	lead-207	700 million years
uranium-234	thorium-230	80,000 years
carbon-14	nitrogen-14	5,700 years

Focus on Carbon-14 Dating

One of the most familiar types of radioactive dating is carbon-14 dating. Carbon-14 forms naturally in Earth's atmosphere when cosmic rays strike atoms of nitrogen-14. Living things take in and use carbon-14, just as they do carbon-12. The carbon-14 in living things gradually decays to nitrogen-14. However, as it decays, it is constantly replaced because living things keep taking in carbon-14. As a result, there is a constant ratio of carbon-14 to carbon-12 in organisms as long as they are alive. This is illustrated in the top part of the **Figure** below.



After organisms die, the carbon-14 they already contain continues to decay, but it is no longer replaced (see the bottom part of the **Figure** below). Therefore, the carbon-14 in a dead organism constantly declines at a fixed rate equal to the half-life of carbon-14. Half of the remaining carbon-14 decays every 5,700 years. If you measure how much carbon-14 is left in a fossil, you can determine how many half-lives (and how many years) have passed since the organism died. Carbon-14 dating is illustrated in the video at this URL: http://www.youtube.com/watch?v=udkQwW6aLik

Q: Why can't carbon-14 dating be used to date specimens older than about 60,000 years?

A: Carbon-14 has a half-life of 5700 years. After about 60,000 years, too little carbon-14 is left in a specimen to be measured.

SUMMARY

- The age of a rock or other specimen can be estimated from the remaining amount of a radioisotope it contains and the radioisotope's known rate of decay, or half-life. This method of dating specimens is called radioactive dating.
- Radioisotopes with longer half-lives are used to date older specimens, and those with shorter half-lives are used to date younger ones.
- Carbon-14 dating is used to date specimens younger than about 60,000 years old. It is commonly used to date fossils of living things and human artifacts.

PRACTICE

Play the radioactive dating game at the following URL, and then answer the questions below.
http://phet.colorado.edu/en/simulation/radioactive-dating-game

1. What is the half-life of carbon-14? What is the half-life of uranium-238?
2. Compare the decay rates of carbon-14 and uranium-238. How long does it take for 75 percent of a sample of carbon-14 atoms to decay? How long does it take for 75 percent of a sample of uranium-238 to decay? Do these rates depend on the number of atoms in the samples?
3. What percentage of carbon-14 remains in a sample after 10,000 years? How many years does it take for uranium-238 to decay to this same percentage?
4. Why would you not use carbon-14 to measure the age of the rock?

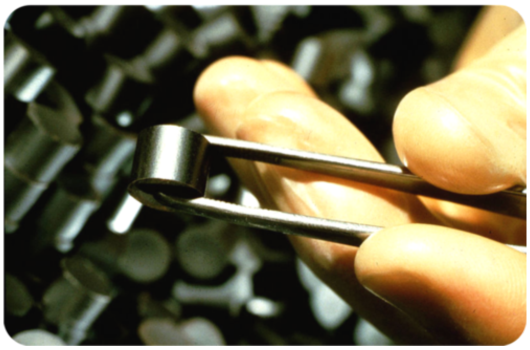
REVIEW

5. What is radioactive dating?
6. Which radioisotope in the table in the article (see above) could you use to date a fossil thought to be about 500 million years old? Explain your choice.
7. Why does the amount of carbon-14 in an organism remain the same throughout the organism's life? Why does the amount change after the organism dies?

ENERGY FROM NUCLEAR FISSION

Nuclear fission is the splitting of the nucleus of an atom into two smaller nuclei. This type of reaction releases a great deal of energy from a very small amount of matter. For example, nuclear fission of a tiny pellet of uranium-235, like the one pictured in **Figure** below, can release as much energy as burning 1,000 kilograms of coal!

Uranium-235

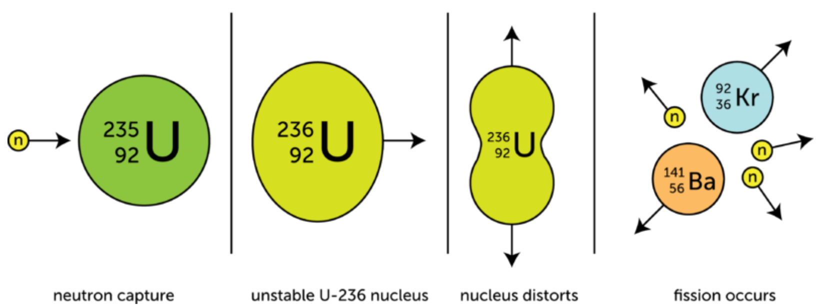


This pellet of uranium-235 can release a huge amount of energy if it undergoes nuclear fission.

Nuclear Energy by Fusion and Fission

As shown in **Figure** below, the reaction begins when a nucleus of uranium-235 absorbs a neutron. This can happen naturally or when a neutron is deliberately crashed into a uranium nucleus in a nuclear power plant. In either case, the nucleus of uranium becomes very unstable and splits in two. In this example, it forms krypton-92 and barium-141. The reaction also releases three neutrons and a great deal of energy.

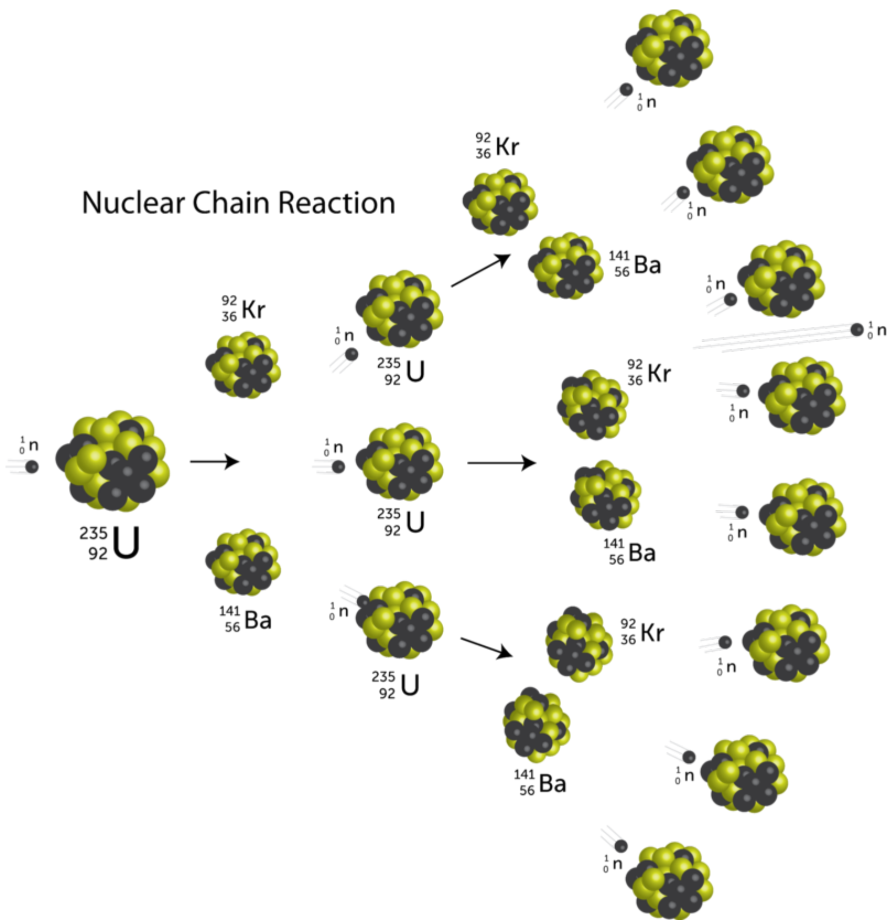
Nuclear Fission



The fissioning of a nucleus of uranium-235 begins when it captures a neutron.

Nuclear Chain Reaction

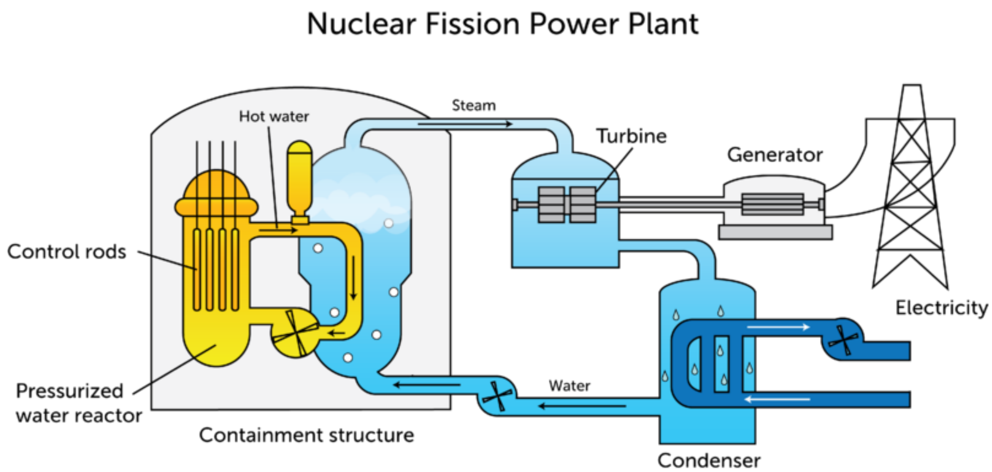
The neutrons released in this nuclear fission reaction may be captured by other uranium nuclei and cause them to fission as well. This can start a nuclear chain reaction (see **Figure** below). In a chain reaction, one fission reaction leads to others, which lead to others, and so on. A nuclear chain reaction is similar to a pile of wood burning. If you start one piece of wood burning, enough heat is produced by the burning wood to start the rest of the pile burning without any further help from you. You can see another example of a chain reaction at this URL:
http://www.youtube.com/watch?v=0v8i4v1mieU (2:54).



In a nuclear chain reaction, each nuclear reaction leads to more nuclear reactions.

Using Energy from Nuclear Fission

If a nuclear chain reaction is uncontrolled, it produces a lot of energy all at once. This is what happens in an atomic bomb. If a nuclear chain reaction is controlled, it produces energy more slowly. This is what occurs in a nuclear power plant. The reaction may be controlled by inserting rods of material that do not undergo fission into the core of fissioning material (see **Figure** below). The radiation from the controlled fission is used to heat water and turn it to steam. The steam is under pressure and causes a turbine to spin. The spinning turbine runs a generator, which produces electricity.



This diagram shows the main parts of a nuclear power plant.

Pros and Cons of Nuclear Fission

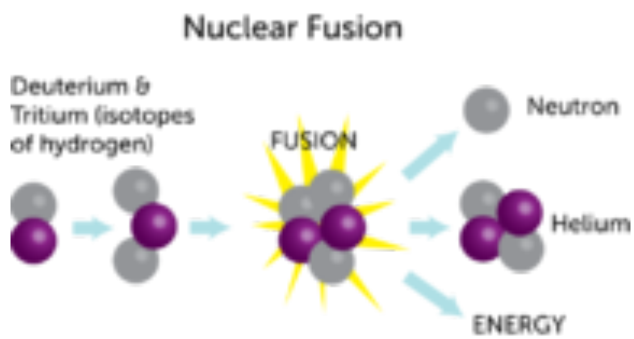
In the U.S., the majority of electricity is produced by burning coal or other fossil fuels. This causes air pollution, acid rain, and global warming. Fossil fuels are also limited and may eventually run out. Like fossil fuels, radioactive elements are limited. In fact, they are relatively rare, so they could run out sooner rather than later. On the other hand, nuclear fission does not release air pollution or cause the other environmental problems associated with burning fossil fuels. This is the major advantage of using nuclear fission as a source of energy.

The main concern over the use of nuclear fission is the risk of radiation. Accidents at nuclear power plants can release harmful radiation that endangers people and other living things. Fukishima nuclear reactor in Japan exploded in 2012 when radioactive water leaked out of the chamber and produced Hydrogen, built up in the chamber, which then exploded. Fish were contaminated in the surrounding ocean. Even without accidents, the used fuel that is left after nuclear fission reactions is still radioactive and very dangerous. It takes thousands of years for it to decay until it no longer releases harmful radiation.

Therefore, used fuel must be stored securely to people and other living things. You can learn more about the problem of radioactive waste at this URL: http://www.youtube.com/watch?v=0PQ97LVRuuM.

Energy from Nuclear Fusion

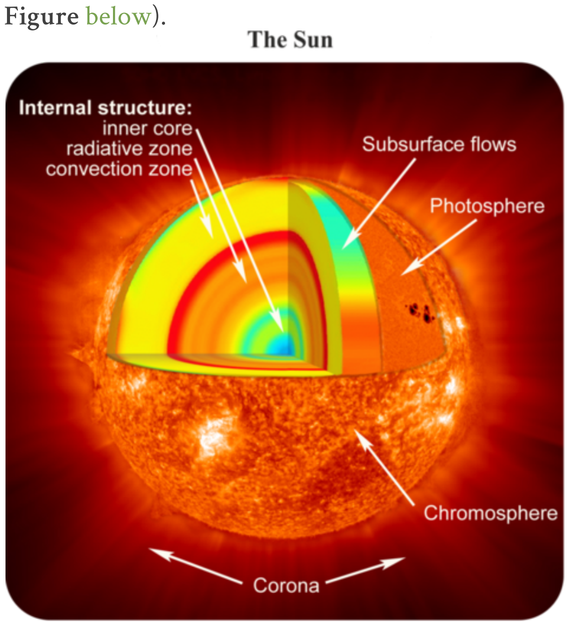
Nuclear fusion is the opposite of nuclear fission. In fusion, two or more small nuclei combine to form a single, larger nucleus. An example is shown in **Figure** below. In this example, two hydrogen nuclei fuse to form a helium nucleus. A neutron and a great deal of energy are also released. In fact, fusion releases even more energy than fission does.



In this nuclear fusion reaction, nuclei of two hydrogen isotopes (tritium and deuterium) fuse together. They form a helium nucleus, a neutron, and energy.

The Power of Stars

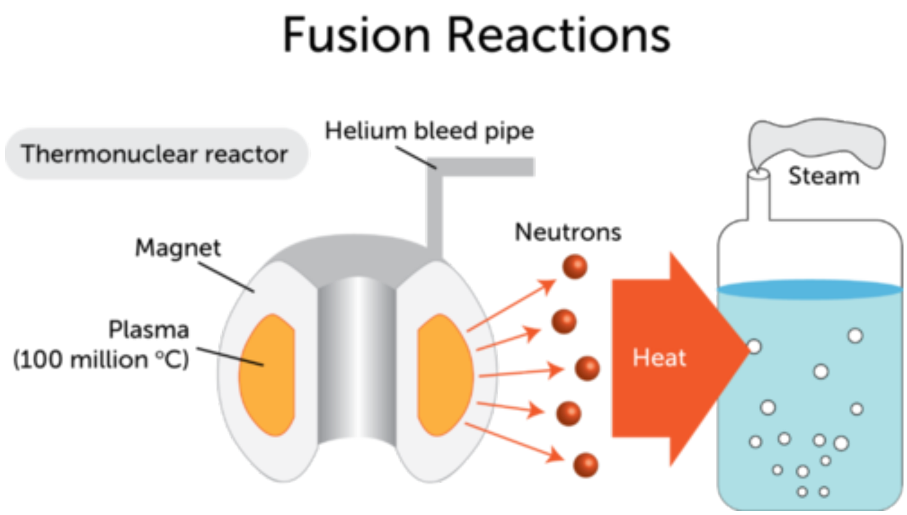
Nuclear fusion of hydrogen to form helium occurs naturally in the sun and other stars. It takes place only at extremely high temperatures. That's because a great deal of energy is needed to overcome the force of repulsion between positively charged nuclei. The sun's energy comes from fusion in its core, where temperatures reach millions of Kelvin (see **Figure** below).



The extremely hot core of the sun radiates energy from nuclear fusion.

Using Nuclear Fusion

Scientists are searching for ways to create controlled nuclear fusion reactions on Earth. Their goal is develop nuclear fusion power plants, where the energy from fusion of hydrogen nuclei can be converted to electricity. How this might work is shown in **Figure** below.



In the thermonuclear reactor modeled here, radiation from fusion is used to heat water and form steam. The steam can then be used to turn a turbine and generate electricity.

The use of nuclear fusion for energy has several pros. Unlike nuclear fission, which involves dangerous radioisotopes, nuclear fusion involves hydrogen and helium. These elements are harmless. Hydrogen is also very plentiful. There is a huge amount of hydrogen in ocean water. The hydrogen in just a gallon of water could produce as much energy by nuclear fusion as burning 1,140 liters (300 gallons) of gasoline! The hydrogen in the oceans would generate enough energy to supply the entire world's people for a very long time. Unfortunately, using energy from nuclear fusion is far from a reality. Scientists are a long way from developing the necessary technology. One problem is raising temperatures high enough for fusion to take place. Another problem is that matter this hot exists only in the plasma state. There are no known materials that can contain plasma, although a magnet might be able to do it. That's because plasma consists of ions and responds to magnetism. You can learn more about research on nuclear fusion at the URL below.

http://www.youtube.com/watch?v=3C5hFQeZCT4&feature=related

Chemical Versus Nuclear Energy

The energy produced in a nuclear reaction is thousands of times more than the energy released by a chemical reaction. The atomic bomb over Hiroshima vaporized everything in a mile of the blast and blew over people who were living fifteen miles away. A chemical reaction from dynamite will blow out windows several blocks away and melt items in the immediate surroundings.

The effects of radiation on humans and other living things:

You may have seen this sign before—maybe in a hospital. The sign means there is danger of radiation in the area. **Radiation** consists of particles and energy that are given off by radioactive isotopes, which have unstable nuclei. But you don't have to go to a hospital to be exposed to radiation. There is radiation in the world all around you.

Radiation in the Environment

A low level of radiation occurs naturally in the environment. This is called background radiation. One source of background radiation is rocks, which may contain small amounts of radioactive elements such as uranium. Another source is cosmic rays. These are charged particles that arrive on Earth from outer space. Background radiation is generally considered to be safe for living things. You can learn more about background radiation with the animation at this URL:
http://www.pbs.org/wgbh/nova/dirtybomb/sources.html

Dangers of Radiation

Long-term or high-dose exposure to radiation can harm both living and nonliving things. Radiation knocks electrons out of atoms and changes them to ions. It also breaks bonds in DNA and other compounds in living things. One source of radiation that is especially dangerous to people is radon. Radon is a radioactive gas that forms in rocks underground. It can seep into basements and get trapped inside buildings. Then it may build up and become harmful to people who breathe it. Long-term exposure to radon can cause lung cancer.

Exposure to higher levels of radiation can be very dangerous, even if the exposure is short-term. A single large dose of radiation can burn the skin and cause radiation sickness. Symptoms of this illness include extreme fatigue, destruction of blood cells, and loss of hair. To learn more about the harmful health effects of radiation, go to this URL: http://www.cbsnews.com/video/watch/?id=7359819n

Nonliving things can also be damaged by radiation. For example, high levels of radiation can weaken metals by removing electrons. This is a problem in nuclear power plants and space vehicles because they are exposed to very high levels of radiation.

Q: Can you tell when you are being exposed to radiation? For example, can you sense radon in the air?

A: Radiation can't be detected with the senses. This adds to its danger. However, there are other ways to detect it.

Detecting Radiation

You generally can't see, smell, taste, hear, or feel radiation. Fortunately, there are devices such as Geiger counters that can detect radiation. A Geiger counter, like the one pictured in the **Figure** below, contains atoms of a gas that is ionized if it encounters radiation. When this happens, the gas atoms change to ions that can carry an electric current. The current causes the Geiger counter to click. The faster the clicks occur, the higher the level of radiation. You can see a video about the Geiger counter and how it was invented at the URL below.
http://vimeo.com/10379389

Using Radiation

Despite its dangers, radioactivity has several uses. For example, it can be used to determine the ages of ancient rocks and fossils. It can also be used as a source of power to generate electricity. Radioactivity can even be used to diagnose and treat diseases, including cancer. Cancer cells grow rapidly and take up a lot of glucose for energy. Glucose containing radioactive elements can be given to patients. Cancer cells take up more of the glucose than normal cells do and give off radiation. The radiation can be detected with special machines like the one in the **Figure** below. The radiation may also kill cancer cells. You can learn more about medical uses of radiation at this URL: http://www.youtube.com/watch?v=v_8xM-mLxJ8

This machine scans a patient's body and detects radiation.

SUMMARY

- A low level of radiation occurs naturally in the environment. This background radiation is generally assumed to be safe for living things.
- Long-term or high-dose exposure to radiation can harm living things and damage nonliving materials such as metals.
- One reason radiation is dangerous is that it generally can't be detected with the senses. It can be detected only with devices such as Geiger counters.
- Radiation has several important uses, including diagnosing and treating cancer.

VOCABULARY

radiation: Energy emitted by the nucleus of a radioisotope or an accelerating particle; transfer of energy by electromagnetic waves that can travel across space as well as through matter.

PRACTICE

Watch the video about uses of radiation at the following URL, and then answer the questions below.
http://www.youtube.com/watch?v=TdbzShLU3Ow

1. Alpha radiation is used in smoke alarms. Explain how a smoke alarm uses this form of radiation.
2. Identify a use of beta radiation, and explain how it works.
3. List three uses of gamma radiation, and describe one of them in detail.

REVIEW

4. What are two sources of background radiation?
5. How can radiation harm living things?
6. What is radon, and why is it harmful to people?
7. How does a Geiger counter detect radiation?
8. What are some uses of radiation?

GLOSSARY

Atom: smallest particle of matter, made of a certain number of protons, neutrons, and electrons.

Element: Type of matter that behaves a certain way in physical and chemical behaviors. There are 92 naturally occurring elements.

Nucleus: The center of an atom containing proton and neutron.

Neutron: A neutral particle having a relative mass of 1 atomic mass unit.

Electron: A negatively charged particle have a relative mass of 1/1846 that of a proton or neutron. It is found outside the nucleus in one of seven orbits.

Isotope: An atom of the same element that has the same number of protons but a different number of neutrons, thus a different atomic mass. Oxygen-17 is an isotope of Oxygen-16.

Periodic table: An arrangement of elements by atomic number that group elements based on similar behavior.

Quanta: An amount of energy that must be obtained before an electron can be moved from one energy level to another.

Wavelength: The distance from one crest of a wave to the next crest, or from a trough to the next trough.

Radiation: Alpha particles, Beta particles, and Gamma rays that produce heat when passed through another body.

Emit: put off

Absorb: taken in

Spectrum: A range of particles or energy from radio waves, microwaves, infrared waves, ultraviolet light to visible light to x-rays to gamma rays.

Half-life: The time it takes for half the radioactive material to decay or break down.

Fission: Splitting of an atom, which releases a large amount of energy

Fusion: Joining two nuclei of atoms, which releases the largest amount of energy

Energy level: there are seven areas around the nucleus of an atom in which electrons fly.

CHAPTER 3

CHEMICAL BONDING

Standard 3: Students will understand chemical bonding and the relationship of the type of bonding to the chemical and physical properties of substances.

Lesson Vocabulary

- chemical bond
- chemical formula
- physical property
- chemical property

STANDARD 3, OBJECTIVE 1: ANALYZE THE RELATIONSHIP BETWEEN THE VALENCE (OUTERMOST) ELECTRONS OF AN ATOM AND THE TYPE OF BOND FORMED BETWEEN ATOMS.

INTRODUCTION TO CHEMICAL BONDING

Lesson Objectives

- Define chemical bond.
- List general properties of compounds

Introduction

There is an amazing diversity of matter in the universe, but there are only about 100 elements. How can this relatively small number of pure substances make up all kinds of matter? Elements can combine in many different ways. When they do, they form new substances called compounds. For a video introduction to compounds, go to this URL: **http://www.youtube.com/watch?v=-HjMoTthEZ0&feature=related (3:53).**

Chemical Bonding

Elements form compounds when they combine chemically. Their atoms join together to form molecules, crystals, or other structures. The atoms are held together by chemical bonds. A chemical bond is a force of attraction between atoms or ions. It occurs when atoms share or transfer valence electrons. Valence electrons are the electrons in the outer energy level of an atom. You can learn more about chemical bonds in this video: **http://www.youtube.com/watch?v=CGA8sRwqIFg&NR=1 (13:21).**
You will read more about valence electrons in the next section.

Chemical Compounds

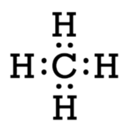
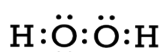
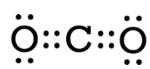
Water (H_2O) is an example of a chemical compound. Water molecules always consist of two atoms of hydrogen and one atom of oxygen. Like water, all other chemical compounds consist of a fixed ratio of elements. It doesn't matter how much or how little of a compound there is. It always has the same composition.

It is important to know that when atoms combine to form compounds, their chemical and physical properties change. A **physical property** is an observable measurable property. Appearance, color, texture, melting point, boiling point, and density are all physical properties. A **chemical property** describes an atom's potential to change through a chemical reaction. The ability to combust, rust, or bubble in acid are all examples of chemical properties. Table salt, called sodium chloride, is formed by bonding a sodium atom (Na) to a chlorine atom (Cl). Elemental sodium is a soft metal that reacts with water to produce a flammable gas. Elemental chlorine is a gas at room temperature and is poisonous. These two nasty chemicals join together to form table salt (NaCl); a substance we eat most every day. Na and Cl atoms do not have the same properties as NaCl. Atoms do not have the same chemical and physical properties as when they join to form compounds.

To see the reaction between Na and Cl go to the following: http://www.youtube.com/watch?v=Mx5JJWI2aaw

Chemical Formulas

Elements are represented by chemical symbols. Examples are H for hydrogen and O for oxygen. Compounds are represented by **chemical formulas**. You've already seen the chemical formula for water. It's H_2O. The subscript 2 after the H shows that there are two atoms of hydrogen in a molecule of water. The O for oxygen has no subscript. When there is just one atom of an element in a molecule, no subscript is used. **Table** below shows some other examples of compounds and their chemical formulas. The table also shows Electron Dot Diagram which will be discussed later in the chapter.

Name of Compound	Electron Dot Diagram	Numbers of Atoms	Chemical Formula
Hydrogen chloride		H = 1 Cl = 1	HCl
Methane		C = 1 H = 4	CH_4
Hydrogen peroxide		H = 2 O = 2	H_2O_2
Carbon dioxide		C = 1 O = 2	CO_2

Problem Solving

Problem: A molecule of ammonia consists of one atom of nitrogen (N) and three atoms of hydrogen (H). What is its chemical formula?
Solution: The chemical formula is NH3.

You Try It!

Problem: A molecule of nitrogen dioxide consists of one atom of nitrogen (N) and two atoms of oxygen (O). What is its chemical formula?

Same Elements, Different Compounds

The same elements may combine in different ratios. If they do, they form different compounds. **Figure** below shows some examples. Both water (H_2O) and hydrogen peroxide (H_2O_2) consist of hydrogen and oxygen. However, they have different ratios of the two elements. As a result, water and hydrogen peroxide are different **compounds** with different properties. If you've ever used hydrogen peroxide to disinfect a cut, then you know that it is very different from water! Both carbon dioxide (CO_2) and carbon monoxide (CO) consist of carbon and oxygen, but in different ratios. How do their properties differ?

Different compounds may contain the same elements in different ratios. How does this affect their properties?

Types of Compounds

There are different types of compounds. They differ in the nature of the bonds that hold their atoms together. The type of bonds in a compound determines many of its properties. Three types of bonds are ionic, covalent, and metallic bonds. You will read about these three types in later lessons.

LESSON SUMMARY

- A chemical bond is a force of attraction between atoms. It occurs when atoms share or transfer electrons.
- A chemical compound is a new substance that forms when atoms of different elements form chemical bonds. A compound always consists of a fixed ratio of elements.

LESSON REVIEW QUESTIONS

1. What is a chemical bond?
2. Define chemical compound.
3. What is the difference between chemical and physical properties?
4. How do the chemical and physical properties differ between compounds and the atoms they are composed of?
5. Which atoms and how many of each make up a molecule of sulfur dioxide? Write the chemical formula for this compound.
6. Why does a molecule of water have a more stable arrangement of electrons than do individual hydrogen and oxygen atoms?
7. In this lesson, you learned about chemical bonds in a water molecule. The bonds form between atoms of hydrogen and oxygen when they share electrons. This type of bond is an example of a covalent bond.

What might be other ways that atoms can bond together?
How might ions form bonds?

VALENCE ELECTRONS

Lesson Objectives:

- Know what valence electrons are and how to determine the number of valence electrons an atom has
- Understand how valence electrons determine chemical reactivity

Vocabulary

- valence electron

Introduction

Did you ever play the card game called go fish? Players try to form groups of cards of the same value, such as four sevens, with the cards they are dealt or by getting cards from other players or the deck. This give and take of cards is a simple analogy for the way atoms give and take valence electrons in chemical reactions.

What Are Valence Electrons?

To understand chemical bonding, we first must understand valence electrons. **Valence electrons** are the electrons in the outer energy level of an atom that can participate in interactions with other atoms. Valence electrons are generally the electrons that are farthest from the nucleus. As a result, they may be attracted as much or more by the nucleus of another atom than they are by their own nucleus.

Electron Dot Diagrams

Because valence electrons are so important, atoms are often represented by simple diagrams that show only their valence electrons. These are called electron dot diagrams, and two are shown below.

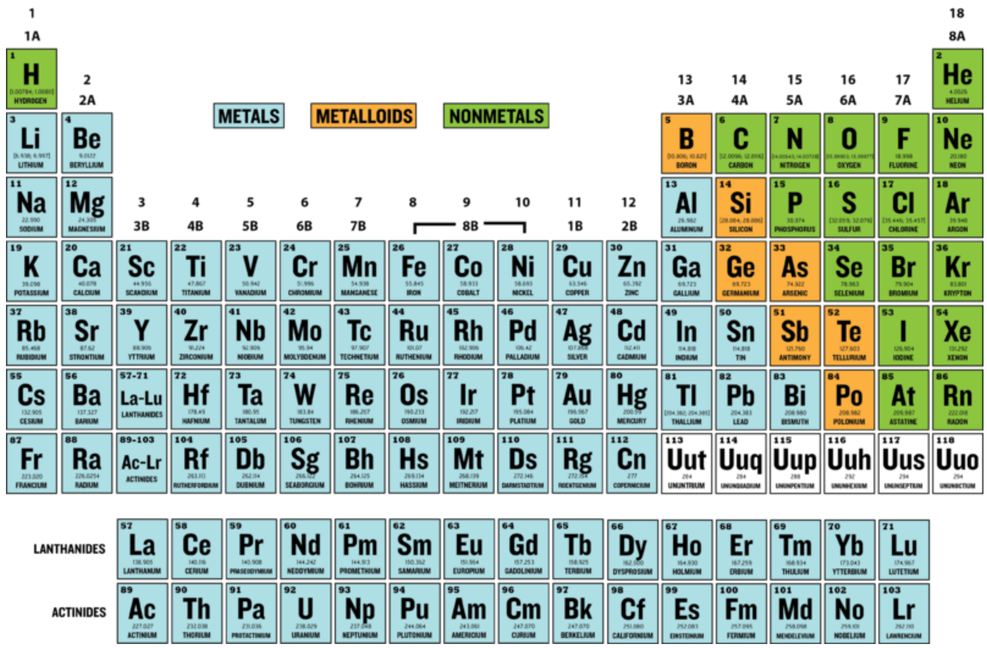
In this type of diagram, an element's chemical symbol is surrounded by dots that represent the valence electrons. Typically, the dots are drawn as if there is a square surrounding the element symbol with up to two dots per side. An element never has more than eight valence electrons, so there can't be more than eight dots per atom.

Q: Carbon (C) has four valence electrons. What does an electron dot diagram for this element look like?

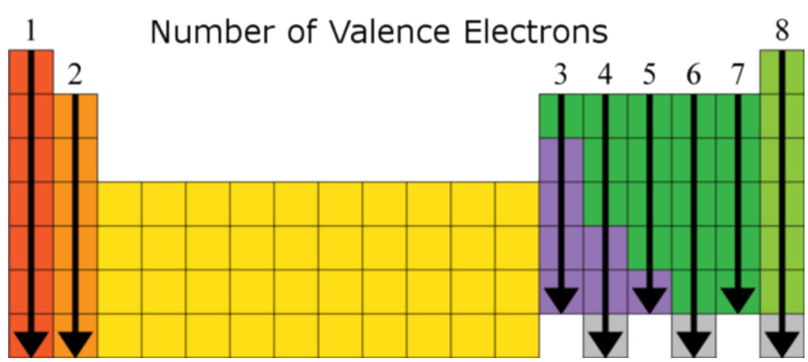
A: An electron dot diagram for carbon looks like this (you draw it in):

Valence Electrons and the Periodic Table

The number of valence electrons in an atom is reflected by its position in the periodic table of the elements (see the periodic table below). Across each row, or period, of the periodic table, the number of valence electrons in groups 1–2 and 13–18 increases by one from one element to the next. Within each column, or group, of the table, all the elements have the same number of valence electrons. This explains why all the elements in the same group have very similar chemical properties.



For elements in groups 1–2 and 13–18, the number of valence electrons is easy to tell directly from the periodic table. This is illustrated in the simplified periodic table in **Figure** below. It shows just the numbers of valence electrons in each of these groups. For elements in groups 3–12, determining the number of valence electrons is more complicated and goes beyond the scope of this course.



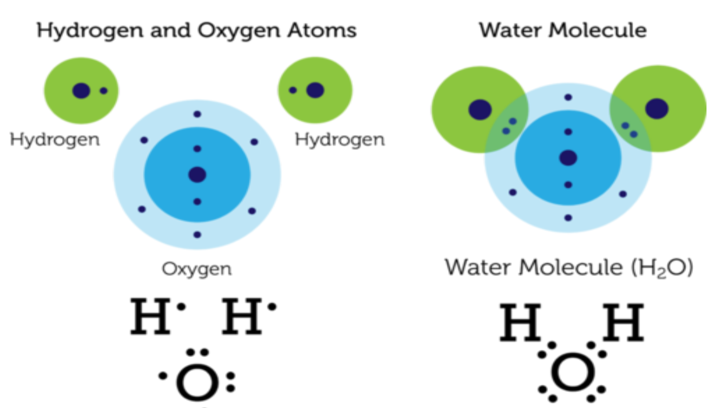
Q: Based on both periodic tables above, what are examples of elements that have just one valence electron? What are examples of elements that have eight valence electrons? How many valence electrons does oxygen (O) have?

A: Any element in group 1 has just one valence electron. Examples include hydrogen (H), lithium (Li), and sodium (Na). Any element in group 18 has eight valence electrons (except for helium, which has a total of just two electrons). Examples include neon (Ne), argon (Ar), and krypton (Kr). Oxygen, like all the other elements in group 16, has six valence electrons.

Valence Electrons and Reactivity

The number of valence electrons of an atom will determine its reactivity with other atoms. Remember that atoms like to be stable and that means they have full outer shells. In other words, stable atoms have 8 valence electrons, or an octet. The noble gas elements are the least reactive of all of the elements on the periodic table, because they already have eight valence electrons. Atoms will bond or react with other atoms to become stable with full outer shells, or full octets. Some atoms will be more reactive than other atoms depending on their valence electrons. For example, fluorine is highly reactive because it has seven valence electrons and only needs one more electron to be stable.

Look at the example of water in Figure below. A water molecule consists of two atoms of hydrogen and one atom of oxygen. Each hydrogen atom has just one electron. The oxygen atom has six valence electrons. In a water molecule, two hydrogen atoms share their two electrons with the six valence electrons of one oxygen atom. By sharing electrons, each atom has electrons available to fill its outer energy level. This gives it a more stable arrangement of electrons that takes less energy to maintain.



These diagrams show the valence electrons of hydrogen and water atoms and a water molecule. The diagrams represent electrons with dots, so they are called electron dot diagrams.

SUMMARY

- Valence electrons are the electrons in the outer energy level of an atom that can participate in interactions with other atoms.
- Because valence electrons are so important, atoms may be represented by electron dot diagrams that show only their valence electrons.
- The number of valence electrons in atoms may cause them to be unreactive or highly reactive. Valence electrons will determine the bond type.

REVIEW

1. What are valence electrons?
2. Draw an electron dot diagram for an atom of nitrogen (N).
3. Which of the following statements about valence electrons and the periodic table is true?
    a. The number of valence electrons decreases from left to right across each period.
    b. The number of valence electrons increases from top to bottom within each group.
    c. All of the elements in group 9 have nine valence electrons.
    d. Elements with the most valence electrons are in group 18.
4. Which element would you expect to be more reactive: phosphorus (P) or fluorine (F)? Explain your answer

ION AND ION FORMATION

Lesson Objectives

- explain why atoms form ions.
- identify the atoms most likely to form positive ions and the atoms most likely to form negative ions.
- given the symbol of a main group element, indicate the most likely number of electrons the atom will gain or lose.
- predict the charge on ions
- describe what polyatomic ions are.
- given the formula of a polyatomic ion, name it, and vice versa.

Vocabulary

- ion
- cation,
- anion

The incredible green lights in this cold northern sky consist of charged particles known as ions. Their swirling pattern is caused by the pull of Earth's magnetic north pole. Called the northern lights, this phenomenon of nature shows that ions respond to a magnetic field. Do you know what ions are? In this section you will learn about ions and how they are formed.

Atoms Are Neutral

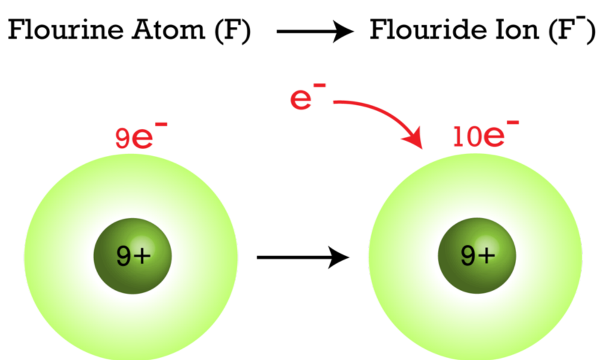
The northern lights aren't caused by atoms, because atoms are not charged particles and therefore would not be attracted to a magnetic field. An atom always has the same number of electrons as protons. Electrons have an electric charge of -1 and protons have an electric charge of +1. Therefore, the charges of an atom's electrons and protons "cancel out." This explains why atoms are neutral in electric charge.

Q: What would happen to an atom's charge if it were to gain extra electrons?

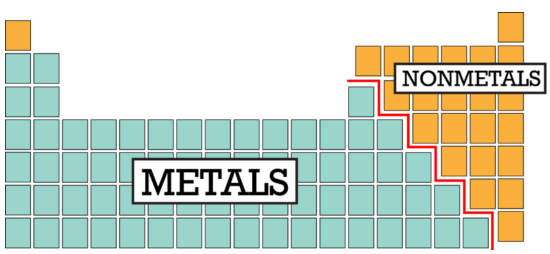
A: If an atom were to gain extra electrons, it would have more electrons than protons. This would give it a negative charge, so it would no longer be neutral.

Atoms to Ions

Atoms cannot only gain extra electrons. They can also lose electrons. In either case, they become **ions**. Ions are atoms that have a positive or negative charge because they have unequal numbers of protons and electrons. Atoms will only gain or lose electrons in this process, the amount of protons will stay the same. If atoms lose electrons, they become positive ions, or **cations**. If atoms gain electrons, they become negative ions, or **anions**. Consider the example of fluorine (see **Figure** below). A fluorine atom has nine protons and nine electrons, so it is electrically neutral. If a fluorine atom gains an electron, it becomes a fluoride ion with an electric charge of -1.



Metals and Nonmetals



As mentioned above, atoms lose or gain electrons to become stable.

Which atoms gain electrons and which atoms lose electrons? Metals, the atoms found on the left side of the table, tend to lose electrons and become cations; while nonmetals tend to gain electrons and become anions. Noble gases do not form ions.

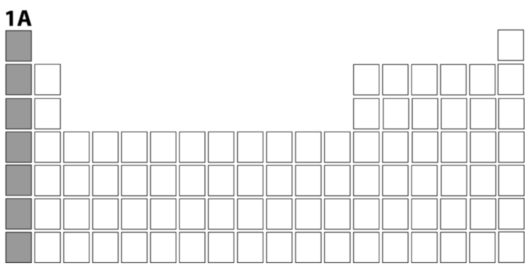
Q: Why do you think atoms lose electrons to, or gain electrons from, other atoms?

A: Atoms form ions by losing or gaining electrons because it makes them more stable. The most stable state for an atom is to have its outermost energy level filled with the maximum possible number of electrons. In the case of metals such as lithium, with just

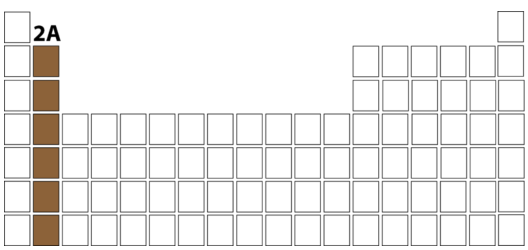
one valence electron in the outermost energy level, a more stable state can be achieved by losing that one outer electron. In the case of nonmetals such as fluorine, which has seven valence electrons in the outermost energy level, a more stable state can be achieved by gaining one electron and filling up the outer energy level. You can learn more about why ions form by watching the video at this URL:
http://www.youtube.com/watch?v=CV53wfl-oV8 (9:35)

Some Common Ions

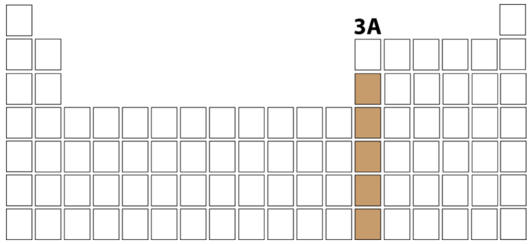
All the metals in family 1A (shown in the figure below) have one valence electron. The entire family forms +1 ions: Li^+, Na^+, K^+, Rb^+, Cs^+, and Fr^+. Note that although hydrogen (H) is in this same column, it is not considered to be a metal. There are times when hydrogen acts like a metal and forms $+1$ ions, but most of the time it bonds with other atoms as a nonmetal. In other words, hydrogen doesn't easily fit into any chemical family.



The metals in family 2A all have two valence electrons. This entire family will form +2 : Be^{+2}, Mg^{+2}, Ca^{+2}, Sr^{+2}, Ba^{+2}, Ra^{+2}.



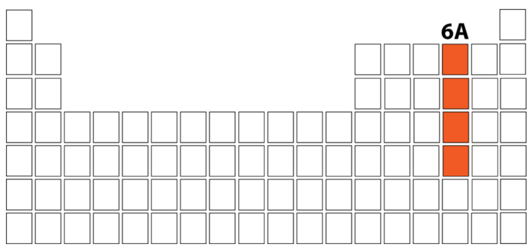
Family 3A have three valence electrons. When these atoms form ions, they will almost always form +3 ions: Al^{+3}, Ga^{+3}, In^{+3}, Ti^{+3}. Notice that boron is omitted from this list. This is because boron falls on the nonmetal side of the metal/nonmetal dividing line. Boron generally doesn't lose all of its valence electrons during chemical reactions.



Family 4A is almost evenly divided into metals and nonmetals. The larger atoms in the family (germanium, tin, and lead) are metals. Since these atoms have 4 valence electrons, they are expected to form ions with charges of +4. All three of the atoms do form such ions (Ge^{4+}, Sn^{4+} and Pb^{4+}), but tin and lead also have the ability to also form +2 ions. Like family 4A, the elements of family 5A are also divided into metals and nonmetals. The smaller atoms in this family behave as nonmetals and form -3 ions, and the larger atoms behave as metals that form +5 ions. For the nonmetals, they each have 5 valence electrons so they will need to gain 3 more hence, the -3 charge.

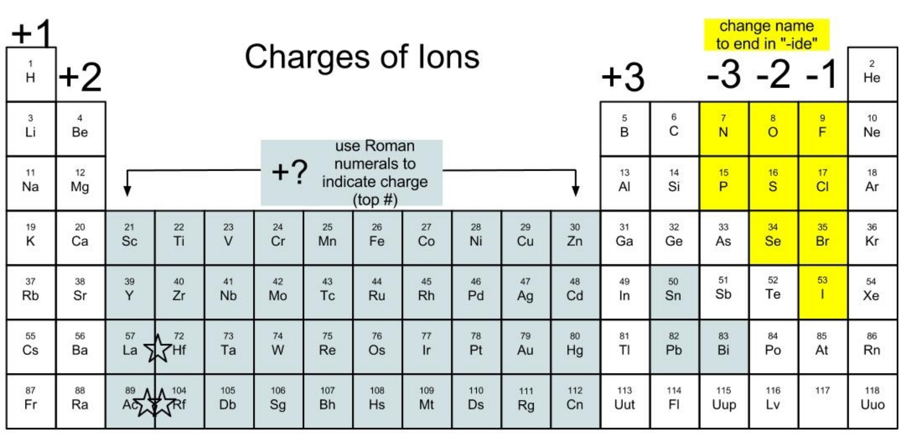
Most of the elements in family 6A (shown in figure below) are nonmetals that have 6 valence electrons. They form –2 ions. If you consider that each has six valence electrons, they will need to gain two more to become stable.

: O^{-2}, S^{-2}, Se^{-2} and Te^{-2}.



Family 7A are all nonmetals. When these atoms form ions, they form –1 ions: F^{-}, Cl^{-}, Br^{-} and I^{-}. They each have seven valence electrons therefore; they need to gain one more to be stable.

Family 8A, of course, is made up of the noble gases, which have no tendency to either gain or lose electrons.



LESSON SUMMARY

- Ions are atoms or groups of atoms that carry electrical charge.
- Metals tend to lose electrons and non-metals tend to gain electrons.
- Atoms that tend to lose electrons will generally lose all the electrons in their outermost energy level.
- Atoms that tend to gain electrons will gain enough electrons to completely fill their outermost energy level.

REVIEW

1. Why are atoms neutral in electric charge?
2. Compare and contrast cations and anions, and give an example of each.
3. Describe how ions form.
4. If the lithium atom becomes an ion, which type of ion will it be, a cation or an anion? What will be the electric charge of this ion? What will the ion be named? What symbol will be used to represent it?
5. How is the number of valence electrons of a metal atom related to the charge on the ion the metal will form?
6. How is the number of valence electrons of a nonmetal related to the charge on the ion the nonmetal will form?
7. If carbon were to behave like a metal and give up electrons, how many electrons would it give up?

PRACTICE

At the following URL, scroll down to the middle of the page and download "Ion Worksheet." Then fill in the missing information in the worksheet.
http://www.powayusd.com/teachers/kvalentine/worksheetspage.htm

STANDARD 3, OBJECTIVE 2: EXPLAIN THAT THE PROPERTIES OF A COMPOUND MAY BE DIFFERENT FROM THOSE OF THE ELEMENTS OR COMPOUNDS FROM WHICH IT IS FORMED.

Lesson Vocabulary

- ionic bond
- ionic compound
- polyatomic ion

CHEMICAL BONDING

IONIC BONDING

Lesson Objectives

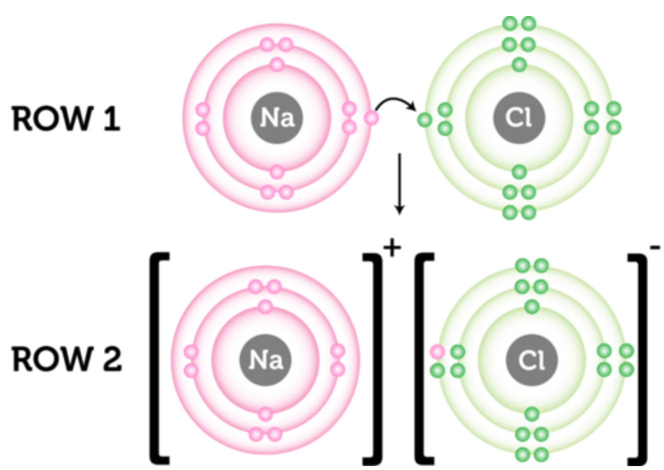
- Describe how ionic bonds form.
- List properties of ionic compounds.
- Name and write the formulas for ionic compounds

Introduction

All compounds form when atoms of different elements share or transfer electrons. In water, the atoms share electrons. In some other compounds, called **ionic compounds**, atoms transfer electrons. The electrons actually move from one atom to another. When atoms transfer electrons in this way, they become charged particles called ions. The ions are held together by ionic bonds.

Formation of Ionic Bonds

An **ionic bond** is the force of attraction that holds together positive and negative ions. It forms when atoms of a metallic element give up electrons to atoms of a nonmetallic element. **Figure** below shows how this happens.



An ionic bond forms when the metal sodium gives up an electron to the nonmetal chlorine.

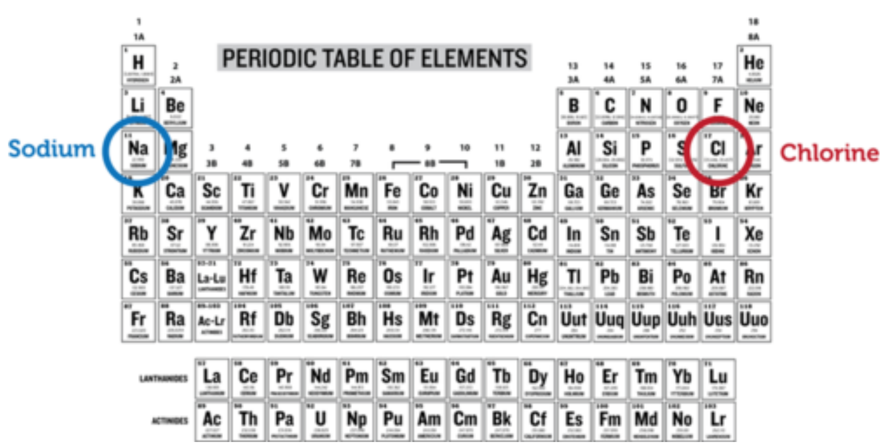
In row 1 of Figure opposite, an atom of sodium donates an electron to an atom of chlorine (Cl).

- By losing an electron, the sodium atom becomes a sodium ion. It now has one less electron than protons, giving it a charge of +1. Positive ions such as sodium are given the same name as the element. The chemical symbol has a plus sign to distinguish the ion from an atom of the element. The symbol for a sodium ion is Na^+.
- By gaining an electron, the chlorine atom becomes a chloride ion. It now has one more electron than protons, giving it a charge of -1. Negative ions are named by adding the suffix –*ide* to the first part of the element name. The symbol for chloride is Cl^-.

Sodium and chloride ions have equal but opposite charges. Opposites attract, so sodium and chloride ions attract each other. They cling together in a strong ionic bond. You can see this in row 2 of Figure above. Brackets separate the ions in the diagram to show that the ions in the compound do not share electrons. You can see animations of sodium chloride forming at these URLs:
http://web.jjay.cuny.edu/~acarpi/NSC/salt.htm
http://www.visionlearning.com/library/module_viewer.php?mid=55

Why Ionic Bonds Form

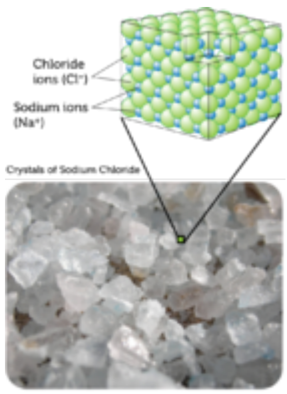
Ionic bonds form only between metals and nonmetals. Metals "want" to give up electrons, and nonmetals "want" to gain electrons. Find sodium (Na) in **Figure** below. Sodium is an alkali metal in group 1. Like other group 1 elements, it has just one valence electron. If sodium loses that one electron, it will have a full outer energy level. Now find fluorine (F) in **Figure** below. Fluorine is a halogen in group 17. It has seven valence electrons. If fluorine gains one electron, it will have a full outer energy level. After sodium gives up its valence electron to fluorine, both atoms have a more stable arrangement of electrons.



Sodium and chlorine are on opposite sides of the periodic table. How is this related to their numbers of valence electrons?

Ionic Compounds

Ionic compounds contain ions of metals and nonmetals held together by ionic bonds. Ionic compounds do not form molecules. Instead, many positive and negative ions bond together to form a structure called a crystal. You can see an example of a crystal in **Figure** below. It shows the ionic compound sodium chloride. Positive sodium ions (Na^+) alternate with negative chloride ions (Cl^-). The oppositely charged ions are strongly attracted to each other.



Sodium chloride crystals are cubic in shape. Other ionic compounds may have crystals with different shapes.

Properties of Ionic Compounds

The rigid crystals are brittle and more likely to break than bend when struck. As a result, ionic crystals tend to shatter. The crystal structure of ionic compounds is strong and rigid. It takes a lot of energy to break all those strong ionic bonds. As a result, ionic compounds are solids with high melting and boiling points (see **Table** below). You can learn more about the properties of ionic compounds by watching the video at this URL:
http://www.youtube.com/watch?v=buWrSgs_ZHk&feature=related (3:34).

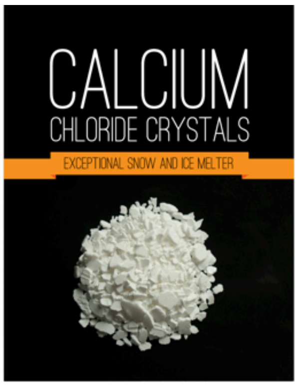
Compare the melting and boiling points of these ionic compounds with those of water (0°C and 100°C), which is not an ionic compound.

Ionic Compound	Melting Point (°C)	Boiling Point (°C)
Sodium chloride ($NaCl$)	801	1413
Calcium chloride ($CaCl_2$)	772	1935
Barium oxide (BaO)	1923	2000
Iron bromide ($FeBr_3$)	684	934

Solid ionic compounds are poor conductors of electricity. The strong bonds between ions lock them into place in the crystal. However, in the liquid state, ionic compounds are good conductors of electricity. Most ionic compounds dissolve easily in water. When they dissolve, they separate into individual ions. The ions can move freely, so they are good conductors of electricity. Dissolved ionic compounds are called electrolytes.

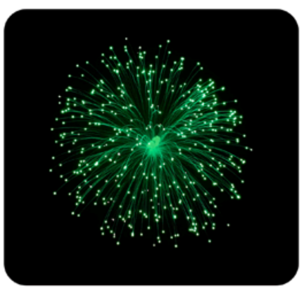
Uses of Ionic Compounds

Ionic compounds have many uses. Some are shown in **Figure** below. Many ionic compounds are used in industry. The human body also needs several ions for good health. Having low levels of the ions can endanger important functions such as heartbeat. Solutions of ionic compounds can be used to restore the ions.



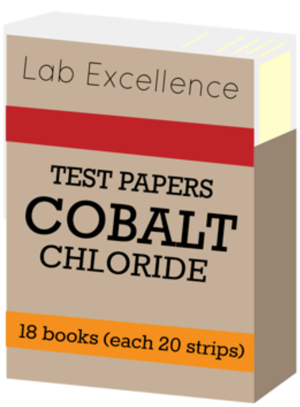
Calcium chloride crystals are used to melt ice and snow. The crystals lower the freezing point of water.

Potassium iodide tablets are given to people who are exposed to high levels of radiation. They protect the thyroid gland from radiation.



Barium chloride is used to make fireworks. It produces green-colored explosions.

Electrolyte solutions are given to children who have lost ions. This can happen with vomiting or diarrhea.



Cobalt chloride test papers are used to detect moisture. They change color when they absorb water.

Lithium iodide is used in batteries. It's an excellent conductor of electricity.

Have you ever used any of these ionic compounds?

Writing Basic Ionic Formulas

In writing formulas for binary ionic compounds (binary refers to two elements, *not* two single atoms), the cation is always written first. Chemists use subscripts following the symbol of each element to indicate the number of that element present in the formula. For example, the formula $\mathrm{Na_2O}$ indicates that the compound contains two atoms of sodium for every one atom of oxygen. When the subscript for an element is 1, the subscript is omitted. The number of atoms of an element with no indicated subscript is always read as 1. When an ionic compound forms, the number of electrons given off by the cations must be exactly the same as the number of electrons taken on by the anions. Therefore, if calcium, which gives off two electrons, is to be combined with fluorine, which takes on one electron, then one calcium atom must combine with two fluorine atoms. The formula would be $\mathrm{CaF_2}$.

Suppose we wish to write the formula for the compound that forms between aluminum and chlorine. To write the formula, we must first determine the charges of the ions that would be formed.

$$\begin{array}{ll} 3+ & 1- \\ \mathrm{Al} & \mathrm{Cl} \end{array}$$

Then, we determine the simplest whole numbers with which to multiply these charges so they will balance (add to zero). In this case, we would multiply the 3+ by 1 and the 1– by 3.

You should note that we could multiply the 3+ by 2 and the 1– by 6 to get 6+ and 6– respectively. These values will also balance, but this is *not* acceptable because empirical formulas, by definition, must have the lowest whole number multipliers. Once we have the *lowest* whole number multipliers, those multipliers become the subscripts for the symbols. The formula for this compound would be $\mathrm{AlCl_3}$.

Here's the process for writing the formula for the compound formed between aluminum and sulfur.

Al^{3+} S^{2-}

Therefore, the formula for this compound would be $\mathrm{Al_2S_3}$.

Another method used to write formulas is called the criss-cross method. It is a quick method, but it often produces errors if the user doesn't pay attention to the results. The example below demonstrates the criss-cross method for writing the formula of a compound formed from aluminum and oxygen. In the criss-cross method, the oxidation numbers are placed over the symbols for the elements just as before.

$\overset{3+}{\mathrm{Al}}$ $\qquad\qquad$ $\overset{2-}{\mathrm{O}}$

In this method, the oxidation numbers are then criss-crossed and used as the subscripts for the other atom (ignoring sign).

$$Al^{3+} \quad O^{2-}$$

This produces the correct formula Al_2O_3 for the compound. Here's an example of a criss-cross error:

$$Pb^{4+} \quad O^{2-}$$

$$Pb^{4+} \quad O^{2-}$$

$$Pb_2O_4$$

$$PbO_2$$

If you used the original method of finding the lowest multipliers to balance the charges, you would get the correct formula PbO_2, but the criss-cross method produces the incorrect formula Pb_2O_4. If you use the criss-cross method to generate an ionic formula, it is essential that you check to make sure that the subscripts correspond to the lowest whole number ratio of the atoms involved.

Polyatomic Ions

Polyatomic ions are those that are made up of more than one atom; they are group of bonded atoms with a charge. Polyatomic ions require additional consideration when you write formulas involving them. Here is a short list of some common polyatomic ions:

- Ammonium ion, NH_4^+
- Acetate ion, $C_2H_3O_2^-$
- Carbonate ion, CO_3^{2-}
- Chromate ion, CrO_4^{2-}
- Dichromate ion, $Cr_2O_7^{2-}$
- Hydroxide ion, OH^-
- Nitrate ion, NO_3^-
- Phosphate ion, PO_4^{3-}
- Sulfate ion, SO_4^{2-}
- Sulfite ion, SO_3^{2-}

Suppose we are asked to write the formula for the compound that would form between calcium and the nitrate ion. We begin by putting the charges above the symbols just as before.

The multipliers needed to balance these ions are 1 for calcium and 2 for nitrate. We wish to write a formula that tells our readers that there are two nitrate ions in the formula for every calcium ion. When we put the subscript 2 beside the nitrate ion in the same fashion as before, we get something strange – $CaNO_{32}$. With this formula, we are indicating 32 oxygen atoms, which is wrong. The solution to this problem is to put parentheses around the nitrate ion before the subscript is added. Therefore, the correct formula is $Ca(NO_3)_2$. Similarly, calcium phosphate would be $Ca_3(PO_4)_2$. If a polyatomic ion does not need a subscript other than an omitted 1, then the parentheses are not needed. Although including these unnecessary parentheses does not change the meaning of the formula, it may cause the reader to wonder whether a subscript was left off by mistake. Try to avoid using parentheses when they are not needed.

Example:

Al^{3+} $C_2H_3O_2^{1-}$

Write the formula for the compound that will form from aluminum and acetate.

The charge on an aluminum ion is $+3$, and the charge on an acetate ion is -1. Therefore, three acetate ions are required to combine with one aluminum ion. This is also apparent by the criss-cross method. However, we cannot place a subscript of 3 beside the oxygen subscript of 2 without inserting parentheses first. Therefore, the formula will be $Al(C_2H_3O_2)_3$.

Example:
Write the formula for the compound that will form from ammonium and phosphate.

NH_4^{1+} PO_4^{3-}

The charge on an ammonium ion is $+1$ and the charge on a phosphate ion is -3. Therefore, three ammonium ions are required to combine with one phosphate ion. The criss-cross procedure will place a subscript of 3 next to the subscript 4. This can only be carried out if the ammonium ion is first placed in parentheses. Therefore, the proper formula is $(NH_4)_3PO_4$.

Example:
Write the formula for the compound that will form from aluminum and phosphate.

Al^{3+} $PO_4{}^{3-}$

Since the charge on an aluminum ion is $+3$ and the charge on a phosphate ion is -3, these ions will combine in a one-to-one ratio. In this case, the criss-cross method would produce an incorrect answer. Since it is not necessary to write the subscripts of 1, no parentheses are needed in this formula. Since parentheses are not needed, it is generally considered incorrect to use them. The correct formula is $AlPO_4$.

More Examples:

Magnesium hydroxide...... $Mg(OH)_2$

Sodium carbonate............ Na_2CO_3

Barium acetate............... $Ba(C_2H_3O_2)_2$

Ammonium dichromate......... $(NH_4)_2Cr_2O_7$

Variable Charge Metals

Metals with variable charges may form multiple different compounds with the same nonmetal. Iron, for example, may react with oxygen to form either FeO or Fe_2O_3. These are very different compounds with different properties. When we name these compounds, it is absolutely vital that we clearly distinguish between them. They are both iron oxides, but in FeO iron has an charge of $+2$, while in Fe_2O_3, it has a charge of $+3$. The rule for naming these compounds is to write the charge of the metal after the name. The charge is written using Roman numerals and is placed in parentheses. For these two examples, the compounds would be named iron(II) oxide and iron(III) oxide. When you see that the compound involves a metal with multiple charges, you must determine the charge of the metal from the formula and indicate it using Roman numerals.

In general, main group metal ions have only one common oxidation state, whereas most of the transition metals have more than one. However, there are plenty of exceptions to this guideline. Main group metals that can have more than one oxidation state include tin (Sn^{2+} or Sn^{4+}) and lead (Pb^{2+} or Pb^{4+}). Transition metals with only one common oxidation state include silver (Ag^+), zinc (Zn^{2+}), and cadmium (Cd^{2+}). These should probably be memorized, but when in doubt, include the Roman numerals for transition metals. Do not do this for main group metals that do not have more than one oxidation state. Referring to AgCl as silver(I) chloride is redundant and may be considered wrong. However, copper chloride is definitely incorrect, because it could refer to either $CuCl$ or $CuCl_2$.

Other than the use of Roman numerals to indicate charge, naming these ionic compounds is no different than what we have already seen. For example, consider the formula $CuSO_4$. We know that the sulfate anion has a charge of -2. Therefore, for this to be a neutral compound, copper must have a charge of +2. The name of this compound is copper(II) sulfate.

How about SnS_2? Tin is a variable charge metal. We need a Roman numeral in the name of this compound. The oxidation number of sulfur is -2. Two sulfide ions were necessary to combine with one tin ion. Therefore, the charge of the tin must be $+4$, and the name of this compound is tin(IV) sulfide.

Examples:

PbO . . . . lead(II) oxide

FeI_2 . . . . iron(II) iodide

$Fe_2(SO_4)_3$ . . . . . . iron(III) sulfate

$AuCl_3$ . . . . gold(III) chloride

CuO . . . . copper(II) oxide

PbS_2 . . . . lead(IV) sulfide

LESSON SUMMARY

- An ionic bond is the force of attraction that holds together oppositely charged ions. It forms when atoms of a metal transfer electrons to atoms of a nonmetal. When this happens, the atoms become oppositely charged ions.
- Ionic compounds form crystals instead of molecules. Ionic bonds are strong and the crystals are rigid. As a result, ionic compounds are brittle solids with high melting and boiling points. In the liquid state or dissolved in water, ionic compounds are good conductors of electricity.
- Cations have the same name as their parent atom.
- Monatomic anions are named by replacing the end of the parent atom's name with "-ide."
- The names of polyatomic ions do not change.
- Ionic compounds are named by writing the name of the cation followed by the name of the anion.
- When naming compounds that include a metal with more than one charge, the charge of the metal ion is indicated with Roman numerals in parentheses between the cation and anion

LESSON REVIEW QUESTIONS

1. What is an ionic bond?
2. Outline the role of energy in the formation of an ionic bond.
3. List properties of ionic compounds.
4. Name the following compounds.

a. CaF_2

b. $(NH_4)_2CrO_4$

c. K_2CO_3

d. $NaCl$

e. $CuSO_4$

f. $Ca(NO_3)_2$

g. $Mg(OH)_2$

5. Write the formulas from the names of the following compounds.

h. Sodium carbonate

i. Calcium hydroxide

j. Iron(III) nitrate

k. Magnesium oxide

l. Aluminum sulfide

6. Create a model to represent the ionic bonds in a crystal of the salt lithium iodide (LiI).
7. A mystery compound is a liquid with a boiling point of 50°C. Is it likely to be an ionic compound? Why or why not?
8. Explain why ionic bonds form only between atoms of metals and nonmetals.

Metallic Bonding

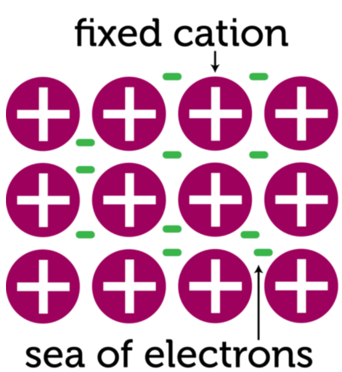
- Vocabulary:
- Metallic bonds
- Malleable
- Conductor

Lesson Objectives:
- Describe how metallic bonds are formed
- List the properties of metallic compounds

The thick, rigid trunk of the oak tree on the left might crack and break in a strong wind. The slim, flexible trunk of the willow tree on the right might bend without breaking. In one way, metals are like willow trees. They can bend without breaking. That's because metals form special bonds called metallic bonds.

What Are Metallic Bonds?

Metallic bonds are forces of attraction between positive metal ions and the valence electrons that are constantly moving around them (see the **Figure** below). The valence electrons include their own and those of other, nearby ions of the same metal. The valence electrons of metals move freely in this way because metals have relatively low electronegativity, or attraction to electrons. The positive metal ions form a lattice-like structure held together by all the metallic bonds.



Metallic bonds.

Q: Why do metallic bonds form only in elements that are metals? Why don't similar bonds form in elements that are nonmetals?

A: Metal atoms readily give up valence electrons and become positive ions whenever they form bonds. When nonmetals bond together, the atoms share valence electrons and do not become ions. For example, when oxygen atoms bond together they form oxygen molecules in which two oxygen atoms share two pairs of valence electrons equally, so neither atom becomes charged.

Metallic Bonds and the Properties of Metals

The valence electrons surrounding metal ions are constantly moving. This makes metals good **conductors** of electricity. The lattice-like structure of metal ions is strong but quite flexible. This allows metals to bend without breaking. Metals are both ductile (can be shaped into wires) and **malleable** (can be shaped into thin sheets). You can learn more about metallic bonds and the properties of metals at this URL: http://www.youtube.com/watch?v=ap5pHBWwpu4

Q: Look at the metalworker in the Figure below. He's hammering a piece of hot iron in order to shape it. Why doesn't the iron crack when he hits it?

A: The iron ions can move within the "sea" of electrons around them. They can shift a little closer together or farther apart without breaking the metallic bonds between them. Therefore, the metal can bend rather than crack when the hammer hits it.

Metal worker shaping iron metal.

SUMMARY

- Metallic bonds are the force of attraction between positive metal ions and the valence electrons that are constantly moving around them. The ions form a lattice-like structure held together by the metallic bonds.
- Metallic bonds explain why metals can conduct electricity and bend without breaking

PRACTICE

Watch the video about metallic bonds at the following URL, and then answer the questions below.
http://www.youtube.com/watch?v=ap5pHBWwpu4

1. What is electricity? Why can metals conduct electricity?
2. What can metals conduct besides electricity?
3. How could you use an empty pop can to demonstrate that metals can bend without breaking?

REVIEW

4. What are metallic bonds?
5. How do metallic bonds relate to the properties of metals?
6. The iron in the metal working picture above is red hot. Infer why the metalworker heats the iron when he shapes it

COVALENT BONDING

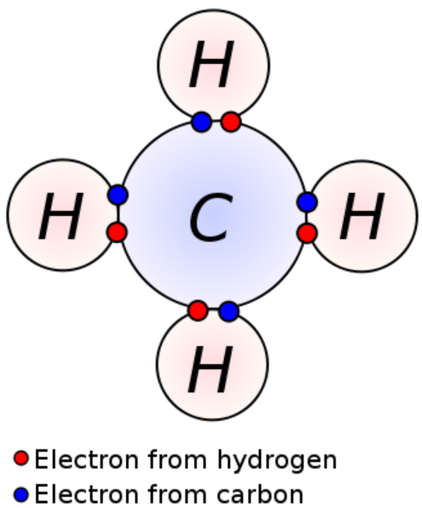
Lesson Objectives

- explain what covalent bonds are.
- explain why covalent bonds are formed.
- compare covalent bonds with ionic and metallic bonds in terms of how their definitions and how they are formed.
- Name and write the formulas for covalent compounds
- draw Lewis structures for simple covalent molecules.

Introduction

In a covalent bond, an atom shares one or more electrons with another atom. Covalent bonds occur between nonmetals.

In the picture of methane (CH_4) below, carbon shares a single electron from each of the the four hydrogens. Covalent bonding is prevalent in organic compounds. In fact, your body is held together by electrons shared by carbons and hydrogens! Covalent bonds are also very strong, meaning it takes a lot of energy to break them apart. Covalent compounds are not good conductors and they typically have low melting and boiling points.



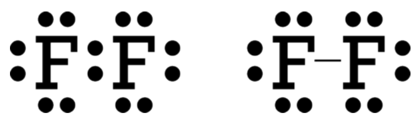
Methane is formed when four hydrogens and one carbon covalently bond.

- Vocabulary
- covalent bond
- polar covalent bond
- electronegativity

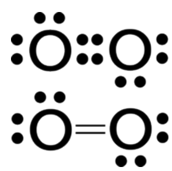
Lewis Dot Structures

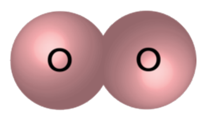
Lewis dot structures, first developed by G.N. Lewis, are a shorthand way of drawing the arrangement of atoms, bonds, and valence electrons in a molecule. When we draw molecules, the diagrams are known as Lewis dot formulas, Lewis structures, or Lewis formulas. The Lewis structures of a molecule show how the valence electrons are arranged among the atoms of the molecule.

In a Lewis structure, each valence electron is represented by a dot, and bonds are shown by placing electrons in between the symbols for the two bonded atoms. Often, a pair of bonding electrons is further abbreviated by a dash. For example, we can represent the covalent bond in the F_2 molecule by either of the Lewis structures shown below.

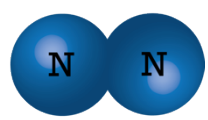


Double bonds (4 electrons shared between two atoms) can be represented either with 4 dots or 2 dashes. The Lewis structure for an oxygen molecule (O_2) is shown below.

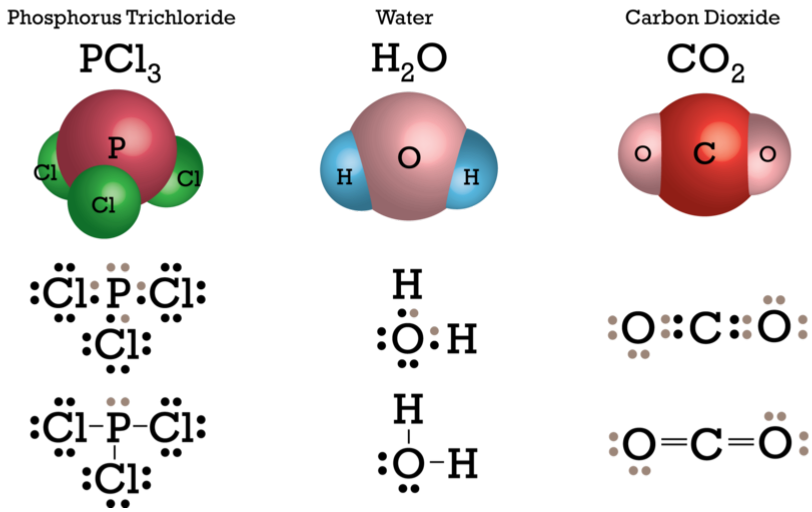




Similarly, triple bonds can be written as 6 dots or 3 dashes. An example of a molecule with triple bonds is the nitrogen molecule, N_2. The Lewis structure for a nitrogen molecule can be represented by either of the two ways shown below. It is important to keep in mind that a dash always represents a pair of electrons. .



Several other examples of representing covalent bonds are shown in the figure below.



Drawing Lewis Structures

The rules outlined below for writing Lewis structures are based on observations of thousands of molecules (note that there will be some exceptions to the rules). We learned earlier that atoms are generally most stable with 8 valence electrons (the octet rule). The major exception is hydrogen, which requires only 2 valence electrons to have a complete valence shell (sometimes called the duet rule). Lewis structures allow one to quickly assess whether each atom has a complete octet (or duet) of electrons.

Rules for Writing Lewis Structures:

1. Decide which atoms are bounded.
    7. Count all the valence electrons of all the atoms.
    8. Place two electrons between each pair of bounded atoms.
    9. Complete all the octets (or duets) of the atoms attached to the central atom.
    10. Place any remaining electrons on the central atom.
    11. If the central atom does not have an octet, look for places to form double or triple bonds.

Example:

Write the Lewis structure for water, H_2O.

Step 1: Decide which atoms are bounded.
Begin by assuming the hydrogen atoms are bounded to the oxygen atom. In other words, assume the oxygen atom is the central atom.

H – O – H

Step 2: Count all the valence electrons of all the atoms.

The oxygen atom has 6 valence electrons, and each hydrogen has 1. The total number of valence electrons is 8.

Step 3: Place two electrons between each pair of bounded atoms.

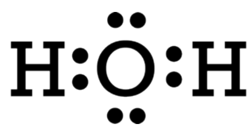
H : O : H

Step 4: Complete all the octets or duets of the atoms attached to the central atom.

The hydrogen atoms are attached to the central atom and require a duet of electrons. Those duets are already present.

Step 5: Place any remaining electrons on the central atom.
The total number of valence electrons is 8, and we have already used 4 of them. The other 4 will fit around the central oxygen atom.

Is this structure correct?



Are the total number of valence electrons correct? Yes
Does each atom have the appropriate duet or octet of electrons? Yes
The structure, then, is correct.

Example:
Write the Lewis structure for carbon dioxide, CO_2.

Step 1: Decide which atoms are bounded.
Begin by assuming the carbon is the central atom and that both oxygen atoms are attached to the carbon.

Step 2: Count all the valence electrons of all the atoms.
The oxygen atoms each have 6 valence electrons and the carbon atom has 4. The total number of valence electrons is 16.

Step 3: Place two electrons between each pair of bounded atoms.

O : C : O

Step 4: Complete all the octets or duets of the atoms attached to the central atom.

Step 5: Place any remaining electrons on the central atom.

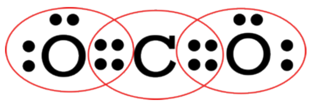
We have used all 16 of the valence electrons so there are no more to place around the central carbon atom.

Is this structure correct?

Is the total number of valence electrons correct? Yes
Does each atom have the appropriate duet or octet of electrons? NO
Each oxygen has the proper octet of electrons, but the carbon atom only has 4 electrons. Therefore, this structure is not correct.

Step 6: If the central atom does not have an octet, look for places to form double or triple bonds.

Double bonds can be formed between carbon and each oxygen atom.



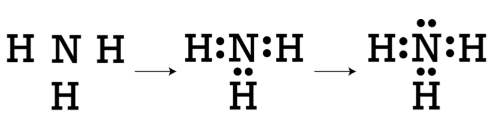
Notice this time that each atom is surrounded by 8 electrons.

Example:
Write the Lewis structure for ammonia, NH_3.

Step 1: Decide which atoms are bounded.
The most likely bonding for this molecule is to have nitrogen as the central atom and each hydrogen bounded to the nitrogen. Therefore, we can start by putting nitrogen in the center and placing the three hydrogen atoms around it.

Step 2: Count all the valence electrons of all the atoms.
The nitrogen atom has five valence electrons, and each hydrogen atom has one. The total number of valence electrons is 8.

Step 3: Place two electrons between each pair of bounded atoms, as seen in the middle drawing of the figure below.



Step 4: Complete all the octets or duets of the atoms attached to the central atom.
In this case, all the non-central atoms are hydrogen, and they already have a duet of electrons.

Step 5: Place any remaining electrons on the central atom.

There are still two electrons left, so they would complete the octet for nitrogen. If the central atom, at this point, does not have an octet of electrons, we would look for places to create a double or triple bond, but that is not the case here. The final drawing on the right in the figure above is the Lewis structure for ammonia.

LESSON SUMMARY

- Covalent bonds are formed when nonmetals share electrons.
- A polar covalent bond is when electrons are unequally shared.
- Naming molecular compounds is based on greek prefixes.
- Lewis dot structures are used to show the structure of molecules.
- There are differences between the three types of bonds.

REVIEW QUESTIONS

1. What is a covalent bond and how does a covalent bond differ from a polar covalent bond?
2. Name the compound CO.
3. Name the compound PCl_3.
4. Name the compound PCl_5.
5. Write the formula for the compound sulfur trioxide.
6. Write the formula for the compound dinitrogen tetrafluoride.
7. Write the formula for the compound oxygen difluoride.
8. Draw the lewis dot structure for the following:
    1. H_2O
    2. NH_3
    3. CH_4
    4. O_2
9. How is a covalent bond different from an ionic bond?
10. How is a covalent bond different from a metallic bond?

STANDARD 3, OBJECTIVE 3: RELATE THE PROPERTIES OF SIMPLE COMPOUNDS TO THE TYPE OF BONDING, SHAPE OF MOLECULES, AND INTERMOLECULAR FORCES.

MOLECULAR GEOMETRY

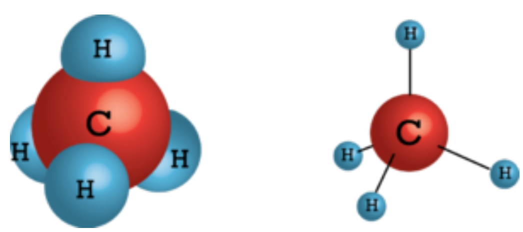
Objectives

- Predict the shape of simple molecules and their polarity from Lewis dot structures.
- Explain the meaning of the acronym VSEPR and state the concept on which it is based.

Introduction

Although a convenient way for chemists to look at covalent compounds is to draw Lewis structures, which shows the location of all of the valence electrons in a compound. Although these are very useful for understanding how atoms are arranged and bonded, they are limited in their ability to accurately represent what shape molecules are. Lewis structures are drawn on flat paper as two dimensional drawings. However, molecules are really three dimensional. In this section you will learn to predict the 3d shape of many molecules given their Lewis structure.

Many accurate methods now exist for determining molecular structure, the three-dimensional arrangement of the atoms in a molecule. These methods must be used if precise information about structure is needed. However, it is often useful to be able to predict the approximate molecular structure of a molecule. A simple model that allows us to do this is called the **valence shell electron pair repulsion (VSEPR) theory.** This model is useful in predicting the geometries of molecules formed in the covalent bonding of non-metals. The main postulate of this theory is that in order to minimize electron-pair repulsion. In other words, the electron pairs around the central atom in a molecule will get as far away from each other as possible.



A central atoms bonded to four other atoms with no lone pairs has a tetrahedral shape.

Predicting the Shape of Molecules

Consider, methane, commonly known as natural gas. In this molecule, carbon has four valence electrons and each hydrogen adds one more so the central atom in methane has four pairs of electrons in its valence shell. The 3d shape of this molecule is dictated by the repulsion of the electrons. Those four pairs of electrons get as far away from each other as possible which forms a shape called **tetrahedral.** In the tetrahedral shape, the bond angle between any two hydrogen atoms is 109.5°.

What if we look at ammonia instead, NH_3? A molecule of ammonia has a nitrogen atom in the middle with three bonds to the hydrogen atoms plus one lone pair of electrons. That means there are four total pairs of electrons around the central atom, and the electrons will still be close to 109.5° apart from each other. However, when discussing the overall shape of the molecule, we only take into account the location of the atoms. When a central atom is bonded to three atoms and has one lone pair of electrons, the overall shape is **trigonal pyramidal**.

We have a similar problem in the case of a molecule such as water, H_2O. In water, the oxygen atom in the middle is bonded to the two hydrogen atoms with two lone pairs. Once again, we only consider the location of atoms when we discuss shape. When a molecule has a central atom bonded to two other atoms with two lone pair of electrons, the overall shape is **bent**.

Electronic Geometry: Tetrahedral
All electron pairs shared.
Molecular Geometry: Tetrahedral

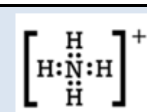
Electronic Geometry: Tetrahedral
One unshared pair of electrons.
Molecular Geometry: Pyramidal

Electronic Geometry: Tetrahedral
Two unshared pairs of electrons.
Molecular Geometry: Angular (bent)

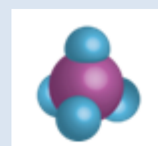
As you can probably imagine, there are different combinations of bonds making different shapes of molecules. Some of the possible shapes are listed in the table. However, it is important to note that some molecules obtain geometries that are not included here.

Summary of Molecular Geometry		
# of atoms bonded to central atom	# of unshared pairs around central atom	Molecular Geometry
2	0	Linear
3	0	Trigonal Planar
2	1	Bent
3	*1*	**Trigonal pyramidal*
2	*2*	**Bent*
4	*0*	**Tetrahedral*

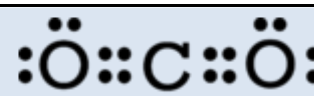
Example: Determine the shape of ammonium, NH_4^+, given by the following Lewis structure:



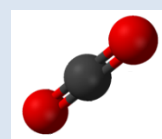
Solution: To answer this question, you need to count the number of atoms around the central atom and the number of unshared pairs. In this example, there are four atoms bonded to the N with zero unshared pairs of electrons. The shape must be tetrahedral.



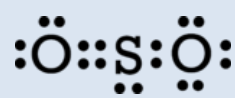
Example: Determine the shape of carbon dioxide, CO_2, given by the following Lewis structure:



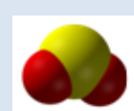
Solution: To answer this question, you need to count the number of atoms around the central atom and the number of unshared pairs. In this example, there are two atoms bonded to the C with zero unshared pairs of electrons. The shape must be linear, according to the table.



Example: Determine the shape of carbon dioxide, SO_2, given by the following Lewis structure:



Solution: To answer this question, you need to count the number of atoms around the central atom and the number of unshared pairs. In this example, there are two atoms bonded to the S with one unshared pair of electrons. The shape must be bent, according to the table.



VOCABULARY

- VSEPR model: A model whose main postulate is that the structure around a given atom in a molecule is determined by minimizing electron-pair repulsion.
- Molecular geometry: The specific three-dimensional arrangement of atoms in molecules.

Further Readings / Supplemental Links

- http://www.up.ac.za/academic/chem/mol_geom/mol_geometry.htm
- http://en.wikipedia.org/wiki/Molecular_geometry
- An animation showing the molecular shapes that are generated by sharing various numbers of electron pairs around the central atom (includes shapes when some pairs of electrons are non-shared pairs). The link must be copied and pasted into your browser to go directly to the animation.
- http://www.classzone.com/cz/books/woc_07/resources/htmls/ani_chem/chem_flash/popup.html?layer=act&src=qtiwf_act065.1.xml

REVIEW QUESTIONS

Predict the 3d shape each of the following molecules will have:

1) CH_3Cl
2) Silicon tetrafluoride,
3) $CHCl_3$
4) Water, H_2O
5) Ammonia, NH_3
6) Methane, CH_4

POLARITY AND HYDROGEN BONDING

Objectives

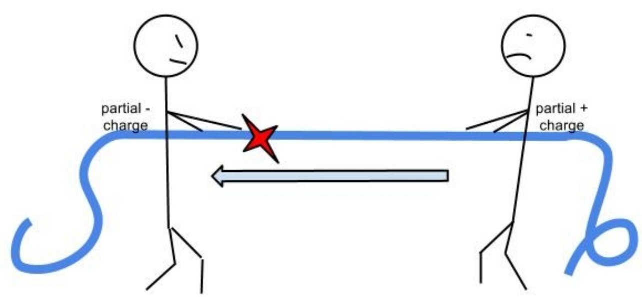
- Explain how polar compounds differ from nonpolar compounds
- Determine if a molecule is polar or nonpolar
- Identify whether or not a molecule can exhibit hydrogen bonding
- List important phenomena which are a result of hydrogen bonding
- Given a pair of compounds, predict which would have a higher melting or boiling point

Introduction

The ability of an atom in a molecule to attract shared electrons is called **electronegativity.** When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur) and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

Polar Covalent Bonds

So far, we have discussed two extreme types of bonds. One case is when two identical atoms bond. They have exactly the same electronegativities, thus the two bonded atoms pull exactly equally on the shared electrons. The shared electrons will be shared exactly equally by the two atoms.



A polar covalent bond is similar to a tug-of-war in which one atom pulls more on the electrons and gains a partial negative charge. The weaker (less electronegative atom) has a partial positive charge.

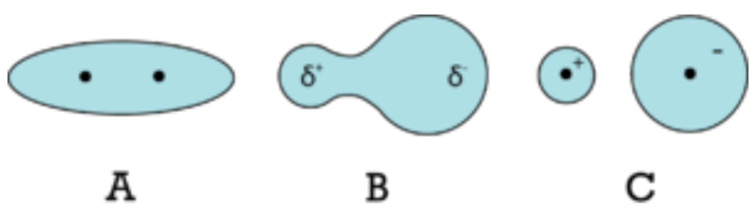
CC – Tracy Poulsen

The other case is when the bonded atoms have a very large difference in their electronegativities. In this case, the more electronegative atom will take the electrons completely away from the other atom and an ionic bond forms.

What about the molecules whose electronegativities are not the same but the difference is not big enough to form an ionic bond? For these molecules, the electrons remain shared by the two atoms but they are not shared equally. The shared electrons are pulled closer to the more electronegative atom. This results in an uneven distribution of electrons over the molecule and causes slight charges on opposite ends of the molecule. The negative electrons are around the more electronegative atom more of the time creating a partial negative side. The other side has a resulting partial positive charge. These charges are not full +1 and -1 charges, they are fractions of charges. For small fractions of charges, we use the symbols **δ+** and **δ–**. These molecules have slight opposite charges on opposite ends of the molecule and said to have a dipole or are called **polar molecules**.

A polar molecule has partially positive and partially negative charges on opposite sides of the molecule.

When atoms combine, there are three possible types of bonds that they can form. In the figure, molecule A represents a covalent bond that would be formed between identical atoms. The electrons would be evenly shared with no partial charges forming. This molecule is **nonpolar**. Molecule B is a polar covalent bond formed between atoms whose electronegativities are not the same but whose electronegativity difference is less than 1.7, making this molecule **polar**. Molecule C is an ionic bond formed between atoms whose electronegativity difference is greater than 1.7.



A) A nonpolar covalent bond in which two identical atoms are sharing electrons

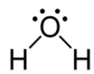
B) a polar covalent bond in which the more electronegative atom pulls the electrons more toward itself (forming partial negative and positive sides)

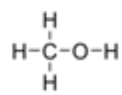
C) an ionic bond in which an extremely electronegative atom is bonded to a very weakly electronegative atom.

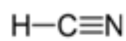
Polar molecules can be attracted to each other due attraction between opposite charges. Polarity underlies a number of physical properties including surface tension, solubility, and melting- and boiling-points. The more attracted molecules are to other molecules, the higher the melting point, boiling point, and surface tension. We will discuss in more detail later how polarity can affect how compounds dissolve and their solubility.

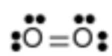
In order to determine if a molecule is polar or nonpolar, it is frequently useful to look a Lewis structures. Nonpolar compounds will be symmetric, meaning all of the sides around the central atom are identical—bonded to the same element with no unshared pairs of electrons. Polar molecules are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded.

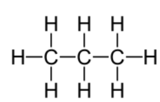
Example: Label each of the following as polar or nonpolar.

1) Water, H_2O: 

2) methanol, CH_3OH: 

3) hydrogen cyanide, HCN: 

4) Oxygen, O_2: 

5) Propane, C_3H_8: 

Solution:

1) Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
2) Methanol is polar. This is not a symmetric molecule. The –OH side is different from the other 3 –H sides.
3) Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
4) Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
5) Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

While molecules can be described as "polar covalent", "non-polar covalent", or "ionic", it must be noted that this is often a relative term, with one molecule simply being *more polar* or *less polar* than another. However, the following properties are typical of such molecules. Polar molecules tend to:

- have higher melting points than nonpolar molecules
- have higher boiling points than nonpolar molecules
- be more soluble in water (dissolve better) than nonpolar molecules
- have lower vapor pressures than nonpolar molecules

Hydrogen Bonding:

When a hydrogen atom is bonded to a very electronegative atom, including fluorine, oxygen, or nitrogen, a very polar bond is formed. The electronegative atom obtains a negative partial charge and the hydrogen obtains a positive partial charge. These partial charges are similar to what happens in every polar molecule. However, because of the big difference in electronegativities between these two atoms and the amount of positive charge exposed by the hydrogen, the dipole is much more dramatic. These molecules will be attracted to other molecules which also have partial charges. This attraction for other molecules which also have a hydrogen bonded to a fluorine, nitrogen, or oxygen atom is called a **hydrogen bond**.

Hydrogen bonds in water

The most important, most common, and perhaps simplest example of a hydrogen bond is found between water molecules. This interaction between neighboring water molecules is responsible for many of the important properties of water.

Hydrogen bonding strongly affects the crystal structure of ice, helping to create an open hexagonal lattice. The density of ice is less than water at the same temperature; thus, the solid phase of water floats on the liquid, unlike most other substances in which the solid form would sink in the liquid form.

Water also has a high boiling point (100°C) compared to the other compounds of similar size without hydrogen bonds. Because of the difficulty of breaking these bonds, water has a very high boiling point, melting point, and viscosity compared to otherwise similar liquids not conjoined by hydrogen bonds.

Water is unique because its oxygen atom has two lone pairs and two hydrogen atoms, meaning that the total number of bonds of a water molecule is up to four. For example, hydrogen fluoride—which has three lone pairs on the F atom but only one H atom—can form only two bonds; (ammonia has the opposite problem: three hydrogen atoms but only one lone pair).

H-F...H-F...H-F.

Have you ever experienced a belly flop? This is also due to the hydrogen bonding between water molecules, causing surface tension. On the surface of water, water molecules are even more attracted to their neighbors than in the rest of the water. This attraction makes it difficult to break through, causing belly flops. It also explains why water striders are able to stay on top of water and why water droplets form on leaves or as they drip out of your faucet.

A. Water beading on a leaf. http://en.wikipedia.org/wiki/File:Dew_2.jpg

B. Water dripping from a tap. http://en.wikipedia.org/wiki/File:Water_drop_animation_enhanced_small.gif

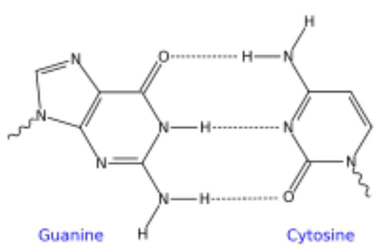
C. Water striders stay atop the liquid due to surface tension. http://en.wikipedia.org/wiki/File:WaterstriderEnWiki.jpg

Hydrogen bonds in DNA and proteins

Hydrogen bonding also plays an important role in determining the three-dimensional structures adopted by proteins and nucleic bases, as found in your DNA. In these large molecules, bonding between parts of the same macromolecule cause it to fold into a specific shape, which helps determine the molecule's physiological or biochemical role. The double helical structure of DNA, for example, is due largely to hydrogen bonding between the base pairs, which link one complementary strand to the other and enable replication. It also plays an important role in the structure of polymers, both synthetic and natural, such as nylon and many plastics.

As a result of the strong attraction between molecules that occurs in a hydrogen bond, the following properties can be summarized. Molecules with hydrogen bonding tend to:

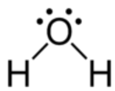
- have higher melting points than polar molecules
- have higher boiling points than polar molecules
- be more soluble in water (dissolve better) than polar molecules



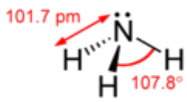
Hydrogen bonding between guanine (G) and cytosine(C), one of two types of base pairs in DNA. The hydrogen bonds are shown by dotted lines.

Example: Label each of the following as polar or nonpolar and indicate which have hydrogen bonding.

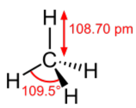
a) H_2O



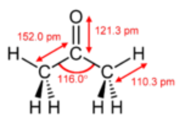
b) Ammonia



c) CH_4



d) acetone, CH_3COCH_3



Solution:

a) This molecule is polar (the unshared pairs of electrons make a polar assymetric shape), and hydrogen bonding (hydrogen is bonded to N, O, or F).

b) This molecule is polar (the unshared pairs of electrons make a polar assymetric shape), and hydrogen bonding (hydrogen is bonded to N, O, or F).

c) This molecule is nonpolar (the molecule is symmetric with H's bonded to all four sides of the central atom), and does not have hydrogen bonding (hydrogen is not bonded to N, O, or F).

d) This molecule is polar (the O is not the same as the CH3 bonded to the central atom) and does not have hydrogen bonding (hydrogen is bonded DIRECTLY to N, O, or F).

Example: For each pair of molecules, indicate which you would expect to have a higher melting point. Explain why. Also, refer to the Lewis structures given to you in the previous example.

a) H_2O vs. acetone
b) CH_4 vs. acetone

Solution:
a) H2O (polar, hydrogen bonding) vs. acetone (polar, no hydrogen bonding). H2O will have a higher melting point because compounds with hydrogen bonding tend to have higher melting points than polar compounds.

b) CH4 (nonpolar, no hydrogen bonding) vs. acetone (polar, no hydrogen bonding). Acetone will have a higher melting point because polar molecules tend to have higher melting points than nonpolar molecules.

Lesson Summary

- Covalent bonds between atoms that are not identical will produce polar bonds.
- Molecules with polar bonds and non-symmetrical shapes will have a dipole.
- Hydrogen bonding is a special interaction felt between molecules, which is a stronger interaction than polar-polar attraction.
- Hydrogen bonding occurs between molecules in which a hydrogen atom is bonded to a very electronegative fluorine, oxygen, or nitrogen atom.
- Compounds with hydrogen bonding tend to have higher melting points, higher boiling points, and greater surface tenstion.
- The unique properties of water are a result of hydrogen bonding
- Hydrogen bonding plays roles in many compounds including DNA, proteins, and polymers.

Vocabulary

- Electronegativity: The tendency of an atom in a molecule to attract shared electrons to itself.
- Polar covalent bond: A covalent bond in which the electrons are not shared equally because one atom attracts them more strongly that the other.

Further Reading / Supplemental Links

http://learner.org/resources/series61.html; The learner.org website allows users to view streaming videos of the Annenberg series of chemistry videos. You are required to register before you can watch the videos but there is no charge. The website has one video that relates to this lesson called Molecular Architecture.

Vision Learning: Water Properties & Behaviors
http://visionlearning.com/library/module_viewer.php?mid=57&l=&c3=

REVIEW QUESTIONS

1. Explain the differences among a nonpolar covalent bond, a polar covalent bond, and an ionic bond.
2. Predict which of the following bonds will be more polar and explain why; P-Cl or S-Cl.
3. What does it mean for a molecule to be "polar"?
4. Which three elements, when bonded with hydrogen, are capable of forming hydrogen bonds?
5. Molecules that are polar exhibit dipole-dipole interaction. What's the difference between dipole-dipole interactions and hydrogen bonding? Which interaction is stronger?
6. Define hydrogen bonding. Sketch a picture of several water molecules and how they interact.
7. After drawing a Lewis structure, indicate whether each is polar or nonpolar. Then indicate whether or not that compound exhibits hydrogen bonding.
   a) CH_3Cl
   b) Silicon tetrafluoride, SiF_4
   c) Water, H_2O
   d) Ammonia, NH_3
   e) Methane, CH_4

CHAPTER 4

CHEMICAL REACTIONS

Vocabulary

- Chemical Reactions
- Reactants

Standard 4: Students will understand that in chemical reactions matter and energy change forms, but the amounts of matter and energy do not change.

STANDARD 4, OBJECTIVE 1: IDENTIFY EVIDENCE OF CHEMICAL REACTIONS AND DEMONSTRATE HOW CHEMICAL EQUATIONS ARE USED TO DESCRIBE THEM.

EVIDENCES OF CHEMICAL REACTIONS

Lesson Objectives

- explain what happens during a chemical reaction.
- identify the reactants and products in any chemical reaction

Products

Yummy! S'mores are on the way! Did you ever toast marshmallows over a campfire? The sweet treats singe on the outside and melt on the inside. Both the fire and the toasted marshmallows are evidence of chemical changes. In the process of burning, the wood changes to ashes and gases, and the outside of the marshmallow turns brown and crispy. Neither the wood nor the marshmallows can change back to their original form. That's because burning is a chemical change and chemical changes are often impossible to undo. In this unit, you'll learn about many types of chemical changes, including how they occur and why you can't live without them.

What Is a Chemical Reaction?

A chemical reaction is a process in which some substances change into different substances. Substances that start a chemical reaction are called reactants. Substances that are produced in the reaction are called products. Reactants and products can be elements or compounds. Chemical reactions are represented by chemical equations, like the one below, in which reactants (on the left) are connected by an arrow to products (on the right).

Reactants → Products

Chemical reactions may occur quickly or slowly. Look at the two pictures in the Figure below. Both represent chemical reactions. In the picture on the left, a reaction inside a fire extinguisher causes foam to shoot out of the extinguisher. This reaction occurs almost instantly. In the picture on the right, a reaction causes the iron tool to turn to rust. This reaction occurs very slowly. In fact, it might take many years for all of the iron in the tool to turn to rust.

Fast

Slow

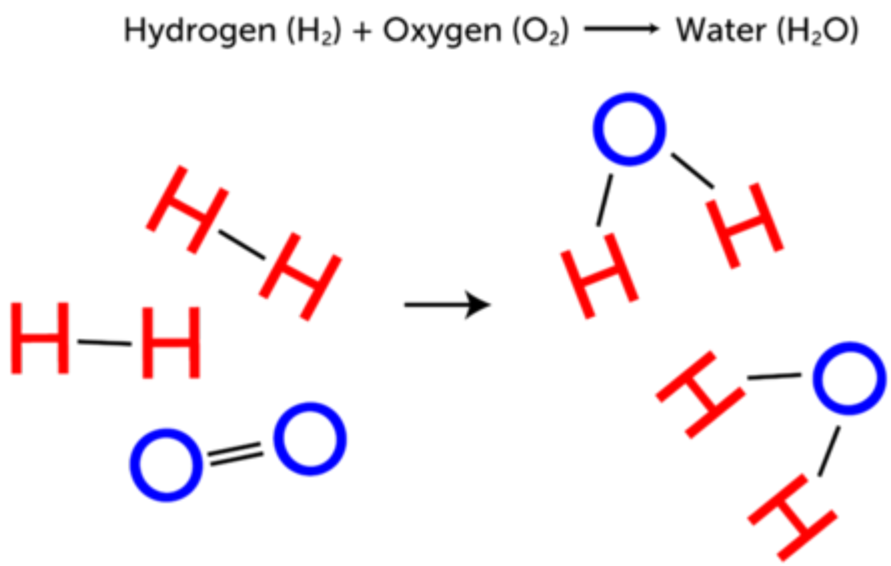
Q: What happens during a chemical reaction? Where do the reactants go, and where do the products come from?

A: During a chemical reaction, chemical changes take place. Some chemical bonds break and new chemical bonds form.

Breaking and Making Chemical Bonds:

The reactants and products in a chemical reaction contain the same atoms, but they are rearranged during the reaction. As a result, the atoms end up in different combinations in the products. This makes the products new substances that are chemically different from the reactants.

Consider the example of water forming from hydrogen and oxygen. Both hydrogen and oxygen gases exist as diatomic ("two-atom") molecules. These molecules are the reactants in the reaction. The Figure below shows that bonds must break to separate the atoms in the hydrogen and oxygen molecules. Then new bonds must form between hydrogen and oxygen atoms to form water molecules. The water molecules are the products of the reaction.



Q: Watch the animation of a similar chemical reaction at the following URL. Can you identify the reactants and the product in the reaction?
http://www.youtube.com/watch?v=SErRUqJ_x30

A: The reactants are hydrogen (H_2) and fluorine (F_2), and the product is hydrogen fluoride (HF).

SUMMARY

- All chemical reactions involve both reactants and products. Reactants are substances that start a chemical reaction, and products are substances that are produced in the reaction.

- A chemical reaction can be represented by the general chemical formula: Reactants → Products.

- Bonds break and reform during chemical reactions. Reactants and products contain the same atoms, but they are rearranged during the reaction, so reactants and products are different substances.

VOCABULARY

- product: Substance produced in a chemical reaction.
- reactant: Substance that starts a chemical reaction.

PRACTICE

Do the activities at the following URL for practice with reactants and products.
http://phet.colorado.edu/en/simulation/reactants-products-and-leftovers

REVIEW

1. Identify the reactants and products in the following chemical reaction:

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

2. How do reactants change into products during a chemical reaction?

Objective:

- **Identify potential signs that a chemical reaction has occurred.**

Look at the girl's hair in the photo above. It has obviously changed color. The process in which this occurred involved chemical reactions. How do you know that chemical reactions have occurred? The change in color is the most obvious clue.

What's Your Sign?

A change in color is just one of several potential signs that a chemical reaction has occurred. Other potential signs include:

- Change in temperature-Heat is released or absorbed during the reaction.
- Production of a gas—Gas bubbles are released during the reaction.
- Production of a solid-A solid settles out of a liquid solution. The solid is called a precipitate.

Examples of Chemical Reactions

Look carefully at the Figures below. All of the photos demonstrate chemical reactions. For each photo, identify a sign that one or more chemical reactions have taken place. You can see other examples of chemical reactions at this URL:
http://www.youtube.com/watch?v=gs0j1EZJ1Uc.

A burning campfire can warm you up on a cold day.

Dissolving an antacid tablet in water produces a fizzy drink.

Adding acid to milk produces solid curds of cottage cheese.

These vividly colored maple leaves were all bright green during the summer. Every fall, leaves of maple trees change to brilliant red, orange, and yellow colors. A change of color is a sign that a chemical change has taken place. Maple leaves change color because of chemical reactions.

Did you ever make a "volcano" by pouring vinegar over a "mountain" of baking soda? If you did, you probably saw the mixture bubble up and foam over. Did a chemical reaction occur? How do you know?

A: Yes, a chemical reaction occurred. You know because the bubbles are evidence that a gas has been produced and production of a gas is a sign of a chemical reaction.

SUMMARY

- Potential signs that chemical reactions have occurred include a change in color, change in temperature, production of a gas, and production of a solid precipitate.

PRACTICE

Watch the lab demonstration at the following URL. Then, describe the sequence of signs of chemical reactions that you observe in the demonstration.
http://www.youtube.com/watch?v=FKb-ZUbbXCA

Did you ever wonder what happens to a candle when it burns? A candle burning is a chemical change in matter. In a **chemical change**, one type of matter changes into a different type of matter, with different chemical properties. Chemical changes occur because of chemical reactions.

From Reactants to Products

All chemical reactions—including a candle burning—involve reactants and products.

- Reactants are substances that start a chemical reaction.
- Products are substances that are produced in the reaction.

When a candle burns, the reactants are the fuel (the candlewick and wax) and oxygen (in the air). The products are carbon dioxide gas and water vapor.

Relating Reactants and Products

The relationship between reactants and products in a chemical reaction can be represented by a chemical equation that has this general form:

$$\text{Reactants} \rightarrow \text{Products}$$

The arrow (→) shows the direction in which the reaction occurs. In many reactions, the reaction also occurs in the opposite direction. This is represented with another arrow pointing in the opposite direction (←).

Q: Write a general chemical equation for the reaction that occurs when a fuel such as candle wax burns.

A: The burning of fuel is a combustion reaction. The general equation for this type of reaction is:

$$\text{Fuel} + O_2 \rightarrow CO_2 + H_2O$$

Q: How do the reactants in a chemical reaction turn into the products?

A: Bonds break in the reactants, and new bonds form in the products. Nothing is lost, things simply are moved around)

Real World Application:
The following situation happens all too often. You have taken apart a piece of equipment to clean it up. When you put the equipment back together, somehow you have an extra screw or two. Or you find out that a piece is missing that was a part of the original equipment. In either case, you know something is wrong. You expect to end up with the same amount of material that you started with, not with more or less than what you had originally.

STANDARD 4, OBJECTIVE 2: ANALYZE EVIDENCE FOR THE LAWS OF CONSERVATION OF MASS AND CONSERVATION OF ENERGY IN CHEMICAL REACTIONS.

VOCABULARY

- law of conservation of mass: Law stating that matter cannot be created or destroyed in chemical reactions.

LAW OF CONSERVATION OF MASS

By the late 1700s, chemists accepted the definition of an element as a substance that cannot be broken down into a simpler substance by ordinary chemical means. It was also clear that elements combine with one another to form more complex substances called compounds. The chemical and physical properties of these compounds are different than the properties of the elements from which they were formed. There were questions about the details of these processes.

In the 1790s, a greater emphasis began to be placed on the quantitative analysis of chemical reactions. Accurate and reproducible measurements of the masses of reacting elements and the compounds they form led to the formulation of several basic laws. One of these is called the law of conservation of mass, which states that during a chemical reaction, the total mass of the products must be equal to the total mass of the reactants. In other words, mass cannot be created or destroyed during a chemical reaction, but must always be conserved.

As an example, consider the reaction between silver nitrate and sodium chloride. These two compounds will dissolve in water to form silver chloride and sodium nitrate. The silver chloride does not dissolve in water, so it forms a solid that we can filter off. When we evaporate the water, we can recover the sodium nitrate formed. If we react 58.5 grams of sodium chloride with 169.9 grams of silver nitrate, we start with 228.4 grams of materials. After the reaction is complete and the materials separated, we find that we have formed 143.4 grams of silver chloride and 85.0 grams of sodium nitrate, giving us a total mass of 228.4 grams for the products. So, the total mass of reactants equals the total mass of products, a proof of the law of conservation of mass.

Lavoisier and Conservation of Mass

How do scientists know that mass is always conserved in chemical reactions? Careful experiments in the 1700s by a

French chemist named Antoine Lavoisier led to this conclusion. Lavoisier carefully measured the mass of reactants and products in many different chemical reactions. He carried out the reactions inside a sealed jar, like the one in the Figure right. In every case, the total mass of the jar and its contents was the same after the reaction as it was before the reaction took place. This showed that matter was neither created nor destroyed in the reactions. Another outcome of Lavoisier's research was the discovery of oxygen. You can learn more about Lavoisier and his important research at:
http://www.youtube.com/watch?v=x9iZq3Zxb08.Antoine Lavoisier.

Q: Lavoisier carried out his experiments inside a sealed glass jar. Why was sealing the jar important for his results? What might his results have been if he hadn't sealed the jar?

A: Sealing the jar was important so that any gases produced in the reactions were captured and could be measured. If he hadn't sealed the jar, gases might have escaped detection. Then his results would have shown that there was less mass after the reactions than before. In other words, he would not have been able to conclude that mass is conserved in chemical reactions.

SUMMARY

- A chemical reaction occurs when some substances change chemically to other substances. Chemical reactions are represented by chemical equations.
- All chemical equations must be balanced because matter cannot be created or destroyed in chemical reactions.
- Antoine Lavoisier did careful experiments to discover the law of conservation of mass in chemical reactions.

PRACTICE

Watch the lab demonstration at the following URL, and then answer the questions below.
http://www.youtube.com/watch?v=774TbEUUM-A

1. What reaction is demonstrated in the video?
2. How can you tell that oxygen is used up in the reaction?
3. How can you tell that the product of the reaction is different from the iron that began the reaction?
4. What evidence shows that mass is conserved in the reaction?

Use the link below to answer the following questions:
https://www.etap.org/demo/grade8_science/lesson5/instruction1tutor.html

5. If you want to say something about chemical reactions, what would you use?
6. What does the Law of Conservation of Mass mean?
7. How much oxygen gas would I need if I react six molecules of hydrogen?

REVIEW

8. Why must all chemical equations be balanced?
9. How did Lavoisier demonstrate that mass is conserved in chemical reactions?
10. State the Law of Conservation of Mass.
11. What does this law mean?

CHEMICAL REACTIONS AND BALANCED EQUATIONS

Lesson Objectives

The student will:

- explain the roles of subscripts and coefficients in chemical equations.
- write a balanced chemical equation when given the unbalanced equation for any chemical reaction.
- explain the role of the law of conservation of mass in a chemical reaction.

Vocabulary

- balanced
- chemical equation
- coefficient
- subscript

Introduction

Even though chemical compounds are broken up to form new compounds during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products. The atoms are merely re-organized into different arrangements. In a complete chemical equation, the two sides of the equation must be balanced. That is, in a balanced chemical equation, the same number of each atom must be present on the reactant and product sides of the equation.

Why must chemical equations be balanced? It's the law! Matter cannot be created or destroyed in chemical reactions. This is the law of conservation of mass. In every chemical reaction, the same mass of matter must end up in the products as started in the reactants. Balanced chemical equations show that mass is conserved in chemical reactions, because the atoms themselves are conserved. The atoms are rearranged, but since they are neither created or destroyed the total mass doesn't change.

A chemical reaction occurs when some substances change chemically to other substances. Chemical reactions are represented by chemical equations. Consider a simple chemical reaction, the burning of methane. In this reaction, methane (CH_4) combines with oxygen (O2) in the air and produces carbon dioxide (CO_2) and water vapor (H_2O). The reaction is represented by the following chemical equation:

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

This balanced chemical equation can be read as 1 mole of CH_4 plus 2 moles of O_2 transforms into 1 mole of CO_2 and 2 moles of H_2O

This equation shows that one molecule of methane combines with two molecules of oxygen to produce one molecule of carbon dioxide and two molecules of water vapor. All chemical equations must be balanced. This means that the same number of each type of atom must appear on both sides of the arrow.

Q: Is the chemical equation for the burning of methane balanced? Count the atoms of each type on both sides of the arrow to find out.

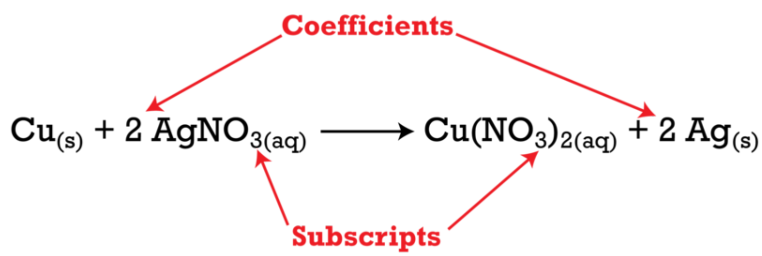
A: Yes, the equation is balanced. There is one carbon atom on both sides of the arrow. There are also four hydrogen atoms and four oxygen atoms on both sides of the arrow.

Balancing Equations

The process of writing a balanced chemical equation involves three steps. As a beginning chemistry student, you will not know whether or not two given compounds will react or not. Even if you saw them react, you would not know what the products are without running any tests to identify them. Therefore, for the time being, you will be told both the reactants and products in any equation you are asked to balance.

Step 1: Know what the reactants and products are, and write a word equation for the reaction.
Step 2: Write the formulas for all the reactants and products.
Step 3: Adjust the coefficients to balance the equation.

There are two types of numbers that appear in chemical equations. There are subscripts, which are part of the chemical formulas of the reactants and products, and there are coefficients that are placed in front of the formulas to indicate how many molecules of that substance are used or produced. In the chemical formula shown below, the coefficients and subscripts are labeled.



The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper(II) nitrate and two moles of solid silver. Recall that a subscript of 1 is not written - when no subscript appears for an atom in a formula, it is understood that only one atom is present. The same is true in writing balanced chemical equations. If only one atom or molecule is present, the coefficient of 1 is omitted.

The subscripts are part of the formulas, and once the formulas for the reactants and products are determined, the subscripts may not be changed. The coefficients indicate the mole ratios of each substance involved in the reaction and may be changed in order to balance the equation. Coefficients are inserted into the chemical equation in order to make the total number of each atom on both sides of the equation equal. Note that equation balancing is accomplished by changing coefficients, never by changing subscripts.

Example:
Write a balanced equation for the reaction that occurs between chlorine gas and aqueous sodium bromide to produce liquid bromine and aqueous sodium chloride.
Step 1: Write the word equation (keeping in mind that chlorine and bromine refer to the diatomic molecules).
chlorine + sodium bromide yields bromine + sodium chloride
Step 2: Substitute the correct formulas into the equation.

$$\mathrm{Cl_2 + NaBr \rightarrow Br_2 + NaCl}$$

Step 3: Insert coefficients where necessary to balance the equation.
By placing a coefficient of 2 in front of NaBr, we can balance the bromine atoms. By placing a coefficient of 2 in front of the NaCl, we can balance the chlorine atoms.

$$\mathrm{Cl_2 + 2\ NaBr \rightarrow Br_2 + 2\ NaCl}$$

A final check (always do this) shows that we have the same number of each atom on the two sides of the equation. We have also used the smallest whole numbers possible as the coefficients, so this equation is properly balanced.

Example:
Write a balanced equation for the reaction between aluminum sulfate and calcium bromide to produce aluminum bromide and calcium sulfate. Recall that polyatomic ions usually remain together as a unit throughout a chemical reaction.
Step 1: Write the word equation.
aluminum sulfate + calcium bromide yields aluminum bromide + calcium sulfate
Step 2: Replace the names of the substances in the word equation with formulas.

$$\mathrm{Al_2(SO_4)_3 + CaBr_2 \rightarrow AlBr_3 + CaSO_4}$$

Step 3: Insert coefficients to balance the equation.
In order to balance the aluminum atoms, we must insert a coefficient of 2 in front of the aluminum compound in the products.

$$\mathrm{Al_2(SO_4)_3 + CaBr_2 \rightarrow 2\ AlBr_3 + CaSO_4}$$

In order to balance the sulfate ions, we must insert a coefficient of 3 in front of the product $\mathrm{CaSO_4}$.

$$\mathrm{Al_2(SO_4)_3 + CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4}$$

In order to balance the bromine atoms, we must insert a coefficient of 3 in front of the reactant $CaBr_2$.

$$Al_2(SO_4)_3 + 3\ CaBr_2 \rightarrow 2\ AlBr_3 + 3\ CaSO_4$$

The insertion of the 3 in front of the reactant $CaBr_2$ also balances the calcium atoms in the product $CaSO_4$. A final check shows that there are two aluminum atoms, three sulfur atoms, twelve oxygen atoms, three calcium atoms, and six bromine atoms on each side. This equation is balanced.

Note that this equation would still have the same number of atoms of each type on each side with the following set of coefficients:

$$2\ Al_2(SO_4)_3 + 6\ CaBr_2 \rightarrow 4\ AlBr_3 + 6\ CaSO_4$$

Count the number of each type of atom on either side of the equation to confirm that this equation is "balanced." While this set of coefficients does "balanced" the equation, they are not the lowest set of coefficients possible. Chemical equations should be balanced with the simplest whole number coefficients. We could divide each of the coefficients in this equation by 2 to get another set of coefficients that still balance the equation and are whole numbers. Since it is required that an equation be balanced with the lowest whole number coefficients, the equation above is not properly balanced. When you have finished balancing an equation, you should not only check to make sure it is balanced, you should also check to make sure that it is balanced with the simplest set of whole number coefficients possible.

Example:

Balance the following skeletal equation. (The term "skeletal equation" refers to an equation that has the correct chemical formulas but does not include the proper coefficients.)

$$Fe(NO_3)_3 + \ NaOH \rightarrow Fe(OH)_3 + \ NaNO_3 \qquad \text{(skeletal equation)}$$

Solution:

We can balance the hydroxide ion by inserting a coefficient of 3 in front of the $NaOH$ on the reactant side.

$$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + NaNO_3$$

We can then balance the nitrate ions by inserting a coefficient of 3 in front of the sodium nitrate on the product side.

$$Fe(NO_3)_3 + 3\ NaOH \rightarrow Fe(OH)_3 + 3\ NaNO_3$$

Counting the number of each type of atom on the two sides of the equation will now show that this equation is balanced.

Example:
Given the following skeletal (un-balanced) equations, balance them.

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

H2SO4(aq) + Al(OH)3(aq) → Al2(SO4)3(aq) + H2O(l)

$$Ba(NO_3)_{2(aq)} + Na_2CO_{3(aq)} \rightarrow BaCO_{3(aq)} + NaNO_{3(aq)}$$

$$C_2H_{6(g)} + O_{2(g)} \rightarrow CO_{2(g)} + H_2O_{(l)}$$

Solution:

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

(sometimes the skeletal equation is already balanced)

3 H2SO4(aq) + 2 Al(OH)3(aq) → Al2(SO4)3(aq) + 6 H2O(l)

$$Ba(NO_3)_{2(aq)} + Na_2CO_{3(aq)} \rightarrow BaCO_{3(aq)} + 2\ NaNO_{3(aq)}$$

2 C2H6(g) + 7 O2(g) → 4 CO2(g) + 6 H2O(l)

LESSON SUMMARY

- A chemical reaction is the process in which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that symbolizes "yields," and the products on the right.

FURTHER READING / SUPPLEMENTAL LINKS

This video shows ten amazing chemical reactions that are fun to watch but dangerous to carry out.
http://listverse.com/2008/03/04/top-10-amazing-chemical-reactions/
Aqueous sodium hydroxide is mixed with gaseous chlorine to produce aqueous solutions of sodium chloride and sodium hypochlorite plus liquid water.
Solid xenon hexafluoride is mixed with liquid water to produce solid xenon trioxide and gaseous hydrogen fluoride.

QUANTIFYING CHEMICAL REACTIONS:

Vocabulary

- formula unit
- stoichiometry

Lesson Objectives

- explain the meaning of the term "stoichiometry."
- interpret chemical equations in terms of molecules, formula units, and moles.

Introduction

You have learned that chemical equations provide us with information about the types of particles that react to form products. Chemical equations also provide us with the relative number of particles and moles that react to form products. In this chapter, you will explore the quantitative relationships that exist between the reactants and products in a balanced equation. This is known as ***stoichiometry***.

Stoichiometry involves calculating the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation. The word stoichiometry actually comes from two Greek words: stoikheion, which means element, and metron, which means measure.

Interpreting Chemical Equations

Recall that a mole is a quantitative measure equivalent to Avogadro's number of particles. How does the mole relate to the chemical equation? Consider the following reaction:

$$N_2O_3 + H_2O \rightarrow 2\ HNO_3$$

We have learned that the coefficients in a chemical equation tell us the relative amounts of each substance involved in the reaction. One way to describe the ratios involved in the reaction above would be, "One molecule of dinitrogen trioxide, N_2O_3, plus one molecule of water yields two molecules of nitrous acid, HNO_3." However, because these are only ratios, this statement would be equally valid using units other than molecules. As a result, we could also say, "One mole of dinitrogen trioxide plus one mole of water yields two moles of nitrous acid."

We can use moles instead of molecules, because a mole is simply an amount equal to Avogadro's number, just like a dozen is an amount equal to 12. It is important to not use units that describe properties other than amount. For example, it would *not* be correct to say that one gram of dinitrogen trioxide plus one gram of water yields two grams of nitrous acid.

Now consider this reaction:

$$2\ CuSO_4 + KI \rightarrow 2\ CuI + 4\ K2SO_4 + I_2$$

Here, we can say, "Two moles of copper(II) sulfate react with four moles of potassium iodide, yielding two moles of copper(I) iodide, four moles of potassium sulfate, and one mole of molecular iodine." Although we can refer to molecules of iodine, I_2, it is generally not correct to refer to molecules of something like KI. Because KI is an ionic substance that exists as crystal lattices instead of discrete molecules, formula unit is used instead.

Example:
Indicate the ratio of compounds involved in the following balanced chemical equations. Describe the ratios in two ways: a) using the number of formula units or molecules and b) using the number of moles present.

1. $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$
2. $KBrO_3 + 6\ KI + 6\ HBr \rightarrow 7\ KBr + 3\ I_2 + 3\ H_2O$

Solution:

a) Two molecules of C_2H_6 plus seven molecules of O_2 yields four molecules of CO_2 plus six molecules of H_2O.
b) Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.
c) One formula unit of $KBrO_3$ plus six formula units of KI plus six molecules of HBr yields seven formula units of KBr plus three molecules of I_2 and three molecules of H_2O.
d) One mole of $KBrO_3$ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I_2 and three moles of H_2O.

LESSON SUMMARY

- Stoichiometry is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation.

FURTHER READING / SUPPLEMENTAL LINKS

This website contains various resources, including PowerPoint lectures, on many topics in chemistry, including one on stoichiometry.
http://www.chalkbored.com/lessons/chemistry-11.htm

REVIEW QUESTIONS

1. Distinguish between formula unit, molecule, and mole. Give examples in your answer.

CALCULATING MOLES

Earlier, we explored mole relationships in balanced chemical equations. In this lesson, we will use the mole as a conversion factor to calculate moles of product from a given number of moles of reactant, or we can calculate the number of moles of reactant from a given number of moles of product. This is called a "mole-mole" calculation. We will also perform "mass-mass" calculations, which allow you to determine the mass of reactant required to produce a given amount of product, or the mass of product you can obtain from a given mass of reactant.

Mole Ratios

A mole ratio is the relationship between two components of a chemical reaction. For instance, one way we could read the following reaction is that 2 moles of $H_{2(g)}$ react with 1 mole of $O_{2(g)}$ to produce 2 moles of $H_2O_{(l)}$.

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

The mole ratio of $H_{2(g)}$ to $O_{2(g)}$ would be:
What is the ratio of hydrogen molecules to water molecules? By examining the balanced chemical equation, we can see that the coefficient in front of the hydrogen is 2, while the coefficient in front of water is also 2. Therefore, the mole ratio can be written as:

Similarly, the ratio of oxygen molecules to water molecules would be:

In the following example, let's try finding the mole ratios by first writing a balanced chemical equation from a "chemical sentence."

Example:
Four moles of solid aluminum are mixed with three moles of gaseous oxygen to produce two moles of solid aluminum oxide. What is the mole ratio of (1) aluminum to oxygen, (2) aluminum to aluminum oxide, and (3) oxygen to aluminum oxide?

Solution:
Balanced chemical equation:

$$4\ Al_{(s)} + 3\ O_{2(g)} \rightarrow 2\ Al_2O_{3(s)}$$

1. mole ratio of aluminum to oxygen =
2. mole ratio of aluminum to aluminum oxide =
3. mole ratio of oxygen to aluminum oxide =

Example:
Write the balanced chemical equation for the reaction of solid calcium carbide (CaC_2) with water to form aqueous calcium hydroxide and acetylene (C_2H_2) gas. When written, find the mole ratios for (1) calcium carbide to water and (2) calcium carbide to calcium hydroxide.

Solution:
Balanced chemical equation:

$$CaC_{2(s)} + 2\ H_2O_{(l)} \rightarrow Ca(OH)_{2(aq)} + C_2H_{2(g)}$$

1. mole ratio of calcium carbide to water = $\frac{1\ \text{mol}\ CaC_2}{2\ \text{mol}\ H_2O}$ or $\frac{2\ \text{mol}\ H_2O}{1\ \text{mol}\ CaC_2}$
2. mole ratio of calcium carbide to calcium hydroxide = $\frac{1\ \text{mol}\ CaC_2}{1\ \text{mol}\ Ca(OH)_2}$

The correct mole ratios of the reactants and products in a chemical equation are determined by the balanced equation. Therefore, the chemical equation must always be balanced before the mole ratios are used for calculations. Looking at the unbalanced equation for the reaction of phosphorus trihydride with oxygen, it is difficult to guess the correct mole ratio of phosphorus trihydride to oxygen gas.

$$PH_{3(g)} + O_{2(g)} \rightarrow P_4O_{10(s)} + H_2O_{(g)}$$

Once the equation is balanced, however, the mole ratio of phophorus trihydride to oxygen gas is apparent.

Balanced chemical equation: $4\ PH_{3(g)} + 8\ O_{2(g)} \rightarrow P_4O_{10(s)} + 6\ H_2O_{(g)}$

The mole ratio of phophorus trihydride to oxygen gas, then, is: $\frac{4\ \text{mol}\ PH_3}{8\ \text{mol}\ O_2}$

Keep in mind that before any mathematical calculations are made relating to a chemical equation, the equation *must* be balanced.

Mole-Mole Calculations

In the chemistry lab, we rarely work with exactly one mole of a chemical. In order to determine the amount of reagent (reacting chemical) necessary or the amount of product expected for a given reaction, we need to do calculations using mole ratios.

Look at the following equation. If only 0.50 moles of magnesium hydroxide, $Mg(OH)_2$, are present, how many moles of phosphoric acid, H_3PO_4, would be required for the reaction?

$$2\ H_3PO_4 + 3\ Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6\ H_2O$$

Step 1: To determine the conversion factor, we want to convert from moles of $Mg(OH)_2$ to moles of H_3PO_4. Therefore, the conversion factor is:

$$\text{mole ratio} = \frac{2\ \text{mol}\ H_3PO_4}{3\ \text{mol}\ Mg(OH)_2}$$

Note that what we are trying to calculate is in the numerator, while what we know is in the denominator.

Step 2: Use the conversion factor to answer the question.

Therefore, if we have 0.50 mol of $Mg(OH)_2$, we would need 0.33 mol of H_3PO_4 to react with all of the magnesium hydroxide. Notice if the equation was not balanced, the amount of H_3PO_4 required would have been calculated incorrectly. The ratio would have been 1:1, and we would have concluded that 0.50 mol of H_3PO_4 were required.

Example:

How many moles of sodium oxide (Na_2O) can be formed from 2.36 mol of sodium nitrate $(NaNO_3)$ using the balanced chemical equation below?

$$10\ Na + 2\ NaNO_3 \rightarrow 6\ Na_2O + N_2O$$

Solution:

$$(2.36\ \text{mol}\ NaNO_3) \cdot \left(\frac{6\ \text{mol}\ Na_2O}{2\ \text{mol}\ NaNO_3}\right) = 7.08\ \text{mol}\ Na_2O$$

Example:
How many moles of sulfur are required to produce 5.42 mol of carbon disulfide, CS_2, using the balanced chemical equation below?

$$C + 2\ S \rightarrow CS_2$$

Solution:

$$(5.42\ \text{mol CS}_2) \cdot (\frac{2\ \text{mol S}}{1\ \text{mol CS}_2}) = 10.84\ \text{mol S}$$

MASS-MASS CALCULATIONS

A mass-mass calculation would allow you to solve one of the following types of problems:

- Determine the mass of reactant necessary to product a given amount of product
- Determine the mass of product that would be produced from a given amount of reactant
- Determine the mass of reactant necessary to react completely with a second reactant

As was the case for mole ratios, it is important to double check that you are using a *balanced* chemical equation before attempting any calculations.

USING DIMENSIONAL ANALYSIS TO SOLVE STOICHIOMETRY PROBLEMS

Many chemists prefer to solve stoichiometry problems with a single line of math instead of writing out the multiple steps. This can be done by using dimensional analysis, also called the factor-label method. Recall that this is simply a method that uses conversion factors to convert from one unit to another. In this method, we can follow the cancellation of units to obtain the correct answer.

For instance: 15.0 g of chlorine gas is bubbled through liquid sulfur to produce disulfur dichloride. How much product is produced in grams?

Step 1: As always, the first step is to correctly write and balance the equation:

$$Cl_{2(g)} + 2\ S_{(l)} \rightarrow S_2Cl_{2(l)}$$

Step 2: Identify what is being given (for this question, $15.0 \text{ g of } Cl_2$ is the given) and what is asked for $(\text{grams of } S_2Cl_2)$.

Step 3: Next, use the correct factors that allow you to cancel the units you don't want and get the unit you do want:

This is the known quantity.

This is the molar ratio from the balanced equation, which converts moles of known to moles of unknown.

This is the final answer.

$$15.0 \text{ g } Cl_2 \times \frac{1 \text{ mol } Cl_2}{71.0 \text{ g } Cl_2} \times \frac{1 \text{ mol } S_2Cl_2}{1 \text{ mol } Cl_2} \times \frac{135 \text{ g } S_2Cl_2}{1 \text{ mol } S_2Cl_2} = 28.5 \text{ g } S_2Cl_2$$

This is the molar mass of the known, which converts the given grams to moles.

This is the molar mass of the unknown, which converts moles to grams.

Example:

Consider the thermite reaction again. This reaction occurs between elemental aluminum and iron(III) oxide, releasing enough heat to melt the iron that is produced. If 500.0 g of iron is produced in the reaction, how much iron(III) oxide was placed in the original container?

Step 1: Write and balance the equation:

$$Fe_2O_{3(s)} + 2\ Al_{(s)} \rightarrow 2\ Fe_{(l)} + \ Al_2O_{3(s)}$$

Step 2: Determine what is given and what needs to be calculated:

$$\text{given} = 500.\ \text{g of Fe} \qquad \text{calculate} = \text{grams of } Fe_2O_3$$

Step 3: Set-up the dimensional analysis system:

$$500\ g\ Fe \cdot \frac{1\ mol\ Fe}{55.85\ g\ Fe} \cdot \frac{1\ mol\ Fe_2O_3}{2\ mol\ Fe} \cdot \frac{159.7\ g\ Fe_2O_3}{1\ mol\ Fe_2O_3} = 717\ g\ Fe_2O_3$$

Example:
Ibuprofen is a common painkiller used by many people around the globe. It has the formula $C_{13}H_{18}O_2$. If 200. g of Ibuprofen is combusted, how much carbon dioxide is produced?

Step 1: Write and balance the equation:

$2\ C_{13}H_{18}O_2(s) + 33\ O_2(g) \rightarrow 26\ CO_2(g) + 18\ H2O(l)$

Step 2: Determine what is given and what needs to be calculated:

$$\text{given} = 200.\text{g of ibuprofen} \qquad \text{calculate} = \text{grams of } CO_2$$

Step 3: Set-up the dimensional analysis system:

$$200.\,g\ C_{13}H_{18}O_2 \cdot \frac{1\ mol\ C_{13}H_{18}O_2}{206.3\ g\ C_{13}H_{18}O_2} \cdot \frac{26\ mol\ CO_2}{2\ mol\ C_{13}H_{18}O_2} \cdot \frac{44.01\ g\ CO_2}{1\ mol\ CO_2} = 555\ g\ CO_2$$

Example:
If sulfuric acid is mixed with sodium cyanide, the deadly gas hydrogen cyanide is produced. How much sulfuric acid must be placed in the container to produce 12.5 g of hydrogen cyanide?

Step 1: Write and balance the equation:

$2\ NaCN(s) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ HCN(g)$

Step 2: Determine what is given and what needs to be calculated:

$$\text{given} = 12.5\text{ g HCN} \qquad \text{calculate} = \text{grams of } H_2SO_4$$

Step 3: Set-up the dimensional analysis system:

$$12.5g\ HCN \cdot \frac{1\ mol\ HCN}{27.0\ g\ HCN} \cdot \frac{1\ mol\ H_2SO_4}{2\ mol\ HCN} \cdot \frac{98.06\ g\ H_2SO_4}{1\ mol\ H_2SO_4} = 22.7\ g\ H_2SO_4$$

LESSON SUMMARY

- The coefficients in a balanced chemical equation represent the relative amounts of each substance in the reaction.
- When the moles of one substance in a reaction is known, the coefficients of the balanced equation can be used to determine the moles of all the other substances.
- Mass-mass calculations can be done using dimensional analysis.

REVIEW QUESTIONS

1. How many moles of water vapor can be produced from 2 moles of ammonia for the following reaction between ammonia and oxygen: $4\ NH_{3(g)} + 5\ O_{2(g)} \rightarrow 4\ NO_{(g)} + 6\ H_2O_{(g)}$?
   a. 3 mol
   b. 6 mol
   c. 12 mol
   d. 24 mol
2. How many moles of bismuth(III) oxide can be produced from 0.625 mol of bismuth in the following reaction: $Bi_{(s)} + O_{2(g)} \rightarrow Bi_2O_{3(s)}$? (Note: equation may not be balanced.)
   a. 0.313 mol
   b. 0.625 mol
   c. 1 mol
   d. 1.25 mol
   e. 2 mol
3. For the following reaction, balance the equation and then determine the mole ratio of moles of $B(OH)_3$ to moles of water: $B_2O_{3(s)} + H_2O_{(l)} \rightarrow B(OH)_{3(s)}$.
   a. $1 : 1$
   b. $2 : 3$
   c. $3 : 2$
   d. None of the above.

4. Write the balanced chemical equation for the reactions below and find the indicated molar ratio.

   a. Gaseous propane (C_3H_8) combusts to form gaseous carbon dioxide and water. Find the molar ratio of O_2 to CO_2.

   b. Solid lithium reacts with an aqueous solution of aluminum chloride to produce aqueous lithium chloride and solid aluminum. Find the molar ratio of $AlCl_{3(aq)}$ to $LiCl_{(aq)}$.

   c. Gaseous ethane (C_2H_6) combusts to form gaseous carbon dioxide and water. Find the molar ratio of $CO_{2(g)}$ to $O_{2(g)}$.

   d. An aqueous solution of ammonium hydroxide reacts with an aqueous solution of phosphoric acid to produce aqueous ammonium phosphate and water. Find the molar ratio of $H_3PO_{4(aq)}$ to $H_2O_{(l)}$.

   e. Solid rubidium reacts with solid phosphorous to produce solid rubidium phosphide. Find the molar ratio of Rb(s) to P(s).

5. For the given reaction (unbalanced): $Ca_3(PO_4)_2 + SiO_2 + C \rightarrow CaSiO_3 + CO + P$

   a. how many moles of silicon dioxide are required to react with 0.35 mol of carbon?

   b. how many moles of calcium phosphate are required to produce 0.45 mol of calcium silicate?

6. For the given reaction (unbalanced): $FeS + O_2 \rightarrow Fe_2O_3 + SO_2$

   a. how many moles of iron(III) oxide are produced from 1.27 mol of oxygen?

   b. how many moles of iron(II) sulfide are required to produce 3.18 mol of sulfur dioxide?

7. Write the following balanced chemical equation. Ammonia and oxygen are allowed to react in a closed container to form nitrogen and water. All species present in the reaction vessel are in the gas state.

   a. How many moles of ammonia are required to react with 4.12 mol of oxygen?

   b. How many moles of nitrogen are produced when 0.98 mol of oxygen are reacted with excess ammonia?

8. How many grams of nitric acid will react with 2.00 g of copper(II) sulfide given the following reaction between copper(II) sulfide and nitric acid: ?

   $3\ CuS(s) + 8\ HNO_3(aq) \rightarrow 3\ Cu(NO_3)_2(aq) + 2\ NO(g) + 4\ H_2O(l) + 3\ S(s)$

   a. $0.49\ \mathrm{g}$
   b. $1.31\ \mathrm{g}$
   c. $3.52\ \mathrm{g}$
   d. $16.0\ \mathrm{g}$

9. Donna was studying the following reaction for a stoichiometry project: $\mathrm{S}_{(s)} + 3\ \mathrm{F}_{2(g)} \rightarrow \mathrm{SF}_{6(s)}$ She wondered how much she could obtain if she used 3.5 g of fluorine. What mass of $\mathrm{SF}_6(s)$ would she obtain from the calculation using this amount of fluorine?

   a. $3.5\ \mathrm{g}$
   b. $4.5\ \mathrm{g}$
   c. $10.5\ \mathrm{g}$
   d. $13.4\ \mathrm{g}$

HEAT ENERGY IN CHEMICAL REACTIONS

Vocabulary

- enthalpy
- surroundings
- system

Lesson Objectives

- define the terms system and surroundings in the context of a chemical reaction.
- identify the system and surroundings in a chemical reaction.
- describe how heat is transferred in endothermic and exothermic reactions.

Introduction

Energy is defined as the ability to do work. Energy is often divided into two types: kinetic energy and potential energy. Kinetic energy is the energy of motion, while potential energy is the energy of position.

Molecules contain potential energy in their physical states and in their chemical bonds. You will remember that when solid substances were changed into liquid, energy had to be added to provide the heat of melting. That energy was used to pull the molecules further apart, changing solid into liquid. That energy is then stored in the liquid as potential energy due to the greater distances between attracting molecules. For similar reasons, energy also has to be added to convert a liquid into a gas.

Chemical bonds store potential energy in a slightly different way. To understand how chemical bonds store energy, we can view a substance as having maximum potential energy in bonds when all the atoms of the substance are separated from each other and are in atomic form (no bonds). The atoms can then form many different bonds. When bonds form, energy is released and the potential energy of the substance decreases.

All Chemical Reactions Involve Energy

Every system or sample of matter has energy stored in it. When chemical reactions occur, the new bonds formed never have exactly the same amount of potential energy as the bonds that were broken. Therefore, all chemical reactions involve energy changes. Energy is either given off or taken on by the reaction.

Before any reaction can occur, reactant bonds need to be broken. A minimum amount of energy, that is the activation energy, must be supplied before any reaction can take place. This minimum energy might be in the form of heat or electrical current.

$$NaCl(aq) + AgNO_3(aq) \rightarrow AgCl(s) + NaNO_3(aq) \quad \Delta H = -166 \text{ kJ}$$

The equation above represents a chemical reaction where energy is produced. This means that there is less energy stored in the bonds of the products than there is in the bonds of the reactants. Therefore, extra energy is left over when the reactants become the products.

In comparison, the decomposition of mercury(II) oxide requires a net input of energy.

$$2\ HgO_{(s)} \rightarrow 2\ Hg_{(l)} + O_{2(g)} \quad \triangle H = 181.7\ kJ$$

In this reaction, there is less energy stored in the bonds of the mercury(II) oxide than is stored in the bonds of the products. Therefore, extra energy had to be added to the reaction to form the products.

These two equations represent the two types of chemical reactions that involve energy transfer and illustrate that chemical reactions involve energy.

Bond Breaking and Bond Forming

In the previous section, we looked at two different types of reactions that involved energy. $\triangle H$, or the heat of reaction, measures the change in the internal energy of the reaction. The internal energy is the sum of all the energy of the chemical system, that is, the potential and the kinetic. $\triangle H$ is also known as change in enthalpy. Enthalpy is the amount of energy a system or substance contains and cannot be measured directly. What can be measured is the change in enthalpy (or $\triangle H$). When the bonds of the reactants contain more energy than the bonds of the products, the reaction is exothermic and $\triangle H$ is negative. Conversely, when the bonds of the products contain more energy than the bonds of the reactants, the reaction is endothermic and $\triangle H$ is positive.

System and Surroundings

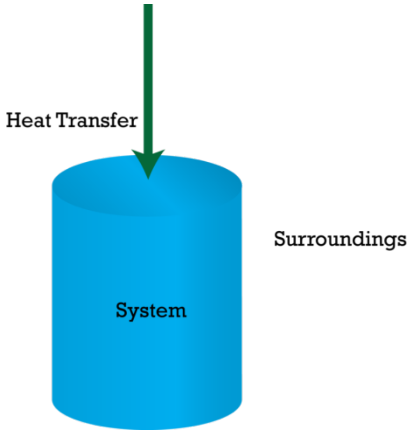
Many chemical reactions take place in an open system. Whether the reaction is exothermic or endothermic, there is energy transfer between the system and the surroundings. For example, you take an ice cube out of the refrigerator and place it on a counter. As the ice cube melts, it requires a small amount of heat to be absorbed from the surroundings (the room) in order to produce the liquid.

$$H_2O_{(s)} \rightarrow H_2O_{(l)} \quad \triangle H = +6.01\ kJ/mol$$

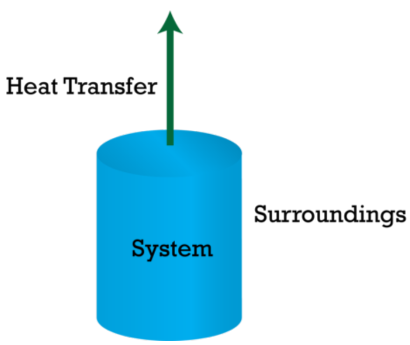
The system in this example is the ice cube melting to form the liquid water. The surroundings are the container, room, and building where the reaction is taking place. In other words, the system involves the reactants and products in the reaction. The surroundings are everything else.

Before moving any further in our discussion, it is important to distinguish between heat and temperature when talking about heat transfer. Heat is the total amount of energy that is transferred between the system and the surroundings, while temperature is the average kinetic energy of a substance. The temperature measures the kinetic energy of the reactant and/or product particles. Consider a cup of water and a bucket of water: both will boil at the same temperature (100^{o}C), but it will require different amounts of heat to bring these two volumes of water to a boil.

An endothermic reaction system absorbs heat from the surroundings and has a positive $\triangle H$. Phase changes from a solid to a liquid to a gas are all endothermic. The diagram below illustrates the transfer of heat energy from the surroundings to the system.



An exothermic reaction system releases heat to the surroundings and has a negative value of $\triangle H$. Just as phase changes from a solid to a liquid to a gas are all endothermic, the reverse of these changes are exothermic. Other reactions that you know as exothermic may be the combustion of fuels. Fuels used to drive your car and heat your home all involve reactions that are exothermic in nature. The figure below illustrates the transfer of heat energy to the surroundings from the system.



The reason that endothermic reactions have positive $\triangle H$ sand exothermic reactions have negative $\triangle H$ sis because the definition of $\triangle H$ is:

$$\triangle H = H_{\text{products}} - H_{\text{reactants}}$$

Therefore, when the reactants contain more potential energy in their bonds than do the products, energy is given off. Looking at the equation above, since a larger number is subtracted from a smaller one, the answer is negative.

Example:
Which of the following processes are endothermic, and which are exothermic?

1. water boiling
2. gasoline burning
3. water vapor condensing
4. iodine crystals subliming
5. ice forming on a pond

Solution:

1. endothermic – state change from liquid to a gas absorbs heat from the surroundings
2. exothermic – combustion releases heat to the surroundings
3. exothermic – state change from gas to a liquid releases heat to the surroundings
4. endothermic – state change from solid to a liquid absorbs heat from the surroundings
5. exothermic – state change from liquid to a solid releases heat to the surroundings

LESSON SUMMARY

- $\triangle H$, or the heat of reaction, measures the change in the enthalpy of the reaction.
- The enthalpy is the sum of all the energy in a chemical system.
- The system involves the reactants and products in the reaction.
- The surroundings are everything else.
- An endothermic reaction system absorbs heat from the surroundings. An exothermic reaction system releases heat to the surroundings.

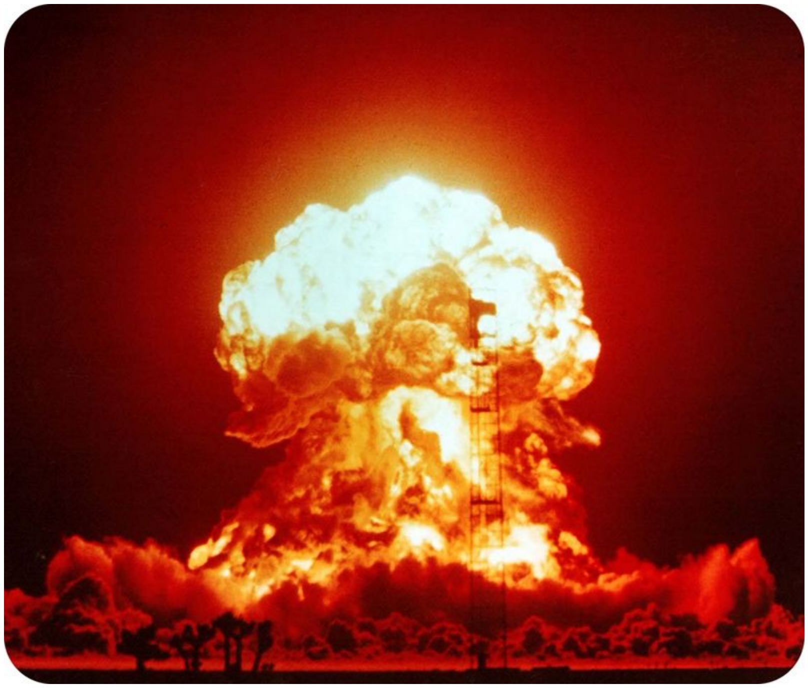
REVIEW QUESTIONS

1. How does a campfire involve energy? Is it endothermic or exothermic? What would be the system and what would be the surroundings?
2. If a chemical reaction absorbs heat from the surroundings, it is said to be what?
    a. in equilibrium
    b. in a closed system
    c. an exothermic reaction
    d. an endothermic reaction
3. If a chemical reaction releases heat to the surroundings, it is said to be what?
    a. in equilibrium
    b. in a closed system
    c. an exothermic reaction
    d. an endothermic reaction
4. Symbolically, change in enthalpy is represented as:
    a. H
    b. $\triangle H$
    c. E
    d. $\triangle E$

5. Which of the following processes would be endothermic?
    a. natural gas burning
    b. melting chocolate
    c. fireworks exploding
    d. Steam condensing

6. Which of the following processes would be exothermic?
    a. gasoline burning
    b. evaporation of ether
    c. melting butter
    d. boiling water

ENERGY TRANSFORMATIONS IN CHEMICAL REACTIONS



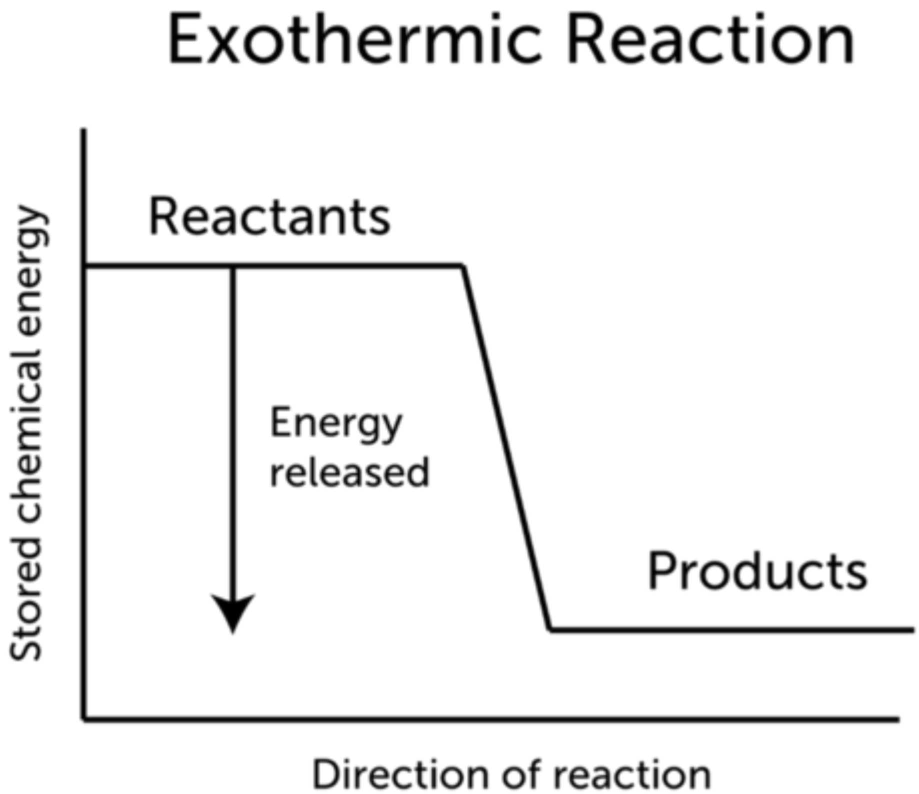
This mushroom cloud was produced in a 1953 nuclear bomb test in Nevada. There's no doubt that the explosion gave off a huge amount of energy. Although not as impressive as nuclear reactions, many chemical reactions also give off energy. These reactions are called exothermic reactions.

What Is an Exothermic Reaction?

All chemical reactions involve energy. Energy is used to break bonds in reactants, and energy is released when new bonds form in products. In some chemical reactions, called **endothermic reactions**, less energy is released when new bonds form in the products than is needed to break bonds in the reactants. The opposite is true of exothermic reactions. In an **exothermic reaction**, it takes less energy to break bonds in the reactants than is released when new bonds form in the products.

Energy Change in Exothermic Reactions

The word *exothermic* literally means "turning out heat." Energy, often in the form of heat, is released as an exothermic reaction proceeds. This is illustrated in the **Figure** below. The general equation for an exothermic reaction is:



Reactants → Products + Energy

If the energy produced in an exothermic reaction is released as heat, it results in a rise in temperature. As a result, the products are likely to be warmer than the reactants.

Q: You turn on the hot water faucet, and hot water pours out. How does the water get hot? Do you think that an exothermic reaction might be involved?

A: A hot water heater increases the temperature of water in most homes. Many hot water heaters burn a fuel such as natural gas. The burning fuel causes the water to get hot because combustion is an exothermic reaction.

Combustion as an Exothermic Reaction

All combustion reactions are exothermic reactions. During a combustion reaction, a substance burns as it combines with oxygen. When substances burn, they usually give off energy as heat and light. Look at the big bonfire in the Figure below. The combustion of wood is an exothermic reaction that releases a lot of energy as heat and light. You can see the light energy the fire is giving off. If you were standing near the fire, you would also feel its heat.

SUMMARY

- An exothermic reaction is a chemical reaction in which less energy is needed to break bonds in the reactants than is released when new bonds form in the products.
- During an exothermic reaction, energy is constantly given off, often in the form of heat.
- All combustion reactions are exothermic reactions. During combustion, a substance burns as it combines with oxygen, releasing energy in the form of heat and light.

VOCABULARY

- exothermic reaction: Chemical reaction that releases energy because it takes less energy to break bonds in the reactants than is released when new bonds form in the product.

PRACTICE

Watch the video about exothermic reactions at the following URL, and then answer the questions below:
http://www.videojug.com/film/a-guide-to-exothermic-reactions

1. Why do exothermic reactions heat up?
2. Explain why this reaction is exothermic: $CH_4 + F_2 \rightarrow CH_3F + HF$.

REVIEW

3. What is an exothermic reaction?
4. Why are the products of an exothermic reaction likely to be warmer than the reactants?
5. Describe an example of an exothermic reaction.

ENDOTHERMIC REACTIONS = CHEMICAL REACTIONS THAT REQUIRE ENERGY

Did you ever use an instant ice pack like? You don't have to pre-cool it in the freezer. All you need to do is squeeze the pack and it starts to get cold. How does this happen? The answer is an endothermic chemical reaction.

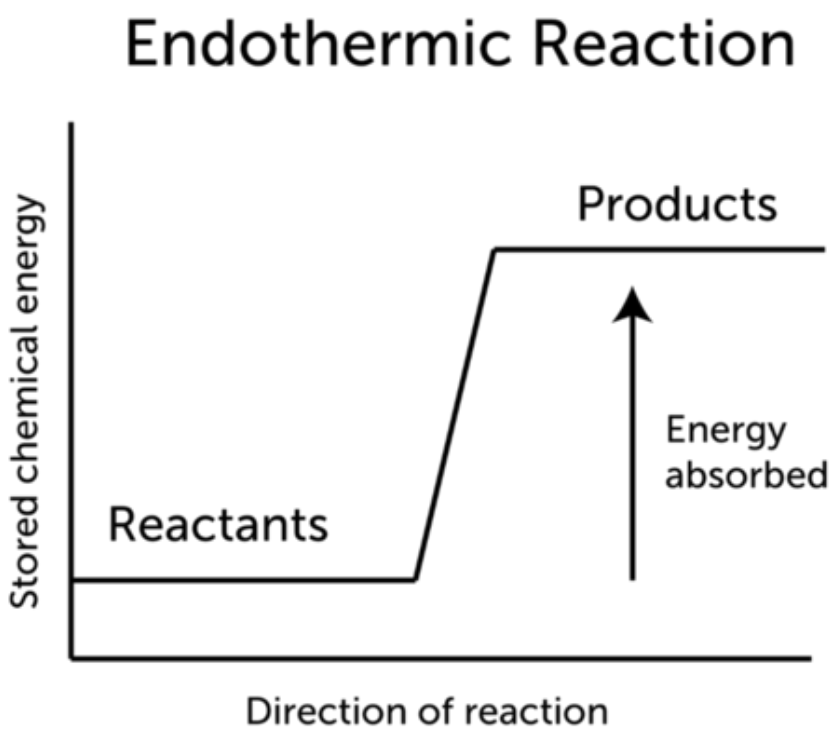
What Is an Endothermic Reaction?

All chemical reactions involve energy. Energy is used to break bonds in reactants, and energy is released when new bonds form in products. In some chemical reactions, called exothermic reactions, more energy is released when new bonds form in the products than is needed to break bonds in the reactants. The opposite is true of endothermic reactions. In an endothermic reaction, it takes more energy to break bonds in the reactants than is released when new bonds form in the products.

Energy Change in Endothermic Reactions

The word endothermic literally means "taking in heat." A constant input of energy, often in the form of heat, is needed to keep an endothermic reaction going. This is illustrated in the Figure below. Energy must be constantly added because not enough energy is released when the products form to break more bonds in the reactants. The general equation for an endothermic reaction is:

$$\text{Reactants} + \text{Energy} \rightarrow \text{Products}$$



In endothermic reactions, the temperature of the products is typically lower than the temperature of the reactants. The drop in temperature may be great enough to cause liquids to freeze.

Q: Now can you guess how an instant cold pack works?

A: Squeezing the cold pack breaks an inner bag of water, and the water mixes with a chemical inside the pack. The chemical and water combine in an endothermic reaction. The energy needed for the reaction to take place comes from the water, which gets colder as the reaction proceeds.

Photosynthesis

One of the most important series of endothermic reactions is photosynthesis. In photosynthesis, plants make the simple sugar glucose (C6H12O6) from carbon dioxide (CO2) and water (H2O). They also release oxygen (O2) in the process. The reactions of photosynthesis are summed up by this chemical equation:

$$6\ CO_2 + 6\ H_2O \rightarrow C_6H_{12}O_6 + 6\ O_2$$

The energy for photosynthesis comes from light. Without light energy, photosynthesis cannot occur. As you can see in the Figure below, plants can get the energy they need for photosynthesis from either sunlight or artificial light.


SUMMARY

- An endothermic reaction is a chemical reaction in which more energy is needed to break bonds in the reactants than is released when new bonds form in the products.
- A constant input of energy, often in the form of heat, is needed to keep an endothermic reaction going.
- One of the most important series of endothermic reactions is photosynthesis. The energy needed for photosynthesis comes from light.

VOCABULARY

- endothermic reaction: Chemical reaction that needs a constant input of energy to continue because it takes more energy to break bonds in the reactants than is released when new bonds form in the products.

REVIEW

1. What is an endothermic reaction?
2. Why is the temperature of products likely to be lower than the temperature of reactants in an endothermic reaction?
3. Describe an example of an endothermic reaction
4. Using collected data, report the loss or gain of heat energy in a chemical reaction.

ELECTROCHEMISTRY

Lesson Objectives

- Describe oxidation-reduction reactions
- Diagram an electrochemical cell

Introduction

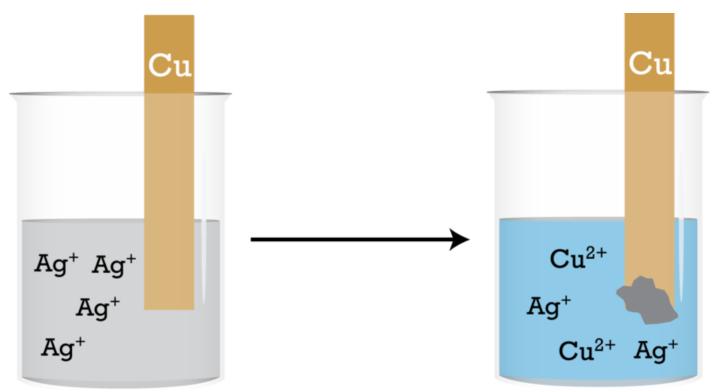
Previously, we have learned how atoms can gain or lose electrons forming charged ions. For instance, we specified that the oxidation state of an atom in the alkali family is $+1$. This occurs because these atoms always lose one electron to form an ion. In comparison, the charge of an ion in the halogen family is -1. Again, this is because a halogen atom tends to gain an electron to form its ions. In this chapter, we will re-examine the concept of charges in more detail.

Some types of chemical reactions are capable of producing an electrical current (or electricity). These are the type of reactions that occur inside batteries. When a reaction is arranged to produce an electric current as it runs, the arrangement is called an electrochemical cell or a Galvanic Cell.

If a strip of copper is placed in a solution of silver nitrate, the following reaction takes place:

$$2\ \mathrm{Ag}^{+} + \mathrm{Cu} \rightarrow 2\ \mathrm{Ag} + \mathrm{Cu}^{2+}$$

In this reaction (illustrated below), copper atoms are donating electrons to silver ions so that the silver ions are reduced to silver atoms and the copper atoms are oxidized to copper(II) ions.

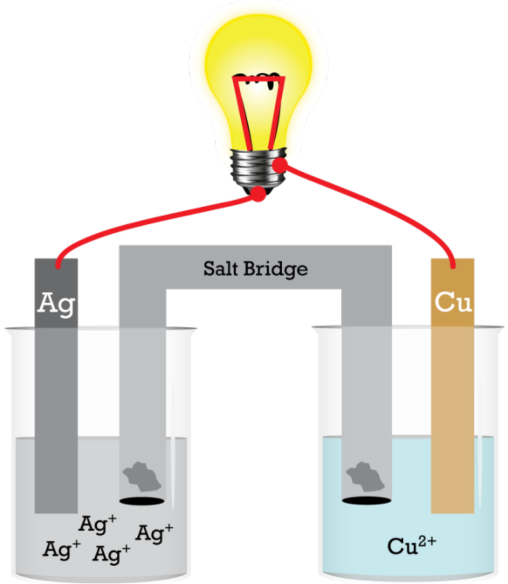


As the reaction occurs, an observer would see the solution slowly turn blue (Cu^{2+} ions are blue in solution), and a mass of solid silver atoms would build up on the copper strip.

Reactions that involve a transfer of electrons are called oxidation-reduction reactions. Atoms that lose electrons (resulting in an increase in the charge of the atom) are said to be oxidized. At the same time, another atom gains those electrons (resulting in a decrease in the charge) and are said to be reduced. Oxidation-reduction reactions can be used to generate an electrical current—or the movement of electrons through a wire.
Although electrons are transferred (moved from the copper atoms to the silver ions) in this example, no electricity is being produced as the electrons are not flowing through a wire.

Electrochemical Cells

The reaction $2\ Ag^{+} + Cu \rightarrow 2\ Ag + Cu^{2+}$ is one that could be physically arranged to produce an external electric current. To do this, the two half-reactions must occur in separate compartments, and the separate compartments must remain in contact through an ionic solution and an external wire.



In the electrochemical cell illustrated above, the copper metal must be separated from the silver ions to avoid a direct reaction. Each electrode in its solution could be represented by the following half-reactions:

$$Cu \rightarrow Cu^{2+} + 2\ e^{-} \qquad\qquad Ag \rightarrow Ag^{+} + e^{-}$$

The external wire will allow electrons to flow between the metal strips. In each half-cell, atoms may be oxidized to ions, leaving excess electrons on the electrode. It can be determined experimentally that electrons will flow in the wire from the copper electrode to the silver electrode. The reason for the direction of electron flow can be explained in several ways. We could say that silver ions have a greater electron affinity than copper ions, so the silver atoms pull the electrons through the wire from copper. On the other hand, we could also say that copper atoms have a greater tendency to give up electrons than silver, so the copper electrode becomes more negative and will push the electrons through the wire toward the silver electrode. Whichever way we look at it, the electrons flow from copper to silver in the external wire. The silver electrode, therefore, will acquire an excess of electrons.

$$\mathrm{Ag}^{+} + e^{-} \rightarrow \mathrm{Ag}$$

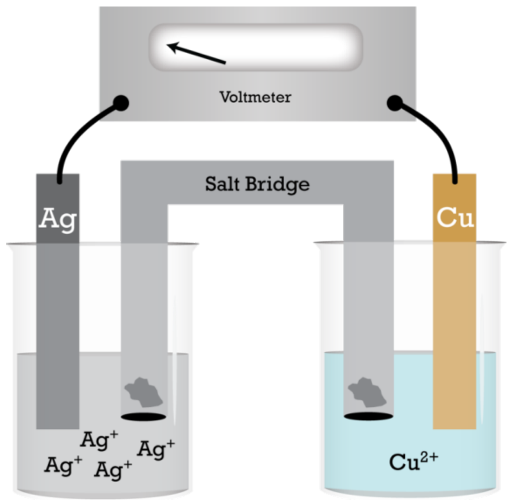
As the electrons on the silver electrode are used up in the reduction of silver ions, more copper atoms are oxidized, and more electrons sent through the wire. The net reaction for the entire cell is:

$$2\ \mathrm{Ag}^{+}_{(aq)} + \mathrm{Cu}_{(s)} \rightarrow 2\ \mathrm{Ag}_{(aq)} + \mathrm{Cu}^{2+}_{(aq)}$$

As the cell runs, electrons are transferred from the copper half-cell to the silver half-cell. Unchecked, this reaction would result in the silver half-cell becoming negatively charged and the copper half-cell becoming positively charged. Once the half-cells became charged in that manner, the reaction could only continue to run if it produced sufficient energy to take electrons away from a positive charge and push them onto an already negatively charged half-cell. Chemical reactions, however, do not produce enough energy to push electrons against a charge gradient.

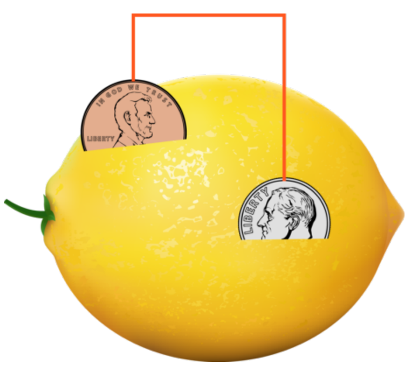
The salt bridge is present to allow negative ions to flow from the silver half-cell to the copper half-cell, and positive ions to flow from the copper half-cell to the silver half-cell. This migration of ions balances the charge movement of the electrons. The salt bridge is the upside-down U-shaped tube connecting the two beakers in the image above. This tube is filled with an ionic solution, and the ends are fitted with porous plugs. An ionic solution is chosen such that neither of its ions will react chemically with any of the other ions in the system. The porous plugs are there to avoid general mixing but to allow ion migration. If the salt bridge were not present or were removed, the reaction would immediately stop. As long as the two beakers remain neutral and there are sufficient reactants (Ag^{+} and Cu) to continue the reaction, the reaction will continue to run.

The electrons that pass through the external circuit can do useful work, such as lighting lights, running cell phones, and so forth. Several cells can be operated together to produce greater current. When we have a series of cells operating together as one, we call the arrangement a battery. If the light bulb is removed from the circuit with the electrochemical cell and replaced with a voltmeter (see illustration below), the voltmeter will measure the voltage (electrical potential energy per unit charge) of the combination of half-cells.



The size of the voltage produced by a cell depends on the temperature, the metals used for electrodes, and the concentrations of the ions in the solutions. If you increase the concentration of the reactant ion (not the product ion), the reaction rate will increase and so will the voltage.

It may seem complicated to construct an electrochemical cell because of all its complexities. Electrochemical cells, however, are actually easy to make and sometimes even occur accidentally. If you take two coins of different metal composition, one copper and one silver for example, and push them part way through the peel of a whole lemon (as illustrated below), upon connecting the two coins with a wire, a small electric current will flow.



Electrochemical cells occasionally occur accidentally when two water pipes of different material are connected. The reactions at the joint cause a great deal of corrosion. Plumbing professionals take great care to make sure such reactions do not occur at pipe joints.

CHAPTER 5

CHEMICAL DYNAMICS

Vocabulary

- activated complex
- activation energy
- collision frequency
- collision theory
- kinetic molecular theory
- threshold energy

Standard 5: Students will understand that many factors influence chemical reactions and some reactions can achieve a state of dynamic equilibrium.

STANDARD 5, OBJECTIVE 1: EVALUATE FACTORS SPECIFIC TO COLLISIONS (E.G., TEMPERATURE, PARTICLE SIZE, CONCENTRATION, AND CATALYSTS) THAT EFFECT THE RATE OF CHEMICAL REACTIONS.

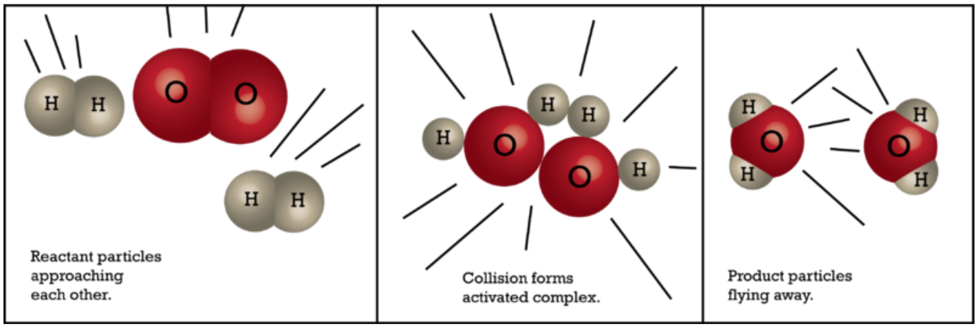
Lesson Objectives

- define the collision theory.
- describe the conditions for successful collisions.
- explain how the kinetic molecular theory applies to the collision theory.
- describe the rate in terms of the conditions of successful collisions.

Introduction

Consider the chemical reaction $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$. In the reactants, the carbon atoms are bonded to hydrogen atoms, and the oxygen atoms are bonded to other oxygen atoms. The bonds that are present in the products cannot form unless the bonds in the reactants are first broken, which requires an input of energy.

The energy to break the old bonds comes from the kinetic energy of the reactant particles. The reactant particles are moving around at random with an average kinetic energy related to the temperature. If a reaction is to occur, the kinetic energy of the reactants must be high enough that when the reactant particles collide, the collision is forceful enough to break the old bonds. Once the old bonds are broken, the atoms in the reactants would be available to form new bonds. At that point, the new bonds of the products could be formed. When the new bonds are formed, potential energy is released. The potential energy that is released becomes kinetic energy that is absorbed by the surroundings (primarily the products, the solvent solution if there is one, and the reaction vessel).



In those cases where the reactants do not collide with enough energy to break the old bonds, the reactant particles will simply bounce off each other and remain reactant particles.

How Reactions Occur

We know that a chemical system can be made up of atoms $(H_2,\ N_2,\ K)$, ions $(NO_3^-,\ Cl^-,\ Na^+)$, or molecules $(H_2O,\ CH_3CH_3,\ C_{12}H_{22}O_{11})$. We also know that in a chemical system, these particles are moving around at random. The collision theory explains why reactions occur between these atoms, ions, and/or molecules at the particle level. The collision theory provides us with the ability to predict what conditions are necessary for a successful reaction to take place. These conditions include:

- the particles must collide with each other
- the particles must have proper orientation
- the particles must collide with sufficient energy to break the old bonds

The rate of the reaction depends on the fraction of molecules that have enough energy and that collide with the proper orientation.

LESSON SUMMARY

- The collision theory explains why reactions occur between atoms, ions, and/or molecules and allows us to predict what conditions are necessary for a successful reaction to take place.
- The kinetic molecular theory provides the foundation for the collision theory on the molecular level.
- The minimum amount of energy necessary for a reaction to take place is known as the threshold energy.
- With increasing temperature, the kinetic energy of the particles and the number of particles with energy greater than the activation energy increases.
- The total number of collisions per second is known as the collision frequency, regardless of whether these collisions are successful or not.

REVIEW QUESTIONS

1. According to the collision theory, it is not enough for particles to collide in order to have a successful reaction to produce products. Explain.

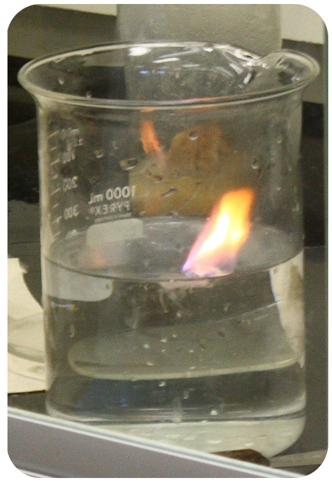
2. Due to the number of requirements for a successful collision, according to the collision theory, the percentage of successful collisions is extremely small. Yet, chemical reactions are still observed at room temperature and some at very reasonable rates. Explain.

3. According to the collision theory, which of the following must happen in order for a reaction to be successful: i. particles must collide, ii. particles must have proper geometric orientation, iii. particles must have collisions with enough energy?
    a. i, ii
    b. i, iii
    c. ii, iii
    d. i, ii, iii

4. **What would happen in a collision between two particles if there was insufficient kinetic energy and improper geometric orientation?**
    a. The particles would rebound and there would be no reaction.
    b. The particles would keep bouncing off each other until they eventually react, therefore the rate would be slow.
    c. The particles would still collide but the byproducts would form.
    d. The temperature of the reaction vessel would increase.

5. **Illustrate the successful collision that would occur between the following:** $2\ NO + 2\ H_2 \rightarrow N_2 + 2\ H_2O$**.**

CHEMICAL REACTION RATE



Potassium reacts violently with water. That's what is happening in the beaker pictured above. Why does potassium have such explosive reactions? It's because the reactions occur so quickly.

How Fast Does It Go?

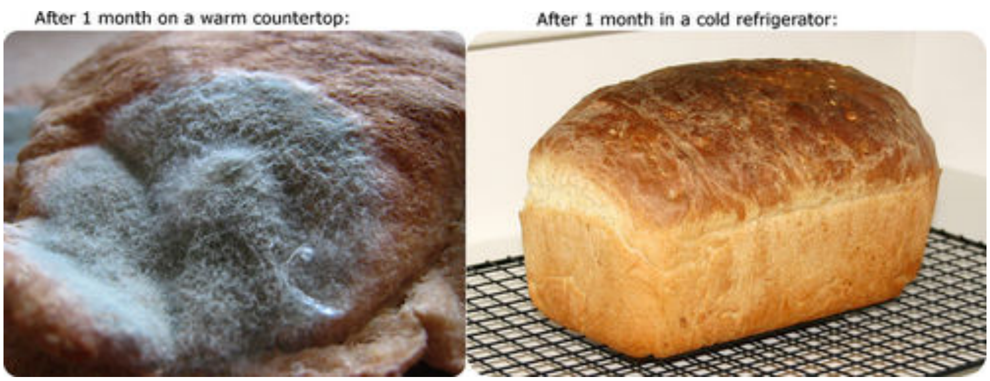
How fast a chemical reaction occurs is called the **reaction rate**. Several factors affect the rate of a given chemical reaction. They include the following:

- temperature of reactants.
- concentration of reactants.
- surface area of reactants.
- presence of a catalyst.

At the following URL, you can see animations showing how these factors affect the rate of chemical reactions.
http://www.kentchemistry.com/links/Kinetics/FactorsAffecting.htm

Temperature of Reactants

When the temperature of reactants increases, the rate of the reaction increases. At higher temperatures, particles of reactants have more energy, so they move faster. As a result, they are more likely to bump into one another and to collide with greater force. For example, food spoils because of chemical reactions, and these reactions occur faster at higher temperatures (see the bread on the left in the Figure below). This is why we store foods in the refrigerator or freezer (like the bread on the right below in the Figure below). The lower temperature slows the rate of spoilage.



Why does increasing the temperature increase the rate of a reaction? When the temperature is increased, the average velocity of the particles is increased. As a result, the average kinetic energy of these particles is also increased. The result is that the particles will collide more frequently because the particles move around faster and will encounter more reactant particles, but this is only a minor part of the reason why the rate is increased. Just because the particles are colliding more frequently does not mean that the reaction will definitely occur.

The major effect of increasing the temperature is that more of the particles that collide will have the amount of energy needed to have an effective collision, or a collision that results in a reaction. In other words, more particles will have the energy needed in order to react.

Concentration of Reactants

Concentration is the number of particles of a substance in a given volume. When the concentration of reactants is higher, the reaction rate is faster. At higher concentrations, particles of reactants are crowded closer together, so they are more likely to collide and react. Did you ever see a sign like the one in the Figure below? You might see it where someone is using a tank of pure oxygen for a breathing problem. Combustion, or burning, is a chemical reaction in which oxygen is a reactant. A greater concentration of oxygen in the air makes combustion more rapid if a fire starts burning.

Q: It is dangerous to smoke or use open flames when oxygen is in use. Can you explain why?

A: Because of the higher-than-normal concentration of oxygen, the flame of a match, lighter, or cigarette could spread quickly to other materials or even cause an explosion.

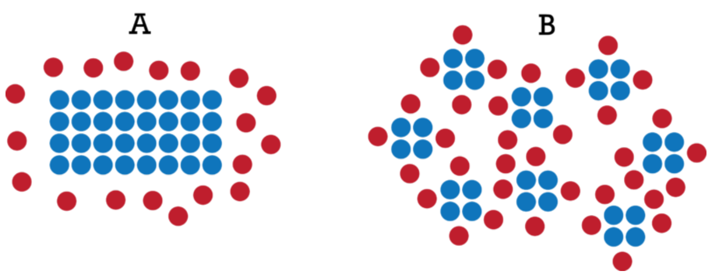
If you had one red ball and one green ball flying around randomly in an enclosed space and undergoing perfectly elastic collisions with the walls and with each other, in a given amount of time, the balls would collide with each other a certain number of times as determined by probability. If you now put two red balls and one green ball in the room under the same conditions, the probability of a collision between a red ball and the green ball would exactly double. The green ball would have twice the chance of encountering a red ball in the same amount of time.

In terms of chemical reactions, a similar situation exists. Particles of two gaseous reactants or two reactants in solution have a certain probability of undergoing collisions with each other in a reaction vessel. If you double the concentration of either reactant, the probability of a collision doubles. The rate of reaction is proportional to the number of collisions per unit time. Assuming that the percent of successful collision does not change, then having twice as many collisions will result in twice as many successful collisions. The rate of reaction is proportional to the number of collisions per unit time, so increasing the concentration of either reactant increases the number of collisions, the number of successful collisions, and the reaction rate.

Surface Area of Reactants

When a solid substance is involved in a chemical reaction, only the matter at the surface of the solid is exposed to other reactants. If a solid has more surface area, more of it is exposed and able to react. Therefore, increasing the surface area of solid reactants increases the reaction rate. Look at the hammer and nails pictured in the Figure below. Both are made of iron and will rust when the iron combines with oxygen in the air. However, the nails have a greater surface area, so they will rust faster.

Consider a reaction between reactant red and reactant blue, where reactant blue is in the form of a single lump (Figure A below). Then compare this to the same reaction where reactant blue has been broken up into many smaller pieces (Figure B below).



In Figure A, only the blue particles on the outside surface of the lump are available for collision with reactant red. The blue particles on the interior of the lump are protected by the blue particles on the surface. If you count the number of blue particles available for collision, you will find that only 20 blue particles could be struck by a particle of reactant red. In Figure B, however, the lump has been broken up into smaller pieces, and all the interior blue particles are now on a surface and available for collision. As a result, more collisions between blue and red will occur. The reaction in Figure B will occur at faster rate than the same reaction in Figure A. Increasing the surface area of a reactant increases the frequency of collisions and increases the reaction rate.

Summary

- How fast a chemical reaction occurs is called the reaction rate.
- Several factors affect the rate of a chemical reaction, including the temperature, concentration, and surface area of reactants, and the presence of a catalyst.

Vocabulary

- catalyst: Substance that increases the rate of a chemical reaction but is not changed or used up in the reaction.
- reaction rate: Speed at which a chemical reaction occurs.

Practice

Watch the video about reaction rate at the following URL, and then answer the questions below.

http://www.youtube.com/watch?feature=endscreen&NR=1&v=F-Cu_AoWOK4

1. What is collision theory?
2. How does collision theory relate to factors that affect reaction rate?

Review

3. Define reaction rate.
4. List factors that influence the rate of a chemical reaction.
5. Choose one of the factors you listed in your answer to question 2, and explain how it affects reaction rate.

ACTIVATION ENERGY

Cheerleaders have a lot of energy! Their role is to get the fans as excited as they are so everyone will cheer for the team. They use their energy to activate the crowd.

Getting Started

Chemical reactions also need energy to be activated. They require a certain amount of energy just to get started. This energy is called activation energy. For example, activation energy is needed to start a car engine. Turning the key causes a spark that activates the burning of gasoline in the engine. The combustion of gas won't occur without the spark of energy to begin the reaction.

Q: Why is activation energy needed? Why won't a reaction occur without it?

A: A reaction won't occur unless atoms or molecules of reactants come together. This happens only if the particles are moving, and movement takes energy. Often, reactants have to overcome forces that push them apart. This takes energy as well. Still more energy is needed to start breaking bonds in reactants.

SUMMARY

- All chemical reactions, including exothermic reactions, need activation energy to get started.
- Activation energy is needed so reactants can move together, overcome forces of repulsion, and start breaking bonds.
- activation energy: Energy needed to start a chemical reaction.

PRACTICE

Watch the animated reactions at the following URL, and then answer the question below.
http://www.youtube.com/watch?v=VbIaK6PLrRM

1. What is the exothermic reaction in the video? What role does the flame play in this reaction?
2. The exothermic reaction releases energy. What is this energy used for?
3. What is the endothermic reaction in the video? What provides the activation energy for this reaction? What product results from this reaction?

REVIEW

4. What is activation energy?
5. Why do reactants need energy in order for a chemical reaction to begin?
6. Make an original graph to represent the role of energy, including activation energy, in an endothermic reaction.

CATALYSTS

The tunnel through this mountain provides a faster route for cars to get to the other side of the mountain. If a chemical reaction were like a road to the other side of a mountain, a catalyst would be like a tunnel.

What Is a Catalyst?

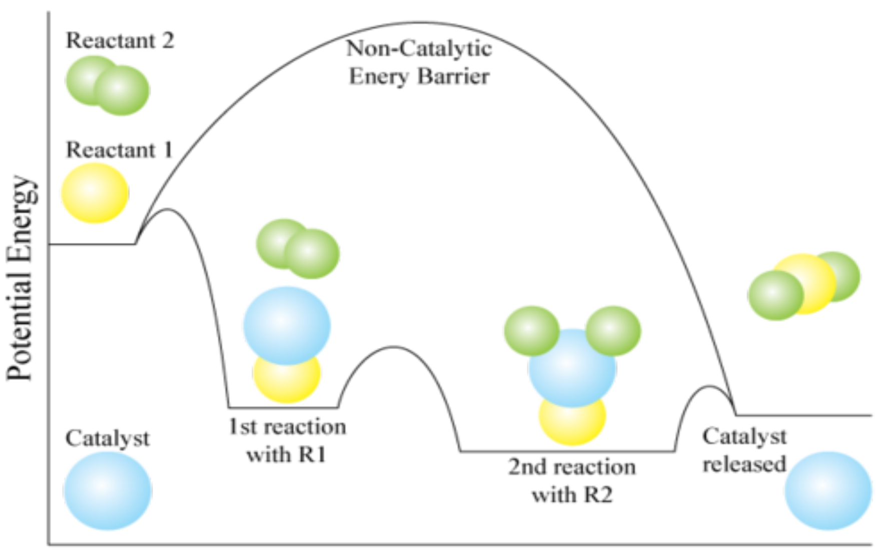
A catalyst is a substance that increases the rate of a chemical reaction. The presence of a catalyst is one of several factors that influence the rate of chemical reactions. (Other factors include the temperature, concentration, and surface area of reactants.) A catalyst isn't a reactant in the chemical reaction it speeds up. As a result, it isn't changed or used up in the reaction, so it can go on to catalyze many more reactions.

Q: How is a catalyst like a tunnel through a mountain?
A: Like a tunnel through a mountain, a catalyst provides a faster pathway for a chemical reaction to occur.

How Catalysts Work

Catalysts interact with reactants so the reaction can occur by an alternate pathway that has a lower activation energy. Activation energy is the energy needed to start a reaction. When activation energy is lower, more reactant particles have enough energy to react so the reaction goes faster. Many catalysts work like the one in the diagram below. (You can see an animated version at the following URL.) The catalyst brings the reactants together by temporarily bonding with them. This makes it easier and quicker for the reactants to react together. Notice how the catalyst is released by the product molecule at the end of the reaction.

http://www.saskschools.ca/curr_content/chem30/modules/module4/lesson5/explainingcatalysts.htm

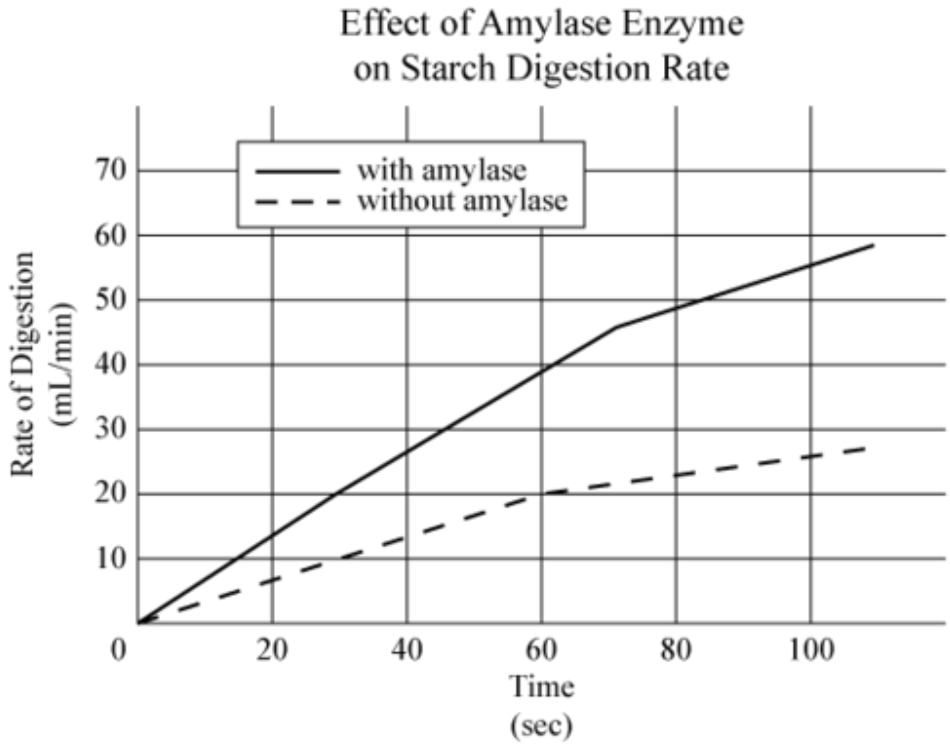


Q: In the diagram above, look at the energy needed in the catalytic and non-catalytic pathways of the reaction. How does the amount of energy compare? How does this affect the reaction rate along each pathway?

A: The catalytic pathway of the reaction requires far less energy. Therefore, the reaction will occur faster by this pathway because more reactants will have enough energy to react.

Catalysts in Living Things

Chemical reactions constantly occur inside living things. Many of these reactions require catalysts so they will occur quickly enough to support life. Catalysts in living things are called enzymes. Enzymes may be extremely effective. A reaction that takes a split second to occur with an enzyme might take many years without it!



More than 1000 different enzymes are necessary for human life. Many enzymes are needed for the digestion of food. An example is amylase, which is found in the mouth and small intestine. Amylase catalyzes the breakdown of starch to sugar. You can see how it affects the rate of starch digestion in the graph below.

Q: If you chew a starchy food such as a soda cracker for a couple of minutes, you may notice that it starts to taste slightly sweet. Why does this happen?

A: The starches in the cracker start to break down to sugars with the help of the enzyme amylase. Try this yourself and see if you can taste the reaction.

SUMMARY

- A catalyst is a substance that increases the rate of a chemical reaction.
- A catalyst provides an alternate pathway for the reaction that has a lower activation energy. When activation energy is lower, more reactant particles have enough energy to react, so the reaction occurs faster.
- Chemical reactions constantly occur inside living things, and many of them require catalysts to occur quickly enough to support life. Catalysts in living things are called enzymes.

VOCABULARY

- catalyst: Substance that increases the rate of a chemical reaction but is not changed or used up in the reaction.

PRACTICE

At the following URL, watch the video showing a chemical reaction both with and without a catalyst. Then answer the questions below.
http://www.youtube.com/watch?v=GSCJ2EYkRqQ

1. Write the chemical equation for the reaction that is demonstrated in the video.
2. What chemical is used to catalyze the reaction?
3. Describe two observations that provide evidence that the reaction has occurred after the addition of the catalyst.

REVIEW

4. What is a catalyst?
5. How does a catalyst speed up a chemical reaction?
6. What are enzymes? Why are they important?

STANDARD 5, OBJECTIVE 2: RECOGNIZE THAT CERTAIN REACTIONS DO NOT CONVERT ALL REACTANTS TO PRODUCTS, BUT ACHIEVE A STATE OF DYNAMIC EQUILIBRIUM THAT CAN BE CHANGED.

Vocabulary

- dynamic equilibrium
- equilibrium
- irreversible reaction
- macroscopic properties
- reversible reaction

Lesson Objectives

- describe what is happening in a system at equilibrium.
- explain what is meant by dynamic equilibrium.
- state the conditions necessary for a system to be in equilibrium.

INTRODUCTION TO EQUILIBRIUM

Consider this generic reaction: $A + B \rightarrow C + D$. Based on what we have learned so far, you might assume that the reaction will keep going forward, forming C and D until either A or B (or both) is completely used up. When this is the case, we would say that the reaction "goes to completion." Reactions that go to completion are referred to as irreversible reactions.

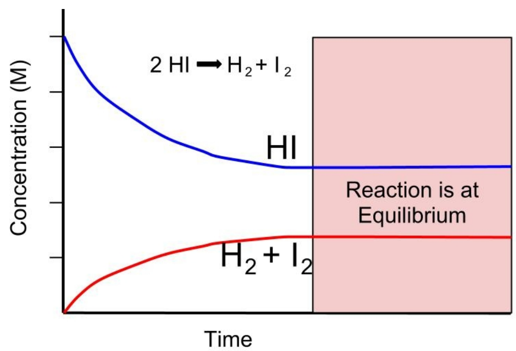
Some reactions, however, are reversible, meaning that products can also react to re-form the reactants. In our example, this would correspond to the reaction $C + D \rightarrow A + B$. During a reversible reaction, both the forward and backward reactions are happening at the same time.

Equilibrium

As we learned earlier, the rate of a reaction depends on the concentration of the reactants. At the very beginning of the reaction $A + B \rightarrow C + D$, we would not expect the reverse reaction to proceed very quickly. If only a few particles of C and D have been created in a large flask of A and B, then it is very unlikely that they will find each other because the concentration of C and D is just too low. If C and D cannot find each other and collide with the correct energy and orientation, no reaction will occur.

However, as more and more C and D are created, it becomes more and more likely that they will find each other and react to re-form A and B. Conversely, as A and B are being used up, the forward reaction slows down for the same exact reason. The concentration of A and B decreases over the course of a reaction because there are less A and B particles in the same size flask. At some point, the rates for the forward and reverse reactions will be equal, at which point the concentrations will no longer change. If A and B are being destroyed at the same rate that they are being created, the overall amount should not change over time. At this point, the system is said to be in **equilibrium**. A qualitative description of this process for the reaction between hydrogen and iodine to make hydrogen iodide is shown below.

Chemists use a double arrow to show that a reaction is in equilibrium. For the reaction above, the chemical equation would be:
$H_2 + I_2 \rightleftharpoons 2HI$



This indicates that both directions of the reaction are occurring. Note that a double-headed arrow ($\leftrightarrow$) should not be used here because this has a different chemical meaning.

Dynamic Equilibrium

One characteristic of a system at equilibrium is that the macroscopic properties do not change over time. Macroscopic properties are those that describe the system as a whole. These can be observed and measured without needing to investigate the system in terms of individual molecules.

One example of a macroscopic property is temperature. Although temperature can be explained on a microscopic level by the speed at which particles are moving around, individual molecules do not have a "temperature." A glass of water with a constant temperature contains molecules with a large range of different speeds, and the speed of any individual molecule is constantly changing as it collides with other molecules and gains or loses kinetic energy. However, if you average out the properties of all these molecules (7.91×10^{24}) for an 8 oz glass of water, the overall temperature appears constant, despite continuous changes on the microscopic level.

Concentration is another example of a macroscopic property. It describes the total amount of a substance, but it does not make any statement about what is occurring on a molecular level. When a reaction is at equilibrium, the concentration of each component is constant over time. As we saw before, both the forward and reverse reactions are still taking place, but since they are moving at the same rate, there is no change for the system as a whole. The term dynamic equilibrium refers to a state where no net change is taking place, despite the fact that both reactions are still occurring. Individual molecules are still being formed and broken down, but the system as a whole is not changing over time.

Other Conditions Necessary for Equilibrium

In order for a reaction to reach an equilibrium state, it needs to take place in a closed system. Consider the following reaction:

$$H_2CO_{3(aq)} \rightleftharpoons H_2O_{(l)} + CO_{2(g)}$$

This equilibrium is established in a closed can of soda because neither the reactants nor the products can leave the system. However, when you open the soda, one of the products (CO_2) is free to escape into the atmosphere. If a molecule of CO_2 escapes, it is no longer available to collide with the water molecules left behind, so it can no longer participate in the reverse reaction. Therefore, the forward reaction will eventually go to completion until there is no more H_2CO_3 available. Note that since H_2CO_3 is an acid, an open soda will become less acidic over time. Acidity is an important component of taste, so flat soda really does taste different for reasons other than texture.

Another requirement for a system to stay in equilibrium is that the temperature and pressure stay constant. As we learned in the previous chapter, temperature and pressure both have an effect on the rate of a reaction. However, the effect might be greater for the forward reaction than the reverse reaction, or vice versa. For example, in the reaction above, adding heat will favor the forward reaction more than the reverse, resulting in the production of more CO_2. This is true even if the can remains closed. Eventually, a new equilibrium will be established at the new temperature, but while the temperature is changing, the system is no longer in equilibrium.

LESSON SUMMARY

- Irreversible reactions will continue to form products until the reactants are fully consumed.
- Reversible reactions will react until a state of equilibrium is reached.
- A reaction is at equilibrium when there is no net change to the system over time.
- Dynamic equilibrium refers to an equilibrium where forward and reverse reactions are still occurring, but they are proceeding at the same rate, so there is no net change.

REVIEW QUESTIONS

1. What does the term dynamic equilibrium mean?
2. List all of the conditions for a dynamic equilibrium.
3. Of the following conditions, which is not required for a dynamic equilibrium?
   a) rate of the forward reaction equals the rate of the reverse reaction.
   b) reaction occurs in an open system
   c) reaction occurs at a constant temperature
   d) reaction occurs in a closed system

4. Which of the following systems, at room temperature and pressure, can be described as a dynamic equilibrium?
   a) an open flask containing air, water and water vapor
   b) a glass of water containing ice cube cubes and cold water
   c) a closed bottle of soda pop
   d) an open flask containing solid naphthalene

5. Is each of the following in a state of equilibrium? Explain.
   a) Ice cubes are melting in a glass of water with a lid on it
   b) Crystals of potassium dichromate were dissolved in water, and now the water is a uniform orange color with a small amount of crystal left in the closed container.
   c) An apple that is left on the counter for a few days, it dries out and turns brown.

6. If the following table (Table below) of concentration vs. time was provided to you for the ionization of acetic acid, how would you know when equilibrium was reached?

Data Table For Problem 6

Time (min)	$[HC_2H_3O_2]$ mol/L
0	0.100
0.5	0.099
1.0	0.098
1.5	0.097
2.0	0.096
2.5	0.095
3.0	0.095
3.5	0.095
4.0	0.095
4.5	0.095
5.0	0.095

TYPES OF EQUILIBRIUM

Lesson Objectives

- demonstrate on specified chemical reactions how Le Châtelier's Principle is applied to equilibrium systems.
- describe the effect of concentration on an equilibrium system.
- describe the effect of pressure as a stress on the equilibrium position.
- describe the effect of temperature as a stress on an equilibrium system.

Vocabulary

- Le Châtelier's Principle

Introduction

When a reaction has reached equilibrium with a given set of conditions, if the conditions are not changed, the reaction will remain at equilibrium forever. The forward and reverse reactions continue at the same equal and opposite rates, and the macroscopic properties remain constant.

It is possible, however, to alter the reaction conditions. For example, you could increase the concentration of one of the products, or decrease the concentration of one of the reactants, or change the temperature. When a change of this type is made in a reaction at equilibrium, the reaction is no longer in equilibrium. When you alter something in a reaction at equilibrium, chemists say that you put stress on the equilibrium. When this occurs, the reaction will no longer be in equilibrium, so the reaction itself will begin changing the concentrations of reactants and products until the reaction comes to a new position of equilibrium. How a reaction will change when a stress is applied can be explained and predicted and is the topic of this lesson.

Le Châtelier's Principle

In the late 1800s, a chemist by the name of Henry-Louis Le Châtelier was studying stresses that were applied to chemical equilibria. He formulated a principle, Le Châtelier's Principle, which states that when a stress is applied to a system at equilibrium, the equilibrium will shift in a direction to partially counteract the stress and once again reach equilibrium. For instance, if a stress is applied by increasing the concentration of a reactant, the equilibrium position will shift toward the right and remove that stress by using up some of the reactants. The reverse is also true. If a stress is applied by lowering a reactant concentration, the equilibrium position will shift toward the left, this time producing more reactants to partially counteracting that stress. The same reasoning can be applied when some of the products is increased or decreased.

Le Châtelier's principle does not provide an explanation of what happens on the molecular level to cause the equilibrium shift. Instead, it is simply a quick way to determine which way the reaction will run in response to a stress applied to the system at equilibrium.

Effect of Concentration Changes

Let's use Le Châtelier's principle to explain the effect of concentration changes on an equilibrium system. Consider the generic equation:

$$A_{(aq)} + B_{(aq)} \rightleftharpoons C_{(aq)} + D_{(aq)}$$

At equilibrium, the forward and reverse rates are equal. The concentrations of all reactants and products remain constant, which keeps the rates constant. Suppose we add some additional A, thus raising the concentration of A without changing anything else in the system. Since the concentration of A is larger than it was before, the forward reaction rate will suddenly be higher. The forward rate will now exceed the reverse rate. Now there is a net movement of material from the reactants to the products. As the reaction uses up reactants, the forward rate that was too high slowly decreases while the reverse rate that was too low slowly increases. The two rates are moving toward each other and will eventually become equal again. They do not return to their previous rates, but they do become equal at some other value. As a result, the system returns to equilibrium.

Le Châtelier's principle says that when you apply a stress (adding A), the equilibrium system will shift to partially counteract the applied stress. In this case, the reaction shifts toward the products so that A and $B$are used up and $C$and $D$are produced. This reduction of the concentration of A is counteracting the stress you applied (adding A). Suppose instead that you removed some A instead of adding some. In that case, the concentration of A would decrease, and the forward rate would slow down. Once again, the two rates are no longer equal. At the instant you remove A, the forward rate decreases, but the reverse rate remains exactly what it was. The reverse rate is now greater than the forward rate, and the equilibrium will shift toward the reactants. As the reaction runs backward, the concentrations of $C$and Ddecrease, slowing the reverse rate, and the concentrations of A and B increase, raising the forward rate. The rates are again moving toward each other, and the system will again reach equilibrium. The shift of material from products to reactants increases the concentration of A, thus counteracting the stress you applied. Le Châtelier's principle again correctly predicts the equilibrium shift.

The effect of concentration on the equilibrium system according to Le Châtelier is as follows: increasing the concentration of a reactant causes the equilibrium to shift to the right, using up reactants and producing more products. Increasing the concentration of a product causes the equilibrium to shift to the left, using up products and producing more reactants. The exact opposite is true when either a reactant or product is removed.

Example:

$$\mathrm{SiCl}_{4(g)} + \mathrm{O}_{2(g)} \rightleftharpoons \mathrm{SiO}_{2(s)} + 2\ \mathrm{Cl}_{2(g)}$$

For the reaction, what would be the effect on the equilibrium system if:

1. $[SiCl_4]$ increases
2. $[O_2]$ increases
3. $[Cl_2]$ increases

Solution:

1. The equilibrium would shift to the right. $[Cl_2]$ would increase, more SiO_2 would be produced (but that does not increase its concentration since its a solid), and $[O_2]$ would decrease.
2. The equilibrium would shift to the right. $[SiCl_4]$ would decrease, more SiO_2 would be produced (but again no change in concentration), and $[Cl_2]$ would increase.
3. The equilibrium would shift left. $[SiCl_4]$ and $[O_2]$ would increase, and SiO_2 would be used up but not change its concentration.

Let's take a moment to consider what happens to the concentration of a reactant or product that is changed. In our theoretical reaction, if you add A, the concentration of A will increase. The equilibrium shifts toward the products, and A is used. Where does the concentration of A end up, higher or lower than the original concentration? The concentration of A increases when you add more A, but it decreases as the equilibrium shifts. A new equilibrium, however, will be reached before the concentration of A gets back down to its original concentration. This is why Le Châtelier's principle says the equilibrium will shift to partially counteract the applied stress. The equilibrium shift will move toward returning the concentration to where it was before you applied the stress, but the concentration never quite gets back to the original value before a new equilibrium is established.

Example:

For the reaction $PCl_{3(g)} + Cl_{2(g)} \rightleftharpoons PCl_{5(g)}$, which way will the equilibrium shift if:

1. $[PCl_3]$ decreases
2. $[Cl_2]$ decreases
3. $[PCl_5]$ decreases

Solution:

1. left
2. left
3. right

Example:

Here's a reaction at equilibrium. Note the phases of each reactant and product.

$$A_{(aq)} + B_{(s)} \rightleftharpoons C_{(aq)} + D_{(aq)}$$

1. Which way will the equilibrium shift if you add some A to the system without changing anything else?
2. After A has been added and a new equilibrium is reached, how will the new concentration of D compare to the original concentration of D?
3. After A has been added and a new equilibrium has been established, how will the new concentration of A compare to the original concentration of A?
4. After A has been added and a new equilibrium has been established, how will the new concentration of B compare to the original concentration of B?
5. Which way will the equilibrium shift if you add some C to the system without changing anything else?
6. After C has been added and a new equilibrium has been established, how will the new concentration of D compare its original concentration?
7. After C has been added and a new equilibrium has been established, how will the new concentration of A compare its original concentration?
8. Which way will the equilibrium shift if you add some B to the system without changing anything else?

Solution:

1. The equilibrium will shift toward the products.

2. The new concentration of D will be higher than the original.
3. The new concentration of A will be higher than the original, but lower than the concentration right after A was added.
4. Since B is a solid, its concentration will be the same as the original. There will be less of it since some was used in the equilibrium shift, but the concentration will be the same.
5. The equilibrium will shift toward the reactants.
6. The new concentration of D will be lower than the original.
7. The new concentration of A will be higher than the original.
8. Since B is a solid, adding B will not change its concentration and therefore has no effect on the equilibrium. It is possible that adding some B will increase the surface area of B and therefore increase the forward reaction rate, but it will also increase the reverse reaction rate by approximately the same amount, hence no shift in equilibrium.

The Haber Process and Concentration Change Effects

The reaction between nitrogen gas and hydrogen gas can produce ammonia, NH_3. However, under normal conditions, this reaction does not produce very much ammonia. Early in the 20th century, the commercial use of this reaction was too expensive because of the low yield.

$$N_{2(g)} + 3\ H_{2(g)} \rightleftharpoons 2\ NH_{3(g)} + \text{energy}$$

A German chemist named Fritz Haber applied Le Châtelier's principle to help solve this problem. Decreasing the concentration of ammonia by immediately removing it from the reaction container causes the equilibrium to shift to the right, so the reaction can continue to produce more product. Haber used Le Châtelier's principle to solve this problem in other ways as well. These will be discussed in the following sections.

SUMMARY

- Increasing the concentration of a reactant causes the equilibrium to shift to the right, producing more products.
- Increasing the concentration of a product causes the equilibrium to shift to the left, producing more reactants.
- Decreasing the concentration of a reactant causes the equilibrium to shift to the left, using up some products.
- Decreasing the concentration of a product causes the equilibrium to shift to the right, using up some reactants.

QUESTIONS

1. What is the effect on the equilibrium position if the [reactants] is increased?
2. What is the effect on the equilibrium position if the [reactants] is decreased?
3. Which of the following will cause a shift in the equilibrium position of the equation: ? i. add C_8H_{18} Ii. add O_2 iii. remove CO_2 iv. remove H_2O

   a. i and ii only
   b. ii and iii only
   c. ii and iv only
   d. i, ii, iii, and iv

4. Which of the following will cause a shift in the equilibrium position of the equation: $CaCO_{3(s)} \rightleftharpoons CaO_{(s)} + CO_{2(g)}$? i. add $CaCO_3$ ii. add CO_2 iii. remove CaO iv. remove CO_2

   a. ii only
   b. i and iii only
   c. ii and iv only
   d. i, ii, iii, and iv only

5. For the reaction: $N_2O_{5(s)} \rightleftharpoons NO_{2(g)} + O_{2(g)}$, what would be the effect on the equilibrium if:

   a. N_2O_5 is added?
   b. NO_2 is removed?
   c. NO_2 is added?
   d. O_2 is added?

6. **Answer the following questions when $[\mathrm{CO}]$ is increased in the following system at equilibrium: .**

    a. Write the equilibrium constant expression.
    b. Which direction will this equilibrium shift?
    c. What effect will this stress have on $[\mathrm{CO_2}]$?

7. **For the reaction , what would be the effect on the equilibrium system if:**

$$\mathrm{C}_{(s)} + \mathrm{H_2O}_{(g)} \rightleftharpoons \mathrm{CO}_{(g)} + \mathrm{H}_{2(g)}$$

    a. $[\mathrm{H_2O}]$ increases?
    b. mass of C decreases?
    c. $[\mathrm{CO}]$ increases?
    d. $[\mathrm{H_2}]$ decreases?

Effect of Changes of Pressure

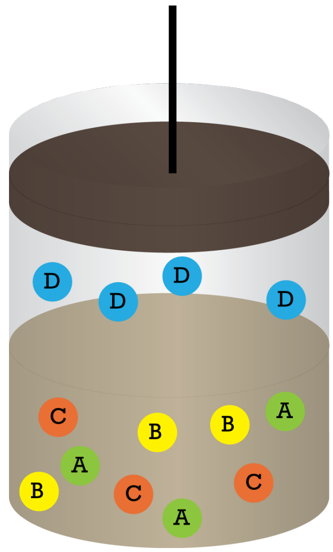
A second type of stress studied by Le Châtelier was the effect of changing the pressure of a system at equilibrium. Reactants and products that are in solid, liquid, or aqueous states are *not* compressible, so raising or lowering the pressure on a system that contains no gases will have no effect on the equilibrium. If, however, the reaction at equilibrium contains at least one gas in either the reactants or products, altering the pressure will alter the equilibrium.

Suppose we have the following reaction at equilibrium.

$$A_{(aq)} + B_{(aq)} \rightleftharpoons C_{(aq)} + D_{(g)}$$

Since the reaction has at least one gaseous substance involved in the reaction, its equilibrium will be affected by a change in the partial pressure of that gas. In order for this reaction to reach equilibrium, it would have to be in a closed reaction vessel so that the gaseous product do not escape. The gas must stay in contact with the solution for the reverse reaction to occur.

There are at least three ways to increase the pressure in the area above the liquid in the cylinder: 1) add some other gas not involved in the reaction, 2) add some gaseous D into the space above the liquid, and 3) push the piston down so the space above the liquid in the cylinder is decreased (see figure below).



As we double the forward reaction rate and quadruple the reverse rate, the equilibrium shift will be toward the reactants.

- Increasing the pressure in the cylinder by adding a gas not involved in the reaction will increase the total gas pressure in the cylinder, but it will not affect the partial pressure of D, so there will be no effect on the equilibrium of this reaction. This concept might be easier to understand if we think about it in terms of concentration rather than pressure. If we have the same amount of gas in the same amount of space, the concentration of the gas and thus its partial pressure will not change. The fact that some other gas was added does not change either of these key values. This is why only changes to the partial pressure of the reaction component, not the total pressure, are able to affect the equilibrium position.
- Adding some gaseous D to the cylinder is the same as adding a reactant or product, which has already been discussed earlier.
- Lowering the piston in the cylinder pushes the D molecules into a smaller space, increasing both the partial pressure of gaseous D and its concentration in the space above the liquid. Since the concentration of D is increased, the reverse reaction rate will increase and the equilibrium will shift toward the reactants.

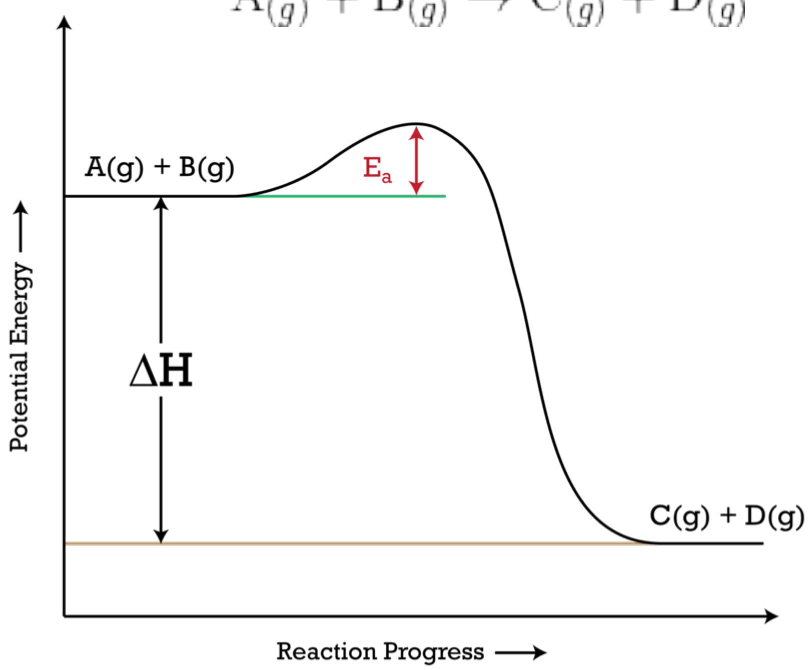
The reverse of this is also true. If you expand the volume of the cylinder by raising the piston, the partial pressure and concentration of D will decrease. When the concentration of D decreases, the reverse rate slows and the forward reaction rate will drive the equilibrium toward the products. This equilibrium shift increases the concentration and partial pressure of D, once again counteracting the stress you applied.

Effect of Changing Temperature

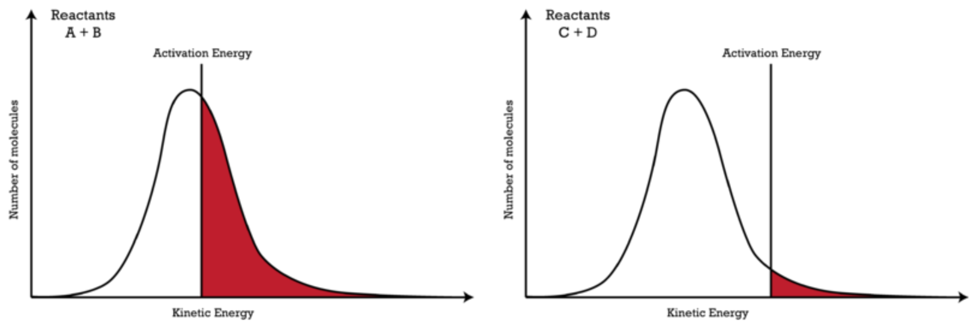
One of the most important factors that determines reaction rate is temperature. Raising the temperature will increase the average speed of the individual particles, thus causing more frequent collisions. Additionally, this increase in energy means that more particles will have the energy necessary to overcome the activation barrier. Overall, a rise in temperature increases both the frequency of collisions and the percentage of successful collisions.

It should be clear that increasing the temperature of the reaction vessel will increase both the forward and reverse reaction rates, but will it increase both rates equally? Let's examine the potential energy diagram of a reaction to see if we can gain any insight there. Here is the potential energy diagram for our usual theoretical reaction:

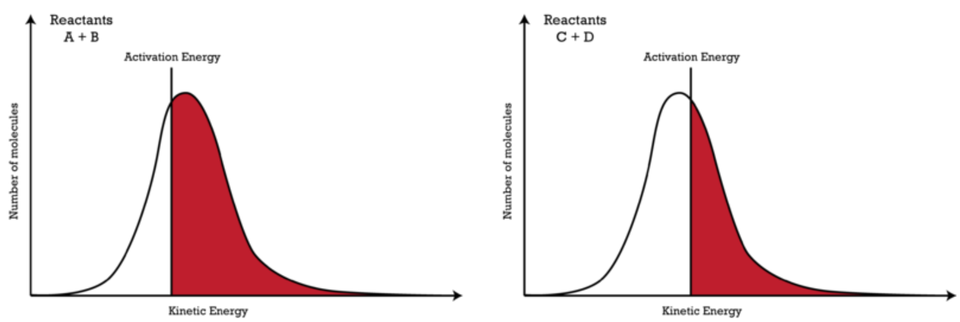
$$A_{(g)} + B_{(g)} \rightarrow C_{(g)} + D_{(g)}$$



As you can see, the forward reaction has a small energy barrier while the reverse reaction has a very large energy barrier. With the reactants and products at the same temperature, the forward reaction will be much faster that the reverse reaction if the concentration of reactants is equal to the concentration of products. In the chapter “Chemical Kinetics,” we used an energy distribution curve to show the percentage of reactant particles that had sufficient activation energy to react. The figure below shows the energy distribution curves for A and B in the forward reaction and for C and D in the reverse reaction.



As you can see in the drawing, a much larger percentage of reactant particles than product particles have the activation energy required for a successful collision. Suppose we were to increase the temperature $10°C$. The activation energy requirements remain the same for both groups of particles, but the curves will shift right to reflect the additional energy these particles possess.



Both reaction rates will increase because a higher percentage of both reactants and products have enough energy to overcome the activation barrier. Additionally, the percentage of particles that exceed the activation energy is still higher for the reactants than the products, so the forward reaction will still be faster. However, if we look at the increase in sufficiently energetic particles, the product side has more of a change than the reactants. Thus, the reverse rate will increase more than the forward rate. The equilibrium will shift left, producing more reactants until a new equilibrium is established.

For nearly every reaction, either the forward or reverse reaction will require more activation energy than the other. The addition of energy to a reaction will increase both reaction rates, but it increases the rate of the slower reaction more. In an exothermic reaction, the products are lower in energy than the reactants, so they have a higher barrier to climb in order to reach the same transition state. Accordingly, the reverse rate will be slower for an exothermic reaction, so increasing the temperature will shift the equilibrium to the left. The reverse is true for an endothermic reaction. By looking at a potential energy diagram, you should be able to tell 1) whether the reaction is exothermic or endothermic, 2) whether the forward or reverse reaction would be slower, assuming equal concentrations of reactants and products, and 3) which direction the equilibrium would shift in respond to a change in temperature.

Following the same reasoning as above, we can see that decreasing the temperature of a reaction produces an equilibrium shift in the opposite direction. Cooling an exothermic reaction results in a shift to the right, and cooling an endothermic reaction causes a shift to the left. Le Châtelier's principle correctly predicts the equilibrium shift when systems are heated or cooled. An increase in temperature is the same as adding energy to the system. Look at the following reaction:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \quad \Delta H = -191\ kJ$$

This could also be written as:

$$2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons 2\ SO_{3(g)} + 191\ kJ$$

When changing the temperature of a system at equilibrium, energy can be thought of as just another product or reactant. For this reaction, 191kJ of energy is produced for every mole of O_2 and 2 moles of SO_2 that react. Therefore, when the temperature of this system is raised, the effect will be the same as increasing any other product. As the temperature is increased, the equilibrium will shift away from the stress, resulting in more reactants and less products. As you would expect, the reverse would be true if the temperature is decreased. A summary of the effect temperature has on equilibrium systems is shown in Table below.

The Effect of Temperature on an Endothermic and an Exothermic Equilibrium System

	Exothermic $(-\triangle H)$	Endothermic $(+\triangle H)$
Increase Temperature	Shifts left, favors reactants	Shifts right, favors products
Decrease Temperature	Shifts right, favors products	Shifts left, favors reactants

Example:
Predict the effect on the equilibrium position if the temperature is increased in each of the following:

1) $H_{2(g)} + CO_{2(g)} \rightleftharpoons CO(g) + H_2O(g) \quad \Delta H = +40\ kJ$

2) $2\ SO_{2(g)} + O_{2(g)} \rightleftharpoons SO_{3(g)} + \text{energy}$

Solution:
The reaction is endothermic. With an increase in temperature for an endothermic reaction, the reaction will shift right, producing more products.
The reaction is exothermic. With an increase in temperature for an exothermic reaction, the reaction will shift left, producing more reactants.

The Haber Process and the Effect of Temperature Change

We have already seen that the maximum amount of ammonia produced in the Haber process can be improved by decreasing the concentration of the ammonia (by removing it from the reaction container) and increasing the pressure.

$$N_{2(g)} + 3\ H_{2(g)} \rightleftharpoons 2\ NH_{3(g)} + \text{energy}$$

One more factor that will affect this equilibrium system is the temperature. Since the forward reaction is exothermic, lowering the temperature will once again shift the equilibrium system to the right and increase the ammonia produced. Unfortunately, this process also has a very high activation energy, so if the temperature is too low, the reaction will slow to a crawl. Thus, a balance must be struck between shifting the equilibrium to favor products and allowing products to be formed at a reasonable rate. It was found that the optimum conditions for this process (the ones that produce the most ammonia the fastest) are 550°C and 250 atm of pressure, with the ammonia being continually removed from the system.

SUMMARY

- For an endothermic reaction, an increase in temperature shifts the equilibrium toward the products, whereas a decrease in temperature shifts the equilibrium toward the reactants.
- For an forward exothermic reaction, an increase in temperature shifts the equilibrium toward the reactant side, whereas a decrease in temperature shifts the equilibrium toward the products.
- Increasing or decreasing the temperature causes the value to change.

QUESTIONS

1. Which direction will a system at equilibrium shift for each of the following?

    a) adding energy to a forward exothermic equilibrium system
    b) adding energy to a reverse exothermic equilibrium system
    c) adding energy to a forward endothermic equilibrium system
    d) adding energy to a reverse endothermic equilibrium system

2. Predict the effect of an increase in temperature on the equilibrium position for each of the following.

    a) $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g) \quad \Delta H = +51.8 kJ$
    b) $P_4O_{10}(s) + 6\ H_2O(l) \rightleftharpoons 4\ H_3PO_4(aq)$ + heat

3. How does an increase in temperature affect the value of and the concentration of the products for each of the following?

    a) $NO_2(g) + NO(g) \rightleftharpoons N_2O(g) + O_2(g) \quad \Delta H = -43\ kJ$
    b) $4\ NH_3(g) + 5\ O_2(g) \rightleftharpoons 4\ NO(g) + 6\ H_2O(g)$ + heat

4. Predict the effect on the equilibrium position if the temperature is increased in each of the following.

a) $C_2H_2(g) + H_2O(g) \rightleftharpoons CH_3CHO(g)$ $\Delta H = -151$ kJ/mol
b) $HC_2H_3O_2(g) + H_2O(g) + \text{energy} \rightleftharpoons CH_3CH_2OH(g) + O_2(g)$

Use this information to answer questions 5 – 8. The reaction below is at equilibrium in a closed container at 25°C.

$$2\ \mathrm{NOBr}_{(g)} \rightleftharpoons 2\ \mathrm{NO}_{(g)} + \mathrm{Br}_{2(g)} \qquad \Delta H = +16.1\ \mathrm{kJ/mol\ Br_2}$$

5. **What will happen to the concentration of NO if the temperature is increased?**

   a) increase
   b) decrease
   c) remain unchanged

6. What will happen to the concentration of if the temperature is increased?

   a) increase
   b) decrease
   c) remain unchanged

7. What will happen to the concentration of if the temperature is increased?

   a) increase
   b) decrease
   c) remain unchanged

8. What will happen to the value of if the temperature is increased?

   a) increase
   b) decrease
   c) remain unchanged

SOLUTIONS

CHAPTER 6

Standard 6: Students will understand the properties that describe solutions in terms of concentration, solutes, solvents, and the behavior of acids and bases.

Standard 6, Objective 1: Describe factors affecting the process of dissolving and evaluate the effects that changes in concentration have on solutions.

Vocabulary

- Solution
- Solute
- Solvent
- Molarity
- Molality
- Concentrated
- Dilute
- Dissolve
- parts per millions
- concentrated
- dilute
- dissolve

Lesson Objectives

- understand properties of solutions
- define a solution.
- describe the composition of solutions.
- define the terms solute and solvent.
- identify the solute and solvent in a solution.
- describe the different types of solutions and give examples of each type.

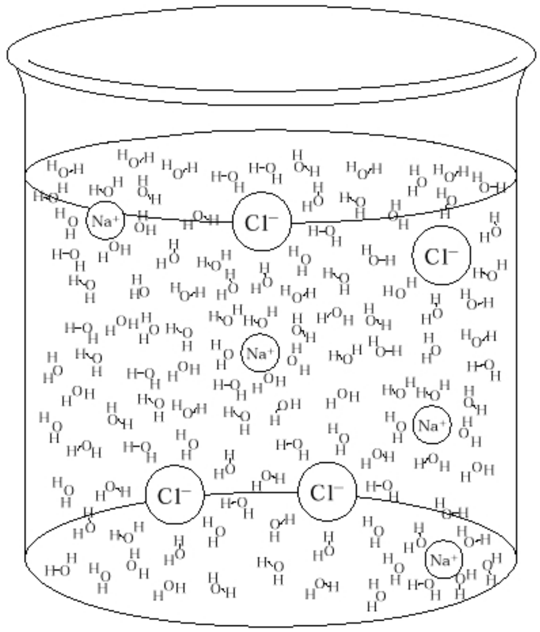
Introduction

Dante was shopping for juice and saw a bottle of juice concentrate. He read the label and learned that some of the water had been removed from the juice before it was bottled, so the concentrate had to be mixed with water before drinking. Next to the bottle of juice was a can of juice. The canned juice was ready to drink. Juice concentrate is an example of a concentrated solution. Ready-to-drink juice is an example of a dilute solution.

Mixtures

A solution is a homogeneous mixture of substances (the prefix "homo-" means "same"), meaning that the properties are the same throughout the solution. Take, for example, the vinegar that is used in cooking. Vinegar is approximately 5% acetic acid in water. This means that every teaspoon of vinegar contains 5% acetic acid and 95% water.

When a solution is said to have uniform properties, the definition is referring to properties at the particle level. What does this mean? Consider brass as an example, which is an alloy made from copper and zinc. An alloy is a homogeneous mixture formed when one solid is dissolved in another. To the naked eye, a brass coin seems like it is just one substance, but at a particle level two substances are present (copper and zinc). Now consider a handful of zinc filings and copper pieces. Is this now a homogeneous solution? The properties of any scoop of the mixture you are holding would not be consistent with any other scoop you removed from the mixture, so the combination of zinc filings and copper pieces in a pile does not represent a homogeneous mixture.



The solvent and solute are the two basic parts of a solution. The solvent is the substance present in the greatest amount. The solute, then, is the substance present in the least amount. When you are making a cup of hot chocolate, you take a teaspoon of cocoa powder and dissolve it in a cup of hot water. Since much less cocoa powder is used than water, the cocoa powder is the solute and the water is the solvent.

Just because solutions have the same composition throughout does not mean the composition cannot be varied. If you were to dissolve a quarter teaspoon of table salt in one cup of water, a solution would form. You can, however, vary the composition of this solution to a point. If you were to add another half teaspoon of salt to the cup of water, you would still make a solution, but the composition of this solution would be different from the last one. What would happen if you tried to dissolve half cup of salt in the water? At this point, the solution has passed its limit as to the amount of salt that can be dissolved in it, so it would no longer be a homogeneous solution. As a result, solutions have a constant composition that can be varied up to a point. There are, however, limits to the amount of substance that can be dissolved into another substance and still remain homogeneous.

LESSON SUMMARY

- In a solution, a solute is present in the least amount (less than 50% of the solution) whereas the solvent is present in the greater amount (more than 50% of the solution).
- A solution is a mixture that has the same properties throughout.

REVIEW QUESTIONS

1. Which of the following is a solution?
   a) milk
   b) blood
   c) gold
   d) air
   e) sugar

2. Which of the following is not a true solution?
   a) vinegar
   b) sand and water
   c) hard water, $CaCO_{3(aq)}$
   d) mercury alloy

MEASURING CONCENTRATION

Lesson Objectives

- define the terms concentrated and dilute.
- define concentration and list the common units used to express the concentration of solutions.
- define molarity, ppm, and molality.
- write the formula for molarity and use the formula to perform calculations involving the molarity, moles of solute, and volume of a solution.
- calculate mass percent, ppm, molality, and mole fraction.

Introduction

Concentration is the measure of how much a given substance is mixed with another substance. Solutions can be said to be dilute or concentrated. A concentrated solution is one in which there is a large amount of solute in a given amount of solvent.
A dilute solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. In order to make juice, you mix the frozen juice from inside these containers with about 3 or 4 times the amount of water. Therefore, you are diluting the concentrated juice. The terms "concentrated" and "dilute," however, only provide a qualitative way of describing concentration. In this lesson, we will explore some quantitative methods of expressing solution concentration.

http://en.wikipedia.org/wiki/File:Orange_juice_1_edit1.jpg

MOLARITY

Of all the quantitative measures of concentration, **molarity** is the one used most frequently by chemists. Molarity is defined as the number of moles of solute per liter of solution. The symbol given for molarity is M, or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[\mathrm{Ag}^+]$ refers to the molarity of the silver ion.

molarity = mol/L $\frac{\text{moles of solute}}{\text{liters of solution}}$ =

Example:

What is the concentration, in mol/L, when 2.34 mol of NaCl has been dissolved in $500. \text{ mL}$ of $\mathrm{H_2O}$?

Solution:

$$[\mathrm{NaCl}] = \frac{2.34 \text{ mol}}{0.500 \text{ liter}} = 4.68 \text{ M}$$

The concentration of the NaCl solution is 4.68 mol/L.

Example:

What would be the mass of potassium sulfate in $500. \text{ mL}$ of a 1.25 mol/L potassium sulfate solution?

Solution:

$\mathrm{M} = \frac{\text{mol}}{\text{L}}, \quad \text{mol} = \mathrm{M} \times \mathrm{L}$ so

$$(1.25 \text{ mol/L})(0.500 \text{ L}) = 0.625 \text{ mol}$$

$$\text{mol} = \frac{\text{mass}}{\text{molar mass}}$$

so $$\text{mass} = (\text{mol}) \cdot (\text{molar mass})$$

$$\text{mass} = (0.625 \text{ mol})(174.3 \text{ g/mol}) = 109 \text{ g}$$

Therefore, the mass of the K_2SO_4 that dissolves in 500. mL of H_2O to make this solution is 109 g.

MOLALITY

Molality is another way to measure concentration of a solution. It is calculated by dividing the number of moles of solute by the number of kilograms of solvent. Molality has the symbol m.

$$\text{molality (m)} = \frac{\text{mol of solute}}{\text{kg of solvent}}$$

Molarity, if you recall, is the number of moles of solute per volume of solution. Volume is temperature dependent. As the temperature rises, the molarity of the solution will actually decrease slightly because the volume will increase slightly. Molality does not involve volume, and mass is not temperature dependent. Thus, there is a slight advantage to using molality over molarity when temperatures move away from standard conditions.

Example:

Calculate the molality of a solution of hydrochloric acid where 12.5 g of hydrochloric acid has been dissolved in 115 g of water.

Solution:

$$\text{mol HCl} = \frac{12.5 \text{ g}}{36.46 \text{g/mol}} = 0.343 \text{ mol}$$

Here is a video of a teacher writing on an electronic blackboard. It shows how to calculate molarity, molality, and mole fraction:
http://www.youtube.com/watch?v=9br3XBjFszs (7:36).

PARTS PER MILLION

Parts per million is another unit for concentration. Parts per million denotes that there is 1 milligram of solute for every kilogram of solvent. It is used most frequently when dealing with environmental issues. You may have heard about parts per million when scientists are referring to drinking water or poisons in fish and other food products. To calculate parts per million, the following formula is used.

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

Example:

Mercury levels in fish have often been at the forefront of the news for people who love to eat fresh fish. Salmon, for instance, contains 0.01 ppm compared to shark which contains 0.99 ppm. In the United States, canned tuna is the most popular selling fish and has a mercury level of 0.12 ppm, according to the FDA statistics. If one were to consume 1.00 kg of canned tuna over a certain time period, how much mercury would be consumed?

Solution:

$$\text{ppm} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 10^6$$

$$\text{mass of solute} = \frac{(\text{mass of solution})(\text{ppm})}{1 \times 10^6}$$

$$\text{mass of solute} = \frac{(1000.\ \text{g})(0.12)}{1 \times 10^6} = 1.2 \times 10^{-4}\ \text{g}$$

LESSON SUMMARY

- Concentration is the measure of how much of a given substance is mixed with another substance.
- Molarity is the number of moles of solute per liter of solution.
- Molality is calculated by dividing the number of moles of solute by the kilograms of solvent. It is less common than molarity but more accurate because of its lack of dependence on temperature.
- Parts per million means that there is 1 milligram of solute for every kilogram of solvent. Therefore, it is the mass of solute per mass of solution multiplied by 1 million.

Review Questions

1. Most times when news reports indicate the amount of lead or mercury found in foods, they use the concentration measures of ppb (parts per billion) or ppm (parts per million). Why use these over the others we have learned?
2. What is the molarity of a solution prepared by dissolving 2.5 g of $\mathrm{LiNO_3}$ in sufficient water to make 60. mL of solution?
   a. 0.036 mol/L
   b. 0.041 mol/L
   c. 0.60 mol/L
   d. 0.060 mol/L
3. A solution is known to have a concentration of 325 ppm. What is the mass of the solute dissolved in 1.50 kg of solvent?
   a. 0.32 mg
   b. 0.49 mg
   c. 325 mg
   d. 488 mg

4. Calculate the molality of a solution of copper(II) sulfate where 11.25 g of the crystals has been dissolved in 325 g of water.

    a. 0.0346 m
    b. 0.0705 m
    c. 0.216 m
    d. None of the above.

5. What is the concentration of each of the following solutions in mol/L?

    a. 3.50 g of potassium chromate dissolved in 100 mL of water
    b. 50.0 g of magnesium nitrate dissolved in 250 mL of water.
    c. Find the mass of aluminum nitrate required to produce 750 g of a 1.5 molal solution.

6. The Dead Sea contains approximately 332 grams of salt per kilogram of seawater. Assume this salt is all NaCl. Given that the density of the Dead Sea water is approximately 1.20 g/mL, calculate the molarity of NaCl.

FACTORS THAT AFFECT THE RATE OF DISSOLUTION

If you're like Tanya in this picture, you prefer your iced tea sweetened with sugar. Sweetened iced tea is a solution in which solid sugar (the solute) is dissolved in cold liquid tea, which is mostly water (the solvent). When you add sugar to tea, particles of water pull apart particles of sugar. The particles of sugar spread throughout the tea, making all of it taste sweet.

Factors that Affect the Rate of Dissolving

Did you ever get impatient and start drinking a sweetened drink before all the sugar has dissolved? As you drink the last few drops, you notice that some of the sugar is sitting on the bottom of the container.

Q: What could you do to dissolve the sugar faster?

A: The rate of dissolving is influenced by several factors, including stirring, temperature of solvent, and size of solute particles. You can see videos demonstrating these factors at the following URLs:
http://www.youtube.com/watch?v=cF55VAk1Nlk
http://www.youtube.com/watch?v=Yb-TdSqmIvc
http://www.youtube.com/watch?v=TO42FOay7rg

Stirring

Stirring a solute into a solvent speeds up the rate of dissolving because it helps distribute the solute particles throughout the solvent. For example, when you add sugar to iced tea and then stir the tea, the sugar will dissolve faster. If you don't stir the iced tea, the sugar may eventually dissolve, but it will take much longer.

Temperature

The temperature of the solvent is another factor that affects how fast a solute dissolves. Generally, a solute dissolves faster in a warmer solvent than it does in a cooler solvent because particles have more energy of movement. For example, if you add the same amount of sugar to a cup of hot tea and a cup of iced tea, the sugar will dissolve faster in the hot tea.

Particle Size

A third factor that affects the rate of dissolving is the size of solute particles. For a given amount of solute, smaller particles have greater surface area. With greater surface area, there can be more contact between particles of solute and solvent. For example, if you put granulated sugar in a glass of iced tea, it will dissolve more quickly than the same amount of sugar in a cube (see Figure below). That's because all those tiny particles of granulated sugar have greater total surface area than a single sugar cube.

SUMMARY

The rate of dissolving of a solute in a solvent is faster when the solute and solvent are stirred, the solvent is warmer, or the solute consists of smaller particles with more surface area.

PRACTICE

Design an experiment to investigate the effects of temperature on the rate of dissolving. You can see an example in the video at the following URL. Identify dependent and independent variables and factors that need to be controlled. Make a list of materials needed for your experiment, and write a step-by-step procedure.
http://www.youtube.com/watch?v=T042FOay7rg

REVIEW

1. List three factors that affect the rate at which a solute dissolves in a solvent.
2. Gina is trying to dissolve bath salts in her bathwater. How could she speed up the rate of dissolving?

STANDARD 6, OBJECTIVE 2: SUMMARIZE THE QUANTITATIVE AND QUALITATIVE EFFECTS OF COLLIGATIVE PROPERTIES.

Vocabulary

- boiling point elevation
- colligative property
- freezing point depression

Colligative Properties

Lesson Objectives

- explain what the term colligative means and list colligative properties.
- indicate what happens to the boiling point, the freezing point, and the vapor pressure of a solvent when a solute is added to it.
- explain why a solution has a lower vapor pressure than the pure solvent of that solution.
- explain why a solution has an elevated boiling point and a depressed freezing point.
- calculate boiling point elevations and freezing point depressions for both non-electrolyte and electrolyte solutions.

Introduction

People who live in colder climates have seen the trucks put salt on the roads when snow or ice is forecast. Why do they do that? When planes fly in cold weather, the planes need to be de-iced before liftoff. Why is that done? It turns out that pure solvents differ from solutions in their boiling points and freezing points. In this lesson, you will understand why these events occur. You will also learn to calculate exactly how much of an effect a specific solute can have on the boiling point or freezing point of a solution. Colligative properties are properties that are due only to the number of particles in solution and not to the chemical properties of the solute.

Boiling Point Elevation

The boiling point of a liquid occurs when the vapor pressure above the surface of the liquid equals the surrounding pressure. At 1 atm of pressure, pure water boils at 100^oC, but salt water does not. When table salt is added to water, the resulting solution has a higher boiling point than water alone. Essentially, the solute particles take up space at the solvent/air interface, physically blocking some of the more energetic water molecules from escaping into the gas phase. This is true for any solute added to a solvent. Boiling point elevation is another example of a colligative property, meaning that the change in boiling point is related only to the number of solute particles in solution, regardless of what those particles are. A 0.20 m solution of table salt would have the same change in boiling point as a 0.20 m solution of KNO_3.

Freezing Point Depression

The effect of adding a solute to a solvent has the opposite effect on the freezing point of a solution as it does on the boiling point. Recall that the freezing point is the temperature at which the liquid changes to a solid. At a given temperature, if a substance is added to a solvent (such as water), the solute-solvent interactions prevent the solvent from going into the solid phase, requiring the temperature to decrease further before the solution will solidify. A common example is found when salt is used on icy roadways. Here the salt is put on the roads so that the water on the roads will not freeze at the normal $0^{o}C$ but at a lower temperature, as low as $-9^{o}C$. The de-icing of planes is another common example of freezing point depression in action. A number of solutions are used, but commonly a solution such as ethylene glycol or a less toxic monopropylene glycol is used to de-ice an aircraft. The aircrafts are sprayed with the solution when the temperature is predicted to drop below the freezing point. The freezing point depression, then, is the difference between the freezing points of the solution and the pure solvent.

http://teachgreenchemnh.wikispaces.com/Greening+freezing+point+depression+lab

Remember that colligative properties are due to the number of solute particles in the solution. Adding 10 molecules of sugar to a solvent will produce 10 solute particles in the solution. However, when the solute is an electrolyte, such as NaCl, adding 10 molecules of solute to the solution will produce 20 ions (solute particles) in the solution. Therefore, adding enough NaCl solute to a solvent to produce a 0.20 m solution will have twice the effect of adding enough sugar to a solvent to produce a 0.20 m solution.

The van't Hoff factor is the number of particles that the solute will dissociate into upon mixing with the solvent. For example, sodium chloride (NaCl) will dissociate into two ions, so the van't Hoff factor for NaCl is i = 2. For lithium nitrate ($LiNO_3$), i = 2, and for calcium chloride ($CaCl_2$), i = 3. More ions will have a greater effect on colligative properties. In other words, the boiling point will increase more when there are more ions and the freezing point will decrease more when there are more ions.

LESSON SUMMARY

- Colligative properties are properties that are due only to the number of particles in solution and not to the chemical properties of the solute.
- Vapor pressure lowering, boiling point elevation, and freezing point depression are colligative properties.
- For electrolyte solutions, the Van't Hoff factor is added to account for the number of ions that the solute will dissociate into in solution.

FURTHER READING / SUPPLEMENTAL LINKS

The following link hosts several videos on how to solve various types of chemistry solutions problem.
http://www.kentchemistry.com/moviesfiles/chemguy/AP/ChemguySolutions.htm

REVIEW QUESTIONS

1. What must be measured in order to determine the freezing point depression?
2. From a list of solutions with similar molalities, how could you quickly determine which would have the highest boiling point?
3. Why would table salt not be a good solution to use when deicing a plane?
4. Determine which of the following solutions would have the lowest freezing point.
   a. 15 g of ammonium nitrate in 100. g of water.
   b. 50 g of glucose in 100. g of water.
   c. 35 g of calcium chloride in 150. g of water.
5. What is the van't Hoff factor for each of the following:
   a. $MgCl_2$
   b. Ammonium sulfate
   c. CH_3OH
   d. Potassium chloride
   e. KCH_3COO

Vocabulary

- indicators
- acid
- base

STANDARD 6, OBJECTIVE 3: DIFFERENTIATE BETWEEN ACIDS AND BASES IN TERMS OF HYDROGEN ION CONCENTRATION.

ACIDS AND BASES

Lesson Objectives

- list the properties of acids
- list the properties of bases

Introduction

We interact with acids on a daily basis without even realizing it. For example, the chemical names for aspirin and vitamin C are acetylsalicylic acid and ascorbic acid; both will produce H^+ ions when dissolved in water. Acetic acid ($HC_2H_3O_2$) is the primary component in vinegar, and formic acid (HCO_2H) is what causes ant bites to sting.

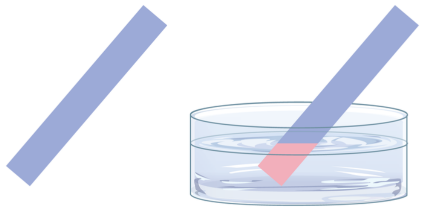
Hydrochloric acid (HCl) is stomach acid, phosphoric acid (H_3PO_4) is commonly found in dark soft drinks, and sulfuric acid (H_2SO_4) is used in car batteries. As you work your way through this chapter, try to notice how the properties of acids and bases manifest themselves in everyday situations.

Properties of Acids

One property that is common to all acids is a sour taste. You are probably most familiar with this in relation to citric acid, which is what makes lemons and other citrus fruits taste sour. In fact, sour taste buds are essentially just complicated H^+ sensors. The fact that one of our primary tastes is concerned solely with determining the acidity of what goes in our mouths further underscores the importance of acids in our lives.

However, testing whether something is acidic by taste is generally not a good idea. Another way to test for acidity is to use an indicator. Indicators are substances that can be used to determine the relative acidity or basicity of a solution, generally through a very distinct color change.

One common type of indicator is litmus paper. If a piece of blue litmus paper turns red when dipped into a solution, it means that the solution is acidic.



Another property common to many acids is that they can react with certain metals to form hydrogen gas. Examples of this type of reaction are shown below. Note that these are all single replacement reactions where a pure element reacts with a compound

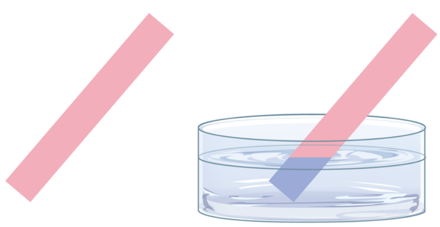
$$Zn_{(s)} + 2\ HCl_{(aq)} \rightarrow ZnCl_{2(aq)} + H_{2(g)}$$

$$Mg_{(s)} + 2\ HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

Properties of Bases

Bases also have a number of characteristic properties. Most bases are slippery and quite bitter (though not all bitter compounds are basic). Caffeine and milk of magnesia (chemical formula $Mg(OH)_2$) are two bases that you may have had the opportunity to taste, although the bitterness is generally masked by other flavors when these compounds are consumed. Other common bases are found in a number of cleaning products, including Drano ($NaOH$) and Windex (NH_4OH).

Like acids, bases can be identified by the use of an indicator. For example, if red litmus paper is dipped into a basic solution, it will turn blue.



This video discusses the properties of acids and bases:
http://www.youtube.com/watch?v=mm7Hcff5b6g (8:02).
A discussion of the difference between strong and weak acids is available at:
http://www.youtube.com/watch?v=XTdkWGlmtSc (9:47).

LESSON SUMMARY

- Acids turn blue litmus paper red, taste sour, and react with metals to produce hydrogen gases.
- Bases turn red litmus paper blue, have a bitter taste, and are slippery to the touch.

FURTHER READING / SUPPLEMENTAL LINKS

- Visit this website to learn more about examples and properties of acids and bases. http://qldscienceteachers.tripod.com/junior/chem/acid.html

REVIEW QUESTIONS

1. What are the properties of acids? Give a common example.
2. Which statement best describes a characteristic of acid solutions?
3. They react with some metals to form hydrogen gas.
4. They turn red litmus paper blue.
5. They taste bitter.
6. They are made from nonmetal oxides.

THE pH CONCEPT

Vocabulary

- hydrogen ion
- pH
- pOH

Lesson Objectives

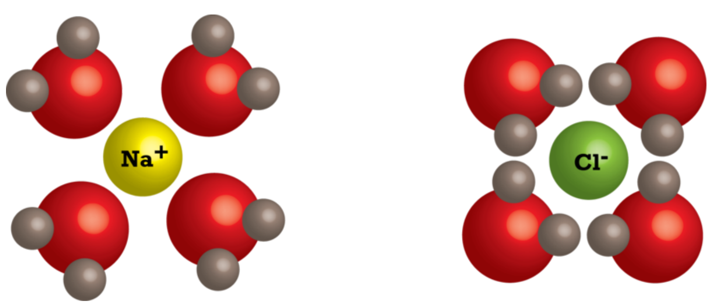
- calculate $[H^+]$ and $[OH^-]$ for a solution of acid or base.
- state the $[H^+]$, $[OH^-]$, and K_w values for the autoionization of water.
- define pH and describe the pH scale.
- write the formulas for pH and pOH and express their values in a neutral solution at 25°C.
- explain the relationships among pH, pOH, and K_w.
- calculate $[H^+]$, $[OH^-]$, pH, and pOH given the value of any one of the other values.
- explain the relationship between the acidity or basicity of a solution and the hydronium ion concentration, $[H_3O^+]$,, and the hydroxide ion concentration, $[OH^-]$, of the solution.

Introduction

We have learned many properties of water, such as pure water does not conduct electricity. The reason pure water does not conduct electricity is because the concentration of ions present when water ionizes is small. In this lesson, we will look a little closer at this property of water and how it relates to acids and bases.

The Hydronium Ion

Recall that ions in solution are hydrated. That is, water molecules are loosely bound to the ions by the attraction between the charge on the ion and the oppositely charged end of the polar water molecules, as illustrated in the figure below. When we write the formula for these ions in solution, we do not show the attached water molecules. It is simply



recognized by chemists that ions in solution are always hydrated.

As with any other ion, a hydrogen ion dissolved in water will be closely associated with one or more water molecules. This fact is sometimes indicated explicitly by writing the hydronium ion, H_3O^+, in place of the hydrogen ion, H^+. Many chemists still use $H^+_{(aq)}$ to represent this situation, but it is understood that this is just an abbreviation for what is really occurring in solution. You are likely to come across both, and it is important for you to understand that they are actually describing the same entity. When using the hydronium ion in a chemical equation, you may need to add a molecule of water to the other side so that the equation will be balanced. This is illustrated in the equations below. Note that you are not really adding anything to the reaction. The (aq) symbol indicates that the various reaction components are dissolved in water, so writing one of these water molecules out explicitly in the equation does not change the reaction conditions.

$$\mathrm{HCl}_{(aq)} \rightarrow \mathrm{H}^+_{(aq)} + \mathrm{Cl}^-_{(aq)} \text{ (not showing hydronium)}$$

$$\mathrm{HCl}_{(aq)} + \mathrm{H_2O}_{(l)} \rightarrow \mathrm{H_3O}^+_{(aq)} + \mathrm{Cl}^-_{(aq)} \quad \text{(showing hydronium)}$$

Relationship Between [H+] and [OH-]

Even totally pure water will contain a small amount of H^+ and OH^-. This is because water undergoes a process known as auto ionization. Auto ionization occurs when the same reactant acts as both the acid and the base. Look at the reaction below.

The ionization of water is frequently written as:

$$\mathrm{H_2O}_{(l)} \rightarrow \mathrm{H}^+ + \mathrm{OH}^-$$

A further definition of acids and bases can now be made:

When $[\mathrm{H_3O}^+] = [\mathrm{OH}^-]$(as in pure water), the solution is *neutral*.
When $[\mathrm{H_3O}^+] > [\mathrm{OH}^-]$, the solution is an *acid*.
When $[\mathrm{H_3O}^+] < [\mathrm{OH}^-]$, the solution is a *base*.

pH AND POH

There are a few very concentrated acid and base solutions used in industrial chemistry and laboratory situations. For the most part, however, acid and base solutions that occur in nature, used in cleaning, and used in biochemistry applications are relatively dilute. Most of the acids and bases dealt with in laboratory situations have hydrogen ion concentrations between 1.0 M and 1.0 M x 10^{-14} M. Expressing hydrogen ion concentrations in exponential numbers can become tedious, so a Danish chemist named Søren Sørensen developed a shorter method for expressing acid strength or hydrogen ion concentration with a non-exponential number. This value is referred to as **pH** and is defined by the following equation:

$$pH = -\log[H^+]$$

where $p = \log$ and H refers to the hydrogen ion concentration. The p from pH comes from the German word *potenz*, meaning power or the exponent of. Rearranging this equation to solve for $[H^+]$, we get $[H^+] = 10^{-pH}$. If the hydrogen ion concentration is between 1.0 M and 1.0 M x 10^{-14} M, the value of the pH will be between 0 and 14.

Example:

Calculate the pH of a solution where $[H^+] = 0.01 \text{ mol/L}$.

Solution:

$$pH = -\log(0.01)$$

$$pH = -\log(1 \times 10^{-2})$$

$$pH = 2$$

Example:

Calculate the $[H^+]$ if the pH is 4.

Solution:

$$[H^+] = 10^{-pH}$$

$$[H^+] = 10^{-4}$$

$$[\mathrm{H}^+] = 1 \times 10^{-4}\ \mathrm{mol/L}$$

Example:

Calculate the pH of saliva, where $[\mathrm{H}^+] = 1.58 \times 10^{-6}\ \mathrm{mol/L}$.

Solution:

$$\mathrm{pH} = -\log[\mathrm{H}^+] = -\log(1.58 \times 10^{-6})$$

$$\mathrm{pH} = 5.8$$

Example:

Fill in the rest of **Table** below.

Hydrogen ion concentration and corresponding pH.

$[H^+]$ **in mol/L**	$-\log[H^+]$	pH
0.1	1.00	1.00
0.2	0.70	0.70
1×10^{-5}	?	?
?	?	6.00
0.065	?	?
?	?	9.00

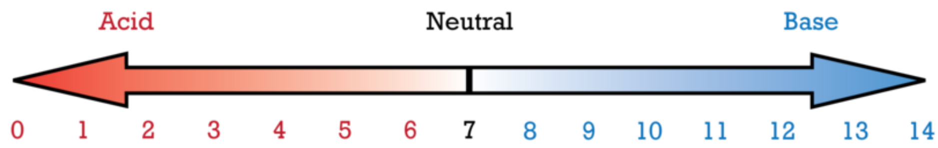
Solution:

The completed table is shown below (**Table** below).

Hydrogen ion concentration and corresponding pH.

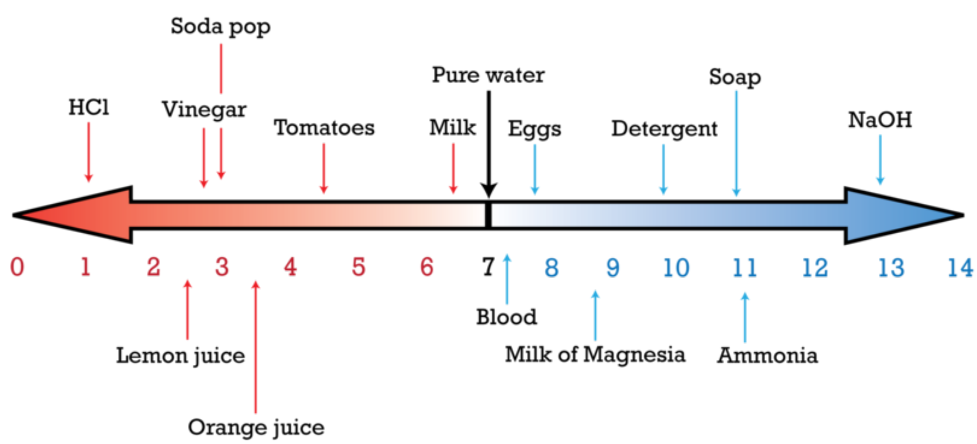
$[H^+]$ **in mol/L**	$-\log[H^+]$	pH
0.1	1.00	1.00
0.2	0.70	0.70
1.00×10^{-5}	5	5
1.00×10^{-6}	6.00	6.00
0.065	1.19	1.19
1.00×10^{-9}	9.00	9.00

An acid with pH = 1, then, is stronger than an acid with pH = 2 by a factor of 10. Simply put, lower pH values correspond to higher H^+ concentrations and more acidic solutions, while higher pH values correspond to higher OH^- concentrations and more basic solutions. This is illustrated in the figure below. It should be pointed out that there are acids and bases that fall outside the pH range depicted. However, we will confine ourselves for now to those falling within the 0-14 range, which covers $[H^+]$ values from 1.0 M all the way down to $1 \times 10^{-14}\text{ M}$.



pH versus Acidity	
pH level	**Solution**
$pH < 7$	Acid
$pH = 7$	Neutral
$pH > 7$	Basic

Have you ever cut an onion and had your eyes water up? This is because of a compound with the formula C_3H_6OS that is found in onions. When you cut the onion, a variety of reactions occur that release a gas. This gas can diffuse into the air and eventfully mix with the water found in your eyes to produce a dilute solution of sulfuric acid. This is what irritates your eyes and causes them to water. There are many common examples of acids and bases in our everyday lives. Look at the pH scale below to see how these common examples relate in terms of their pH.



For a classroom demonstration of pH calculations (5d, 5f; 1e I&E Stand.), see: http://www.youtube.com/watch?v=lca_puB1R8k (9:45).

LESSON SUMMARY

- Water ionizes to a very slight degree according to the equation $H_2O_{(l)} \leftrightarrows [H^+] + [OH^-]$.
- In pure water at 25°C, $[H^+] = [OH^-] = 1.00 \times 10^{-7}\ M$.
- The equilibrium constant for the dissociation of water, K_w, is equal to 1.00×10^{-14} at $25°C$.
- $pH = -\log[H^+]$
- $pOH = -\log[OH^-]$
- $pK_w = -\log K_w$
- $pH + pOH = pK_w = 14.0$

FURTHER READING / SUPPLEMENTAL LINKS

The websites below have more information about pH.

- http://www.johnkyrk.com/pH.html
- http://purchon.com/chemistry/ph.htm

REVIEW QUESTIONS

1. What is the $[H^+]$ ion concentration in a solution of $0.350 \text{ mol/L } H_2SO_4$?

   a. 0.175 mol/L

   b. 0.350 mol/L

   c. 0.700 mol/L

   d. $1.42 \times 10^{-14} \text{ mol/L}$

2. A solution has a pH of 6.54. What is the concentration of hydronium ions in the solution?

   a. $2.88 \times 10^{-7} \text{ mol/L}$

   b. $3.46 \times 10^{-8} \text{ mol/L}$

   c. 6.54 mol/L

   d. 7.46 mol/L

3. A solution has a pH of 3.34. What is the concentration of hydroxide ions in the solution?

   a. $4.57 \times 10^{-4} \text{ mol/L}$

   b. $2.19 \times 10^{-11} \text{ mol/L}$

   c. 3.34 mol/L

   d. 10.66 mol/L

4. Fill in **Table** below and rank the solutions in terms of increasing acidity.

Solutions	$[H^+]$ (mol/L)	$-\log\ [H^+]$	pH
A	0.25	0.60	0.60
B	?	2.90	?
C	1.25×10^{-8}	?	?
D	0.45×10^{-3}	?	?
E	?	1.26	?

5. Wine is considered to be an acidic solution. Determine the concentration of hydronium ions in wine with $\mathrm{pH}\ 3.81$.

NEUTRALIZATION

Vocabulary

- neutralization

Lesson Objectives

- explain what is meant by a neutralization reaction and give an example of one.
- write a balanced equation for the reaction that occurs when an acid reacts with a base.
- predict the salt that will be produced from the neutralization reaction between a given acid and base.

Introduction

Neutralization is a reaction between an acid and a base that produces water and a salt. The general reaction is shown below:

$$\text{acid} + \text{base} \rightarrow \text{salt} + \text{water}$$

After reviewing the concept of ionic compounds, we will examine neutralization reactions and look at the different types of salts that can be formed from acids and bases as they react with one another.

Neutralization Reactions

Acids are a combination of a hydrogen cation and a nonmetal anion. Examples include HCl, HNO_3, and $HC_2H_3O_2$. Many bases are a combination of metal cations and nonmetal anions. Examples include $NaOH$, KOH, and $Mg(OH)_2$. According to the Arrhenius definitions of acids and bases, the acid will contribute an H^+ ion that will react to neutralize the OH^- ion contributed by the base, producing neutral water molecules.

All acid-base reactions produce salts. The anion from the acid will combine with the cation from the base to form the ionic salt. Examples are shown below.

$$HClO_{4(aq)} + NaOH_{(aq)} \rightarrow NaClO_{4(aq)} + HOH_{(l)}$$

$$H_2SO_{4(aq)} + 2KOH_{(aq)} \rightarrow K_2SO_{4(aq)} + 2HOH_{(l)}$$

(Note: $HOH = H_2O$)

No matter what the acid or the base may be, the products of this type of reaction will always be a salt and water. Aside from the fact that the H^+ ion will neutralize the OH^- ion to form water, we also know that these reactions are double displacement reactions, because they consist of cations exchanging anions.

LESSON SUMMARY

- A neutralization reaction between an acid and a base will produce a salt and water.
- In a neutralization reaction, the acid will produce H^+ ions that react to neutralize the OH^- ions produced by the base, forming neutral water. The other product will be an ionic salt.

TITRATION

Vocabulary

- titration

Lesson Objectives

- explain what an acid/base indicator is.
- explain what a titration is.
- describe how titrations can be used to determine the concentration of an acid or a base in solution.

Introduction

The typical laboratory procedure for determining the concentration of acid and/or base in a solution is to complete a titration. There are three main types of titration experiments. As we go through this lesson, we will take apply the knowledge we have obtained about acids and bases, chemical reactions, and molarity calculations to the concept of titrations.

Indicators

Recall from the chapter on “Acids-Bases” that an indicator is a substance that changes color at a specific pH and is used to indicate the pH of the solution. One example of an indicator is litmus paper. Litmus paper is paper that has been dipped in a substance that will undergo a color change when it is exposed to either an acid or a base. If red litmus paper turns blue, the solution is basic (pH > 7), and if blue litmus turns red the solution is acidic (pH < 7).

There are two requirements for a substance to function as an acid-base indicator: 1) the substance must have an equilibrium affected by hydrogen ion concentration, and 2) the two forms of the compound on opposite sides of the equilibrium must have different colors. Most indicators function in the same general manner and can be presented by a generic indicator equation. In the equation below, we represent in the indicator ion with a hydrogen ion attached as HIn, and we represent the indicator ion without the hydrogen attached as In^-.

Colors and pH Ranges for Common Indicators		
Indicator	**pH Range**	**Color Change**
Methyl Violet	0.0 – 1.6	Yellow - Blue
Thymol Blue	1.2 – 2.8	Red - Yellow
Orange IV	1.3 – 3.0	Red - Yellow
Methyl Orange	3.2 – 4.4	Red - Orange
Bromophenol Blue	3.0 – 4.7	Orange/Yellow - Violet
Congo Red	3.0 – 5.0	Blue - Red
Bromocresol Green	3.8 – 5.4	Yellow - Blue
Methyl Red	4.8 – 6.0	Red - Yellow
Litmus	5.0 – 8.0	Red - Blue
Chlorophenol Red	4.8 – 6.2	Yellow - Red
Bromothymol Blue	6.0 – 7.6	Yellow - Blue
Phenol Red	6.4 – 8.2	Yellow - Red/Violet
Thymol Blue	8.0 – 9.6	Yellow - Blue
Phenolphthalein	8.2 – 10.0	Colorless - Pink
Alizarin Yellow R	10.1 – 12.0	Yellow - Red
Methyl Blue	10.6 – 13.4	Blue - Pale Violet
Indigo Carmine	11.4 – 13.0	Blue - Yellow

There are many more indicators than are shown in Table above, but these are ones that you may find in common chemistry classroom laboratories. One example of an indicator not found in the table is known as the universal indicator. The universal indicator is a solution that has a different color for each pH from 0 – 14. Universal indicator is produced by creatively mixing many of the individual indicators together so that a different color is achieved for each different pH. It is used for many types of experiments to determine if solutions are acids or bases and where on the pH scale the substance belongs. The chart below indicates the colors of universal indicator for different pH values.

Color of Universal Indicator at Various pH Values

Example:

If the pH of the solution is 4.8, what would be the color of the solution if the following indicators were added?

1. Universal indicator
2. Bromocresol Green
3. Phenol red

Solution:

1. Universal indicator = Orange to orange-yellow
2. Bromocresol Green = green (midway pH = 4.6)
3. Phenol red = yellow

Example:

A solution found in the laboratory was tested with a number of indicators. These were the results:

- Phenolphthalein was colorless
- Bromocresol green was blue
- Methyl red was yellow
- Phenol red was yellow

What was the pH of the solution?

Solution:

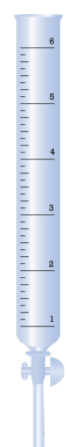
- Phenolphthalein was colorless, $\mathrm{pH} < 8.0$
- Bromocresol green was blue, $\mathrm{pH} > 5.4$
- Methyl red was yellow, $\mathrm{pH} > 6.0$
- Phenol red was yellow, $\mathrm{pH} < 6.4$

Therefore, the pH of the solution must be between 6.0 and 6.4.

The Titration Process

One of the properties of acids and bases is that they neutralize each other. In the laboratory setting, an experimental procedure where an acid is neutralized by a base (or vice versa) is known as titration. Titration is the addition of a known concentration of base (or acid), to a solution of acid (or base) of unknown concentration. Since both volumes of the acid and base are known, the concentration of the unknown solution is then mathematically determined.

When doing a titration, you need to have a few pieces of equipment. A burette like the one shown below is used to accurately dispense the volume of the solution of known concentration. An Erlenmeyer flask is used to hold a known volume of the solution whose concentration is unknown. A few drops of the indicator are added to the flask before you begin the titration. The endpoint is the point where the indicator changes color, which tells us that the acid is neutralized by the base. The equivalence point is the point where the number of moles of acid exactly equals the number of moles of base.



The Mathematics of Titration

For the calculations involved here, we will only use our acid and base examples where the stoichiometric ratio of H^+ and OH^- is 1:1. To determine the volume required to neutralize an acid or a base, or in other words, to reach the equivalence point, we will use a formula similar to the dilution formula:

$$M_a \times V_a = M_b \times V_b$$

where M_a is the molarity of the acid, V_a is the volume of the acid, M_b is the molarity of the base, and V_b is the volume of the base. Note that if the acid and base do not neutralize each other in a 1:1 ratio, this equation does not hold true.

Example:

When $10.0\ \mathrm{mL}$ of a $0.125\ \mathrm{mol/L}$ solution of hydrochloric acid, HCl, is titrated with a $0.100\ \mathrm{mol/L}$ solution of potassium hydroxide, KOH, what is the volume of the hydroxide solution required to neutralize the acid? What type of titration is this?
Solution:

$$M_a \times V_a = M_b \times V_b$$

$$M_a = 0.125\ \mathrm{mol/L}$$

$$V_a = 10.0\ \mathrm{mL}$$

$$M_b = 0.100\ \mathrm{mol/L}$$

$$V_b = ?$$

$$M_a \times V_a = M_b \times V_b$$

Therefore, for this strong acid-strong base titration, the volume of base required is 12.5 mL.

This video shows the technique for performing a titration using an indicator: http://www.youtube.com/watch?v=9DkB82xLvNE (5:03).

LESSON SUMMARY

- An indicator is a substance that changes color at a specific pH and is used to indicate the pH of the solution relative to that point.
- Universal indicator is a mixture of indicators that produces a different color for each pH from 0 – 14.
- A titration is the addition of a known concentration of base (or acid) to a solution of acid (or base) of unknown concentration.
- The endpoint is the point in the titration where the indicator changes color.
- A standard solution is a solution whose concentration is known exactly and is used to find the exact concentration of the titrant.
- For titrations where the stoichiometric ratio of mol H^+ to mol OH^- is 1:1, the concentrations or volumes for the unknown acid or base can be calculated with the formula $Ma \times Va = Mb \times Vb$.

FURTHER READING / SUPPLEMENTAL LINKS

- The following link is to a video about acid-base neutralization and titration.

  http://link.brightcove.com/services/player/bcpid9113583001?bctid=1405713919

- The video at the link below shows the lab techniques needed for titration.

  http://chem-ilp.net/labTechniques/TitrationVideo.htm

- This video is a ChemStudy film called "Acid Base Indicators." The film is somewhat dated but the information is accurate.

  http://www.youtube.com/watch?v=yi8QrjmV6Sw

REVIEW QUESTIONS

1. Why do you think there would be more experimental error when using an indicator instead of a pH meter during a titration?

2. If 22.50 mL of a sodium hydroxide solution is necessary to neutralize 18.50 mL of 0.1430 mol/L HNO_3 solution, what is the concentration of $NaOH$?
   a. 0.1176 mol/L
   b. 0.1430 mol/L
   c. 0.1740 mol/L
   d. 2.64 mol/L

3. Calculate the concentration of hypochlorous acid if 25.00 mL of $HClO$ is neutralized by 32.34 mL of a 0.1320 mol/L solution of sodium hydroxide.

GLOSSARY

acid: ionic compound that produces positive hydrogen ions (H^+) when dissolved in water

activation energy: energy needed to start a chemical reaction. With increasing temperature, the kinetic energy of the particles and the number of particles with energy greater than the activation energy increases.

alkali earth metals: group 2a of the periodic table

alkali metals: group 1a of the periodic table

alpha particle: an alpha particle is a helium nucleus, having a +2 charge and a mass of 4, having two protons and two neutrons

anion: a nonmetal that has gained an electron

atom: smallest part of an element that can't be broken down by ordinary means

atomic mass unit (amu): a unit of mass equal to one-twelfth the mass of a carbon-twelve atom

atomic number: the number of protons in the nucleus of an atom

avogadro's number: the number of objects in a mole; equal to 6.02 x 1023

balanced chemical equation: a chemical equation in which the number of each type of atom is equal on the two sides of the equation

base: ionic compound that produces negative hydroxide ions (OH^-) when dissolved in water.

beta particle: a beta particle is a high speed electron from the breakdown of a neutron

big bang theory: the idea that the universe was originally extremely hot and dense at some finite time in the past and has since cooled by expanding to the present state and continues to expand today

boiling point: the point at which the vapor pressure of a substance is equal to the environmental pressure surrounding the substance. At this point, particles can overcome intermolecular forces and escape as gas into the surroundings.

catalyst: substance that increases the rate of a chemical reaction but is not changed or used up in the reaction

cation: a metal that has lost an electron

chemical bond: an attraction between atoms by sharing or transferring electrons chemical change: alteration that occurs when one substance is turned into another substance

chemical formula: indicates the atoms and number of atoms in a chemical compound

chemical property: a chemical property is any of a material's properties that becomes evident during a chemical reaction; that is, any quality that can be established only by changing a substance's chemical identity *(source: Wikipedia)*

chemical reaction: the process in which one or more substances are changed into one or more new substances

colligative properties: properties that are due only to the number of particles in solution and not to the chemical properties of the solute

collision frequency: the total number of collisions per second is known as the collision frequency, regardless of whether these collisions are successful or not

collision theory: explains why reactions occur between atoms, ions, and/or molecules and allows us to predict what conditions are necessary for a successful reaction to take place

concentrated: a solution that contains a large amount of solute.

concentration: the number of particles of a substance in a given volume.

conductivity: a measure of a material's ability to conduct an electric current *(source: wikipedia)*

conversion factor: is a factor used to convert one unit of measurement into another unit

cosmic background radiation: energy in the form of radiation leftover from the early big bang

covalent bond: a type of bond in which electrons are shared between bonded atoms

covalent compound: a substance that is made up of one more atom bonded together

Dalton’s atomic theory: the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom

dilute: solution with little dissolved solute has a low concentration

dimensional analysis: is a technique that involves the study of the dimensions (units) of physical quantities dipole-dipole: type of intermolecular force between polar molecules

dissolve: to cause to disperse

ductile: can be drawn out into thin wires

dynamic equilibrium: a state of equilibrium where change is still occurring on a molecular level even though the macroscopic properties remain constant; occurs when the rate of the forward reaction is equal to the rate of the reverse reaction

electrochemical cell: an arrangement of electrodes and ionic solutions in which a spontaneous redox reaction is used to produce a flow of electrons in an external circuit

electron: negatively charged particle that orbits the nucleus of an atom at high speed. It also has a mass of 1/1840th atomic mass unit—it doesn't weigh very much!

electronegativity: the attractiveness an atom's nucleus has for electrons

element: same type of atom that behaves in a certain way, such as oxygen

endothermic reaction: means "taking in heat." It takes more energy to break bonds in the reactants than is released when new bonds form in the products.

energy: the ability to do work or cause change

exothermic reaction: means "turning out heat." It takes less energy to break bonds in the reactants than is released when new bonds form in the products.

experiment: is a controlled method of testing a hypothesis under the conditions we want at a time and place of our choosing

freezing point: the temperature at which a substance freezes (or becomes a solid) is known as its freezing point

frequency: the number of waves that pass a fixed point in a given amount of time

fusion: joining of two nuclei under high pressure and temperature

gamma ray: gamma radiation is the highest energy on the spectrum of electromagnetic radiation—it has no mass or charge

group (family): a vertical column in the periodic table

half-life: length of time it takes for half of a given amount of a radioisotope to decay halogens: group 7a of the periodic table

hydrogen bond: is the strongest type of intermolecular force. The attractive interaction of a polar hydrogen atom in a molecule or chemical group and an electronegative atom, such as nitrogen, oxygen or fluorine, from another molecule or chemical group *(source: wikipedia)*

hydronium ion: a positively charged ion consisting of three hydrogen atoms and one oxygen atom (H_3O^+)

hypothesis: is a proposed explanation for a phenomenon or event

indicator: substance used to determine the relative acidity or basicity of a solution, generally through a very distinct color change

intermolecular forces: are forces of attraction or repulsion which act between neighboring particles (atoms, molecules or ions) *(source: wikipedia)*

ion: an atom or group of atoms with an excess positive or negative charge, lost or gained electrons

ionic bond: formed between the attraction between a positive ion and a negative ion

ionic compound: a bond between ions resulting from the transfer of electrons from a metal to a nonmetal

irreversible reactions: reactions that go to completion

isotopes: atoms of the same element that have the same number of protons but different numbers of neutrons

kelvin temperature scale: has its zero at absolute zero (determined to be -273 c) and uses the same degree size as a degree on the Celsius scale

law of conservation of mass: law stating that matter cannot be created or destroyed in chemical reactions.

law of conservation of matter and energy: states that the total amount of mass and energy in the universe is conserved (does not change)

length: is the measurement of anything from end to end

London dispersion forces: weakest type of intermolecular force. It is attraction between adjacent nonpolar molecules malleability: the ability of a solid to bend or be hammered into other shapes without breaking *(source: wikipedia)*

malleable: can be hammered into thin sheets

mass: a measure of the amount of matter in an object

mass number: the total number of protons and neutrons in the nucleus of an atom

Mendeleev: the Russian chemist credited with organizing the periodic table in the form we use today

metallic bond: force of attraction between a positive metal ion and the valence electrons it shares with other ions of the metal

meter: is the SI unit of length

model: is a descriptive, graphic, or three-dimensional representation of a hypothesis or theory used to help enhance understanding

molarity: the number of moles of solute per liter of solution

mole: an Avogadro's number of objects; 1.00 mole of carbon-12 atoms has a mass of 12.0 grams and contains atoms

Moseley: the chemist credited with finding that each element has a unique atomic number

neutralization: a reaction between an acid and a base that produces water and a salt

neutron: particle with no charge in the nucleus of an atom that has a relative mass of 1 atomic mass unit

noble gases: group 8a of the periodic table

nucleus: the center of an atom, containing one or more protons and neutrons

oxidation: a loss of electrons in an atom or an increase in the oxidation state of an atom

oxidizing agent: a substance that gains electrons in a chemical reaction or undergoes an increase in its oxidation state

parts per million (ppm): the mass of solute per mass of solution multiplied by 1 million

percentage composition: composition expressed by percentages of constituents periodic law: states that the properties of the elements recur periodically as their atomic numbers increase

periodic table: a tabular arrangement of the chemical elements according to atomic number

pH: the negative logarithm of the hydrogen ion concentration

physical change: transformation that does not alter the identity of a substance

physical property: a physical property is any property that is measurable whose value describes a physical system's state *(source: wikipedia)*

polar covalent bond: bond formed by the unequal sharing of electrons between two covalently bonded atoms

polarity: a concept in chemistry which describes how equally bonding electrons are shared between atoms *(source: wikipedia)*

polyatomic ion: a group of bonded atoms with a positive or negative charge

products: materials present at the end of a reaction

proton: positively charge particle in the nucleus of an atom that has a relative mass of 1 atomic mass unit

radioactive dating: method of determining the age of fossils or rocks that is based on the rate of decay of radioisotopes

radioactive decay: process in which the unstable nucleus of a radioisotope becomes stable by emitting charged particles and energy and changing to another element

reactant: substance that starts a chemical reaction

reactants: the starting materials in a reaction

reaction rate: speed at which a chemical reaction occurs

reduction: the gain of electrons or decrease in oxidation state in a chemical reaction

reversible reaction: a reaction that can also proceed in the reverse direction

salt bridge: a u-shaped tube containing an electrolyte that connects two half-cells in an electrochemical cell scientific law: is a statement that summarizes the results of many observations and experiments

scientific method: is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions

scientific notation: is a shorthand method for writing very large and very small numbers are expressed as the product of a number between and multiplied by some power of 10

second: the SI unit for time

SI: The International System of Units (abbreviated "SI" from French: Le Système International d'unités) is the modern form of the metric system *(source: Wikipedia)*

solubility: property of a solid, liquid, or gaseous chemical substance called solute to dissolve in a solid, liquid, or gaseous solvent *(source: Wikipedia)*

solute: the substance in a solution present in the least amount

solution: a homogeneous mixture of substances

solvent: the substance in a solution present in the greatest amount

subatomic particles: particles that are smaller than the atom

subscript: part of the chemical formula that indicates the number of atoms of the preceding element

synthesis reaction: a reaction in which two or more reactants combine to make one product

system: the reactants and products in the reaction

temperature: the average kinetic energy of the particles that make up a material

threshold energy: the minimum amount of energy necessary for a reaction to take place

titration: the process in which a known concentration of base (or acid) is added to a solution of acid (or base) of unknown concentration

transition elements: groups 3 to 12 of the periodic table

valence electrons: the electrons in the outermost energy level of an atom

wavelength: is the distance between two corresponding points on adjacent waves
electromagnetic spectrum: full range of wavelengths of electromagnetic waves, from radio waves to gamma rays

THE PERIODIC TABLE OF THE ELEMENTS

element name / atomic number / symbol / average mass: Silver
47
Ag
107.87

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Hydrogen 1 H 1.01																	Helium 2 He 4.00
Lithium 3 Li 6.94	Beryllium 4 Be 9.01											Boron 5 B 10.81	Carbon 6 C 12.01	Nitrogen 7 N 14.01	Oxygen 8 O 16.00	Fluorine 9 F 19.00	Neon 10 Ne 20.18
Sodium 11 Na 22.99	Magnesium 12 Mg 24.31											Aluminum 13 Al 26.98	Silicon 14 Si 28.09	Phosphorus 15 P 30.97	Sulfur 16 S 32.07	Chlorine 17 Cl 35.45	Argon 18 Ar 39.95
Potassium 19 K 39.10	Calcium 20 Ca 40.08	Scandium 21 Sc 44.96	Titanium 22 Ti 47.88	Vanadium 23 V 50.94	Chromium 24 Cr 52.00	Manganese 25 Mn 54.94	Iron 26 Fe 55.85	Cobalt 27 Co 58.93	Nickel 28 Ni 58.69	Copper 29 Cu 63.55	Zinc 30 Zn 65.39	Gallium 31 Ga 69.72	Germanium 32 Ge 72.61	Arsenic 33 As 74.92	Selenium 34 Se 78.96	Bromine 35 Br 79.90	Krypton 36 Kr 83.80
Rubidium 37 Rb 85.47	Strontium 38 Sr 87.62	Yttrium 39 Y 88.91	Zirconium 40 Zr 91.22	Niobium 41 Nb 92.91	Molybdenum 42 Mo 95.94	Technetium 43 Tc (98)	Ruthenium 44 Ru 101.07	Rhodium 45 Rh 102.91	Palladium 46 Pd 106.42	Silver 47 Ag 107.87	Cadmium 48 Cd 112.41	Indium 49 In 114.82	Tin 50 Sn 118.71	Antimony 51 Sb 121.76	Tellurium 52 Te 127.60	Iodine 53 I 126.90	Xenon 54 Xe 131.29
Cesium 55 Cs 132.91	Barium 56 Ba 137.33	57-70 *	Hafnium 72 Hf 178.49	Tantalum 73 Ta 180.95	Tungsten 74 W 183.84	Rhenium 75 Re 186.21	Osmium 76 Os 190.23	Iridium 77 Ir 192.22	Platinum 78 Pt 195.08	Gold 79 Au 196.97	Mercury 80 Hg 200.59	Thallium 81 Tl 204.38	Lead 82 Pb 207.20	Bismuth 83 Bi 208.98	Polonium 84 Po (209)	Astatine 85 At (210)	Radon 86 Rn (222)
Francium 87 Fr (223)	Radium 88 Ra (226)	89-102 **	Rutherfordium 104 Rf (267)	Dubnium 105 Db (268)	Seaborgium 106 Sg (271)	Bohrium 107 Bh (272)	Hassium 108 Hs (270)	Meitnerium 109 Mt (276)	Darmstadtium 110 Ds (281)	Roentgenium 111 Rg (280)	Copernicium 112 Cn (285)	Ununtrium 113 Uut (284)	Ununquadium 114 Uuq (289)	Ununpentium 115 Uup (288)	Ununhexium 116 Uuh (293)	Ununseptium 117 Uus (294?)	Ununoctium 118 Uuo (294)

ALKALI METALS · ALKALINE EARTH METALS · TRANSITION METALS · POST-TRANSITION METALS · METALLOIDS · HALOGENS · NOBLE GASES

UNKNOWN CHEMICAL PROPERTIES

SYNTHETIC ELEMENTS

INNER TRANSITION METALS

*LANTHANIDES	Lanthanum 57 La 138.91	Cerium 58 Ce 140.12	Praseodymium 59 Pr 140.91	Neodymium 60 Nd 144.24	Promethium 61 Pm (145)	Samarium 62 Sm 150.36	Europium 63 Eu 151.97	Gadolinium 64 Gd 157.25	Terbium 65 Tb 158.93	Dysprosium 66 Dy 162.50	Holmium 67 Ho 164.93	Erbium 68 Er 167.26	Thulium 69 Tm 168.93	Ytterbium 70 Yb 173.04	Lutetium 71 Lu 174.97
**ACTINIDES	Actinium 89 Ac (227)	Thorium 90 Th 232.04	Protactinium 91 Pa 231.04	Uranium 92 U 238.03	Neptunium 93 Np (237)	Plutonium 94 Pu (244)	Americium 95 Am (243)	Curium 96 Cm (247)	Berkelium 97 Bk (247)	Californium 98 Cf (251)	Einsteinium 99 Es (252)	Fermium 100 Fm (257)	Mendelevium 101 Md (258)	Nobelium 102 No (259)	Lawrencium 103 Lr (262)

THIS BOOK BELONGS TO: