Introduction to Chemistry

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1. I can distinguish between observable (qualitative) and numeric (quantitative) data.
2. I can construct and analyze data tables and graphs.
3. I can identify independent, dependant, and controlled variables in an experiment description, data table or graph.
4. I can write a laboratory summary in a Claim-Evidence Format

Unit 2: The Structure of the Atom
Nature of Science Goal—Scientific understanding changes as new data is collected.

1. I can use atomic models to explain why theories may change over time.
2. I can identify the relative size, charge and position of protons, neutrons, and electrons in the atom.
3. I can find the number of protons, neutrons and electrons in a given isotope of an element if I am given a nuclear symbol or name of element and mass number.
4. I can describe the difference between atomic mass and mass number.
5. I can describe the relationship between wavelength, frequency, energy and color of light (photons).
6. I can describe the process through which the electrons give off photons (energy) and describe the evidence that electrons have specific amounts of energy.
7. I can identify an unknown element using a flame test or by comparison to an emission spectra.
8. I can write electron configurations for elements in the ground state.

Unit 3: The Organization of the Elements
Nature of Science Goal—Classification systems lead to better scientific understanding.

1. I can describe the advantages of Mendeleev's Periodic Table over other organizations.
2. I can compare the properties of metals, nonmetals, and metalloids.
3. I can determine the number of valence electrons for elements in the main block.
4. I can explain the similarities between elements within a group or family.
5. I can identify patterns found on the periodic table such as reactivity, atomic radius, ionization energy and electronegativity.

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1. I can indicate the type of bond formed between two atoms and give properties of ionic, covalent, metallic bonds and describe the properties of materials that are bonded in each of those ways.
2. I can compare the physical and chemical properties of a compound to the elements that form it.
3. I can predict the charge an atom will acquire when it forms an ion by gaining or losing electrons using the octet rule.
4. I can write the names and formulas of ionic compounds.

5. I can indicate the shape and polarity of simple covalent compounds from a model or drawing.
6. I can describe how hydrogen bonding in water affects physical, chemical, and biological phenomena.

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Nature of Science Goal— Mathematics is a tool to increase scientific understanding.
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1. I can use the terms solute and solvent in describing a solution.
2. I can sketch a solution, colloid, and suspension at the particle level.
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4. I can calculate concentration in terms of molarity and molality.
5. I can describe the colligative properties of solutions. (Boiling point elevation, Freezing point depression, Vapor pressure lowering) in terms of every day applications.
6. I can identify which solution of a set would have the lowest freezing point or highest boiling point.

Unit 7: Describing Chemical Reactions
Nature of Science Goal—Conservations laws are investigated to explore science relationships.
1. I can classify a change as chemical or physical and give evidence of chemical changes reactions.
2. I can describe the principles of collision theory and relate frequency, energy of collisions, and addition of a catalyst to reaction rate.
3. I can write a chemical equation to describe a simple chemical reaction.
4. I can balance chemical reactions and recognize that the number of atoms in a chemical reaction does not change.
5. I can classify reactions as synthesis, decomposition, single replacement, double replacement or combustion.
6. I can use molar relationships in a balanced chemical reaction to predict the mass of product produced in a simple chemical reaction that goes to completion.
7. I can explain the concept of dynamic equilibrium as it relates to chemical reactions.
8. I can describe whether reactants or products are favored in equilibrium when given the equilibrium constant.
9. I can predict the effect of adding or removing either a product or a reactant or the effect of changing temperature to shift equilibrium.

Unit 8: Describing Acids and Bases

Nature of Science Goal--Nature is moving toward equilibrium
1. I can describe properties of acids and bases and identify if a solution is acidic or basic.
2. I can calculate the pH of a solution.
3. I can write a neutralization reaction between an acid and base.
4. I can calculate the concentration of an acid or base from data collected in a titration.

Unit 9: Energy of Chemical Changes
Nature of Science Goal—Science provides technology to improve lives.
1. I can classify evidence of energy transformation (temperature change) as endothermic or exothermic.
2. I can describe how electrical energy can be produced in a chemical reaction and identify which element gained and which element lost electrons.
3. I can identify the parts of a battery, including anode, cathode, and salt bridge.

Unit 10: Nuclear Changes
Nature of Science Goal—Correct interpretation of data replaces fear and superstition.
1. I can compare the charge, mass, energy, and penetrating power of alpha, beta, and gamma radiation and recognize that of the products of the decay of an unstable nucleus include radioactive particles and wavelike radiation.
2. I can interpret graphical data of decay processes to determine half-life and the age of a radioactive substance.
3. I can compare and contrast the amount of energy released in a nuclear reaction to the amount of energy released in a chemical reaction.
4. I can describe the differences between fission and fusion.
5. I can describe scientific evidence that all matter in the universe has a common origin.

Chapter 1: Introduction to Chemistry & the Nature of Science

1.1: The Process of Science

Objectives

- Explain the necessity for experimentation
- In an experiment, identify the independent, dependent, and controlled variables.

Introduction

Socrates (469 B.C. - 399 B.C.), Plato (427 B.C. - 347 B.C.), and Aristotle (384 B.C. - 322 B.C.) are among the most famous of the Greek philosophers. Plato was a student of Socrates, and Aristotle was a student of Plato. These three were probably the greatest thinkers of their time. Aristotle's views on physical science profoundly shaped medieval scholarship, and his influence extended into the Renaissance (14th century - 16th century). Aristotle's opinions were the authority on nature until well into the 1300s. Unfortunately, many of Aristotle's opinions were wrong. It is not intended here to denigrate Aristotle's intelligence; he was without doubt a brilliant man. It was simply that he was using a method for determining the nature of the physical world that is inadequate for that task. The philosopher's method was logical thinking, not making observations on the natural world. This led to many errors in Aristotle's thinking on nature. Let's consider two of Aristotle's opinions as examples.

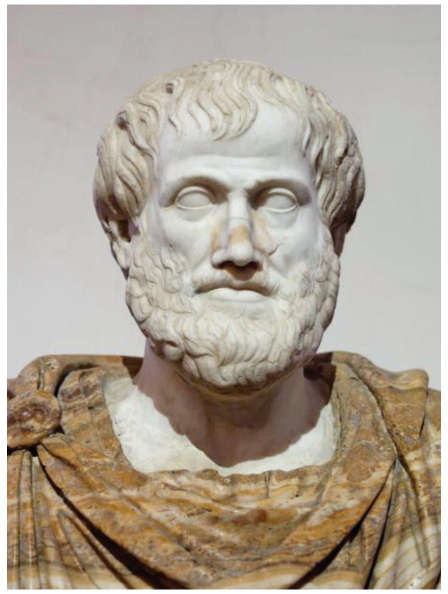


Image obtained from: http://upload.wikimedia.org/wikipedia/commons/a/ae/Aristotle_Altemps_Inv8575.jpg

In Aristotle's opinion, men were bigger and stronger than women; therefore, it was logical to him that men would have more teeth than women. Thus, Aristotle concluded it was a true fact that men had more teeth than women. Apparently, it never entered his mind to actually look into the mouths of both genders and count their teeth. Had he done so, he would have found that men and women have exactly the same number of teeth.

In terms of physical science, Aristotle thought about dropping two balls of exactly the same size and shape but of different masses to see which one would strike the ground first. In his mind, it was clear that the heavier ball would fall faster than the lighter one and he concluded that this was a law of nature. Once again, he did not consider doing an experiment to see which ball fell faster. It was logical to him, and in fact, it still seems logical. If someone told you that the heavier ball would fall faster, you would have no reason to disbelieve it. In fact, it is not true and the best way to prove this is to try it.

Eighteen centuries later, Galileo decided to actually get two balls of different masses, but with the same size and shape, and drop them off a building (Legend says the Leaning Tower of Pisa), and actually see which one hit the ground first. When Galileo actually did the experiment, he discovered, by observation, that the two balls hit the ground at exactly the same time . . . Aristotle's opinion was, once again, wrong.

Scientific Methods of Problem Solving

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

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. Other scientists oppose this listing of steps because not all steps occur in every case, and sometimes the steps are out of order. The scientific method is listed in a series of steps here because it makes it easier to study. You should remember that not all steps occur in every case, nor do they always occur in order.

The Steps in the Scientific Method

- **Step 1:** Identify the problem or phenomenon that needs explaining. This is sometimes referred to as "defining the problem."
- **Step 2:** Gather and organize data on the problem. This step is also known as "making observations."
- **Step 3:** Suggest a possible solution or explanation. A suggested solution is called a **hypothesis**.
- **Step 4:** Test the hypothesis by making new observations.
- **Step 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 return to Step 3.

Experimentation

Experimentation is the primary way through which science gathers evidence for ideas. It is more successful for us to cause something to happen at a time and place of our choosing. When we arrange for the phenomenon to occur at our convenience, we can have all our measuring instruments present and handy to help us make observations, and we can control other variables. Experimentation involves causing a phenomenon to occur when and where we want it and under the conditions we want. An **experiment** is a controlled method of testing an idea or to find patterns. When scientists conduct experiments, they are usually seeking new information or trying to verify someone else's data.

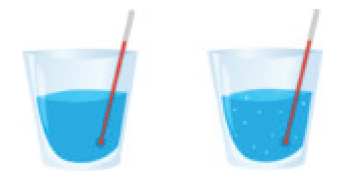
Experimentation involves changing and looking at many variables. The **independent variable** is the part of the experiment that is being changed or manipulated. There can only be one independent variable in any experiment. Consider, for example, that you were trying to determine the best fertilizer for your plants. It would be important for you to grow your plants with everything else about how they are grown being the same except for the fertilizer you were using. You would be changing the type of fertilizer you gave the plants and this would be the independent variable. If you also changed how much water the plants received, the type of plants you were growing, and some of the plants were grown inside and others outside, you could not determine whether or not it was actually the fertilizer that caused the plants to grow better or if it was something else you had changed. This is why it is important that there is only one independent variable.

The **dependent variable** is what is observed or measured as a result of what happened when the independent variable was changed. In the plant experiment described above, you might measure the height of the plant and record their appearance and color. These would be the dependent variables. The dependent variable is also sometimes called the resultant variable.

Controlled variables are conditions of the experiment that are kept the same for various trials of the experiment. Once again, if we were testing how fertilizer affected how well our plants grew, we would want everything else about how the plants are grown to be kept the same. We would need to use the same type of plant (maybe green beans), give them the same amount of water, plant them in the same location (all outside in the garden), give them all the same pesticide treatment, etc. These would be controlled variables.

Suppose a scientist, while walking along the beach on a very cold day following a rainstorm, observed two pools of water in bowl shaped rocks near each other. One of the pools was partially covered with ice, while the other pool had no ice on it. The unfrozen pool seemed to be formed from seawater splashing up on the rock from the surf, but the other pool was too high for seawater to splash in, so it was more likely to have been formed from rainwater.

The scientist wondered why one pool was partially frozen and not the other, since both pools were at the same temperature. 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, and she decided to go home and try an experiment to see if this were true. So far, the scientist has identified a question, gathered a small amount of data, and suggested an explanation. In order to test this hypothesis, the scientist will conduct an experiment during which she can make accurate observations.



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 liquid and these were put in a freezer. The scientist then observed the conditions and temperatures of the two liquids at regular intervals.

The Temperature and Condition of Fresh Water in a Freezer

Time (min)	Temp (°C)	Condition
0	25	Liquid
5	20	Liquid
10	15	Liquid
15	10	Liquid
20	5	Liquid
25	0	Frozen
30	-5	Frozen

The Temperature and Condition of Salt Water in a Freezer

Time (min)	Temp (°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 the hypothesis from this experiment; fresh water freezes at a higher temperature than salt water. Much more support would be needed before the scientist would be confident of this hypothesis. Perhaps she would 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 changing air pressure will alter the freezing point of water, so this and other variables must be kept the same, or they must be controlled variables.

Example: In the experiment described above, identify the:
a) independent variable(s)
b) dependent variable(s)
c) controlled variable(s)

Solution:
a) Remember, the independent variable is what the scientist changed in his/her experiment. In this case, the scientist added salt to one container and not to another container. The independent variable is whether or not salt was added.
b) Dependent variables are what we look for as a result of the change we made. The scientist recorded the temperature and physical state (liquid or solid) over time. These are the dependent variables.
c) Controlled variables are kept the same throughout all of the trials. The scientist selected identical containers, put the same amount of water in the containers, and froze them in the same conditions in the same freezer. These are all controlled variables.

Suppose you wish to determine which brand of microwave popcorn (independent variable) leaves the fewest unpopped kernels (dependent variable). You will need a supply of various brands of microwave popcorn to test and you will need a microwave oven. If you used different brands of microwave ovens with different brands of popcorn, the percentage of unpopped kernels could be caused by the different brands of popcorn, but it could also be caused by the different brands of ovens. Under such circumstances, the experimenter would not be able to conclude confidently whether the popcorn or the oven caused the difference. To eliminate this problem, you must use the same microwave oven for every test. By using the same microwave oven, you control many of the variables in the experiment. What if you allowed the different samples of popcorn to be cooked at different temperatures? What if you allowed longer heating periods? In order to reasonably conclude that the change in one variable was caused by the change in another specific variable, there must be no other variables in the experiment. All other variables must be kept constant or controlled.

When stating the purpose of an experiment, it is important to clarify the independent and dependent variables. The purpose is frequently stated in a sentence such as:

"To see how changing _____________ *affects* ____________*."*

in which the independent variable is listed in the first blank, and the dependent variable is listed in the second blank.

In the popcorn experiment, we would state the purpose as: "To see how changing the brand of popcorn affects the percentage of unpopped kernels". The independent variable is the brand of popcorn and the dependent variable is what percentage of the popcorn didn't pop. In the salt water experiment described earlier, we would state the purpose as "To see how adding salt to water affects the temperature the water freezes."

Lesson Summary

- Scientists use experimentation to test their ideas.
- In an experiment, it is important to include only one independent variable (to change only one thing in the experiment)
- The dependent variable is what is measured or observed as a result of how the independent variable changed.
- Controlled variables are those which are kept the same throughout various trials in the experiment.

Vocabulary

- Experiment: A controlled method of testing a hypothesis.
- Controlled experiment: An experiment that compares the results of an experimental sample to a control sample.

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 two videos that apply to this lesson. One is a video called ***The World of Chemistry*** that relates chemistry to other sciences and daily life. Another video called ***Thinking Like Scientists*** relates to the scientific method. The audience on the video is young children but the ideas are full grown.
- Website of the James Randi Foundation. James Randi is a staunch opponent of fake science. http://www.randi.org/site/
- Websites dealing with the history of the scientific method.
  http://www.historyguide.org/earlymod/lecture10c.html
  http://www.history.boisestate.edu/WESTCIV/science/

1.1: Review Questions

Use the following paragraph to answer questions 1-4:

Gary noticed that two plants which his mother planted on the same day that were the same size when planted were different in size after three weeks. Since the larger plant was in the full sun all day and the smaller plant was in the shade of a tree most of the day, Gary believed the sunshine was responsible for the difference in the plant sizes. In order to test this, Gary bought ten small plants of the same size and type. He made sure they had the same size and type of pot. He also made sure they have the same amount and type of soil. Then Gary built a frame to hold a canvas roof over five of the plants while the other five were nearby but out in the sun. Gary was careful to make sure that each plant received exactly the same amount of water and plant food every day.

1) What scientific reason might Gary have for insisting that the container size for the all plants be the same?
   a) Gary wanted to determine if the size of the container would affect the plant growth.

   b) Gary wanted to make sure the size of the container did not affect plant growth in his experiment.
   c) Gary wanted to control how much plant food his plants received.
   d) Gary wanted his garden to look organized.
   e) There is no possible scientific reason for having the same size containers.
2) What scientific reason might Gary have for insisting that all plants receive the same amount of water every day?
   a) Gary wanted to test the effect of shade on plant growth and therefore, he wanted to have no variables other than the amount of sunshine on the plants.
   b) Gary wanted to test the effect of the amount of water on plant growth.
   c) Gary's hypothesis was that water quality was affecting plant growth.
   d) Gary was conserving water.
   e) There is no possible scientific reason for having the same amount of water for each plant every day.
3) What was the variable being tested in Gary's experiment (what is the independent variable)?
   a) The amount of water
   b) The amount of plant food
   c) The amount of soil
   d) The amount of sunshine
   e) The type of soil
4) Which of the following factors may be varying in Gary's experimental setup that he did not control?
   a) Individual plant variation
   b) Soil temperature due to different colors of containers
   c) Water loss due to evaporation from the soil
   d) The effect of insects which may attack one set of plants but not the other

5) A student decides to set up an experiment to determine the relationship between the growth rate of plants and the presence of detergent in the soil. He sets up 10 seed pots. In five of the seed pots, he mixes a precise amount of detergent with the soil. The other five seed pots have no detergent in the soil. The five seed pots with detergent are placed in the sun and the five seed pots with no detergent are placed in the shade. All 10 seed pots receive the same amount of water and the same number and type of seeds. He grows the plants for two months and charts the growth every two days. What is wrong with his experiment?
   a) The student has too few pots.
   b) The student has two independent variables.
   c) The student has two dependent (resultant) variables.
   d) The student has no experimental control on the soil.

A scientist plants two rows of corn for experimentation. She puts fertilizer on row 1 but does not put fertilizer on row 2. Both rows receive the same amount of sun and water. She checks the growth of the corn over the course of five months.

6) What is the independent variable in this experiment?
7) What is the dependent variable in this experiment?
8) What variables are controlled in this experiment?

1.2: Hypothesis, Law, & Theory

Objectives

- Describe the difference between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.
- Explain the concept of a model.
- Explain why scientists use models.
- Explain the limitations of models as scientific representations of reality.

Introduction

Although all of us have taken science classes throughout the course of our study, many people have incorrect or misleading ideas about some of the most important and basic principles in science. We have all heard of hypotheses, theories, and laws, but what do they really mean? Before you read this section, think about what you have learned about these terms before. What do these terms mean to you? What do you read contradicts what you thought? What do you read supports what you thought?

Hypotheses

One of the most common terms used in science classes is a "hypothesis". The word can have many different definitions, depending on the context in which it is being used:

- "An educated guess" – because it provides a suggested solution based on the evidence. Note that it isn't just a random guess. It has to be based on evidence to be a scientific hypothesis.
- Prediction – if you have ever carried out a science experiment, you probably made this type of hypothesis, in which you predicted the outcome of your experiment.
- Tentative or Proposed explanation – hypotheses can be suggestions about why something is observed, but in order for it to be scientific, we must be able to test the explanation to see if it works, if it is able to correctly predict what will happen in a situation, such as: if my hypothesis is correct, we should see ___ result when we perform ___ test. A hypothesis is very tentative; it can be easily changed.

Theories

The *United States National Academy of Sciences* describes what a theory is as follows:

> "Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."
>
> "A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses" but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that

> everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact. "

Note some key features of theories that are important to understand from this description:

- Theories are explanations of natural phenomenon. They aren't predictions (although we may use theories to make predictions). They are explanations why we observe something.
- Theories aren't likely to change. They have so much support and are able to explain satisfactorily so many observations, that they are not likely to change. Theories can, indeed, be facts. Theories can change, but it is a long and difficult process. In order for a theory to change, there must be many observations or evidence that the theory cannot explain.
- Theories are not guesses. The phrase "just a theory" has no room in science. To be a scientific theory carries a lot of weight; it is not just one person's idea about something.

Laws

Scientific laws are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are typically well-supported by observations and/or experimental evidence. Usually scientific laws refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories are more overarching explanations of how nature works and why it exhibits certain characteristics. As a comparison, theories explain why we observe what we do and laws describe what happens.

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after many, many tests, that certain patterns existed in the observations on gas behavior. If the temperature of the gas increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do describe why the patterns exist.

A common misconception is that scientific theories are rudimentary ideas that will eventually graduate into scientific laws when enough data and evidence has been accumulated. A theory does not change into a scientific law with the accumulation of new or better evidence. Remember, theories are explanations and laws are patterns we see in large amounts of data, frequently written as an equation. A theory will always remain a theory; a law will always remain a law.

A **model** is a description, graphic, or 3-D representation of 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. 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, or as complex as mathematical equations representing light.

Chemists rely on both careful observation and well-known physical laws. By putting observations and laws together, chemists develop models. Models are really just ways of predicting what will happen given a certain set of circumstances. Sometimes these models are mathematical, but other times, they are purely descriptive.

If you were asked to determine the contents of a box that cannot be opened, you would 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/or determine its mass. When you completed your experiments, you would develop an idea of what is inside; that is, you would make a model of what is inside a box that cannot be opened.

A good example of how a model is useful to scientists is how models were used to explain the development of the atomic theory. As you will learn in a later chapter, the idea of the concept of an atom changed over many years. In order to understand each of the different theories of the atom according to the various scientists, models were drawn, and the concepts were more easily understood.

Chemists make up models about what happens when different chemicals are mixed together, or heated up, or cooled down, or compressed. Chemists invent these models using many observations from experiments in the past, and they use these models to predict what might happen during experiments in the future. Once chemists have models that predict the outcome of experiments reasonably well, those working models can help to tell them what they need to do to achieve a certain desired result. That result might be the production of an especially strong plastic, or it might be the detection of a toxin when it's present in your food.

Lesson Summary

- A hypothesis is a tentative explanation that can be tested by further investigation.
- A theory is a well-supported explanation of observations.
- A scientific law is a statement that summarizes the relationship between variables.
- An experiment is a controlled method of testing a hypothesis.
- A model is a description, graphic, or 3-D representation of 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

- Hypothesis: A tentative explanation that can be tested by further investigation.
- Theory: A well-established explanation
- Scientific law: A statement that summarizes the relationship between variables.
- Model: A description, graphic, or 3-D representation of theory used to help enhance understanding.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki/Scientific_theory

- http://en.wikipedia.org/wiki/Hypothesis
- Video on Demand – Modeling the Unseen (http://www.learner.org/resources/series61.html?pop=yes&pid=793#)

1.2: Review Questions

Multiple Choice

1) A number of people became ill after eating oysters in a restaurant. Which of the following statements is a hypothesis about this occurrence?
   a) Everyone who ate oysters got sick.
   b) People got sick whether the oysters they ate were raw or cooked.
   c) Symptoms included nausea and dizziness.
   d) Bacteria in the oysters may have caused the illness.
2) If the hypothesis is rejected (proved wrong) by the experiment, then:
   a) The experiment may have been a success.
   b) The experiment was a failure.
   c) The experiment was poorly designed.
   d) The experiment didn't follow the scientific method.
3) A hypothesis is:
   a) A description of a consistent pattern in observations.
   b) An observation that remains constant.
   c) A theory that has been proven.
   d) A tentative explanation for a phenomenon.
4) A scientific law is:
   a) A description of a consistent pattern in observations.
   b) An observation that remains constant.
   c) A theory that has been proven.
   d) A tentative explanation for a phenomenon.
5) A well-substantiated explanation of an aspect of the natural world is a:
   a) Theory.
   b) Law.
   c) Hypothesis.
   d) None of these.
6) Which of the following words is closest to the same meaning as hypothesis?
   a) Fact
   b) Law
   c) Formula
   d) Suggestion
   e) Conclusion
7) Why do scientists sometimes discard theories?
   a) The steps in the scientific method were not followed in order.
   b) Public opinion disagrees with the theory.
   c) The theory is opposed by the church.
   d) Contradictory observations are found.
8) True/False: When a theory has been known for a long time, it becomes a law.

1.3: Graphing

Objectives

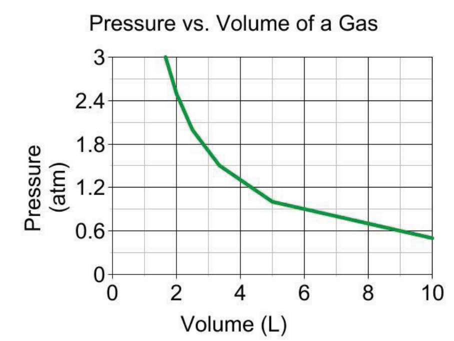
- Correctly graph data utilizing dependent variable, independent variable, scale and units of a graph, and best fit curve.
- Recognize patterns in data from a graph.
- Solve for the slope of given line graphs.

Introduction

Scientists search for regularities and trends in data. Two common methods of presenting data that aid in the search for regularities and trends are tables and graphs. The table below presents data about the pressure and volume of a sample of gas. You should note that all tables have a title and include the units of the measurements.

Volume (liters)	Pressure (atm)
10.0	0.50
5.0	1.00
3.33	1.50
2.50	2.00
2.00	2.50
1.67	3.00

You may note a regularity that appears in this table; as the volume of the gas decreases (gets smaller), its pressure increases (gets bigger). This regularity or trend becomes even more apparent in a graph of this data. A **graph** is a pictorial representation of patterns using a coordinate system. When the data from the table is plotted as a graph, the trend in the relationship between the pressure and volume of a gas sample becomes more apparent. The graph gives the scientist information to aid in the search for the exact regularity that exists in these data.



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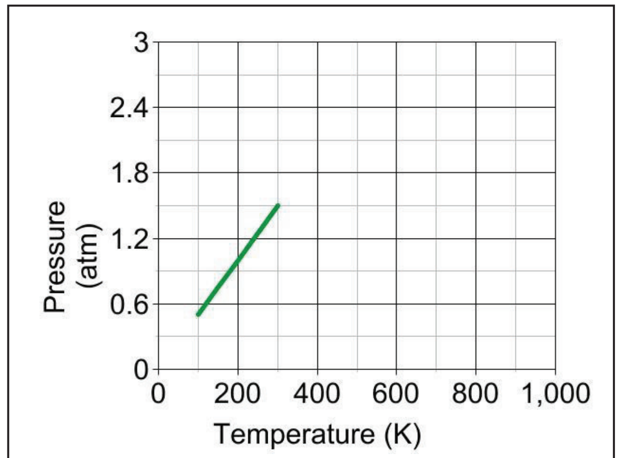
When scientists record their results in a data table, the independent variable is put in the first column(s), the dependent variable is recorded in the last column(s) and the controlled variables are typically not included at all. Note in the data table that the first column is labeled "Volume (in liters)" and that the second column is labeled "Pressure (in atm). That indicates that the volume was being changed (the independent variable) to see how it affected the pressure (dependent variable).

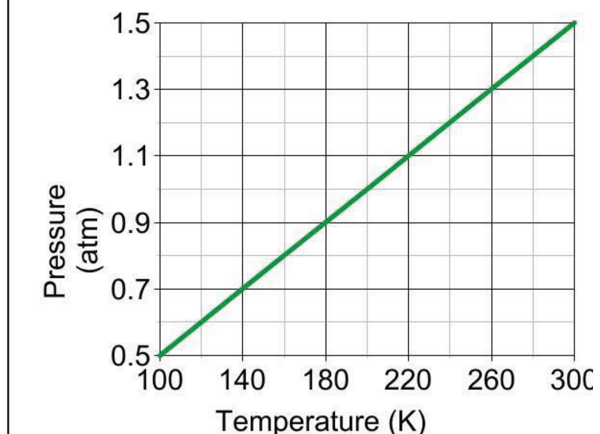
In a graph, the independent variable is recorded along the x-axis (horizontal axis) or as part of a key for the graph, the dependent variable is recorded along the y-axis (vertical axis), and the controlled variables are not included at all. Note in the data table that the X-axis is labeled "Volume (liters)" and that the Y-axis is labeled "Pressure (atm). That indicates that the volume was being changed (the independent variable) to see how it affected the pressure (dependent variable).

Drawing Line Graphs

Reading information from a line graph is easier and more accurate as the size of the graph increases. In the two graphs shown below, the first graph uses only a small fraction of the space available on the graph paper. The second graph uses all the space available for the same graph. If you were attempting to determine the pressure at a temperature of 260 K, using the graph on the left would give a less accurate result than using the graph on the right.

When you draw a line graph, you should arrange the numbers on the axis to use as much of the graph paper as you can. If the lowest temperature in your data is 100 K and the highest temperature in your data is 160 K, you should arrange for 100 K to be on the extreme left of your graph and 160 K to be on the extreme right of your graph. The creator of the graph on the left did not take this advice and did not produce a very good graph. You should also make sure that the axis on your graph are labeled and that your graph has a title.





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When constructing a graph, there are some general principles to keep in mind:

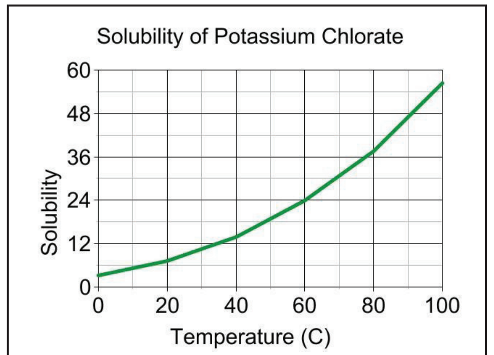
- Take up as much of the graph paper as possible. The lowest x-value should be on the far left of the paper and the highest x-value should be on the far right side of the paper. Your lowest y-value should be near the bottom of the graph and the highest y-value near the top. Choose your scale to allow you to do this. You do not need to start counting at zero.
- Count your x- and y-scales by consistent amounts. If you start counting your x-axis where every box counts as 2-units, you must count that way the course of the entire axis. Your y-axis may count by a different scale (maybe every box counts as 5 instead), but you must count the entire y-axis by that scale.
- Both of your axis should be labeled, including units. What was measured along that axis and what unit was it measured in?
- For X-Y scatter plots, draw a best-fit-line or curve that fits your data, instead of connecting the dots. You want a line that shows the overall trend in the data, but might not hit exactly all of your data points. What is the overall pattern in the data?

Reading Information from a Graph

When we draw a line graph from a set of data points, we are creating data points between known data points. This process is called **interpolation**. Even though we may have four actual data points that were measured, we assume the relationship that exists between the quantities at the actual data points also exists at all the points on the line graph between the actual data points. Consider the following set of data for the solubility of $KClO_3$ in water.

The table shows that there are exactly six known data points. When the data is graphed, however, the graph maker assumes that the relationship between the temperature and the solubility remains the same. The line is drawn by interpolating the data points between the actual data points.

Temperature (°C)	*Solubility (g/100 mL H_2O)*
0	3.3
20	7.3
40	13.9
60	23.8
80	37.5
100	56.3



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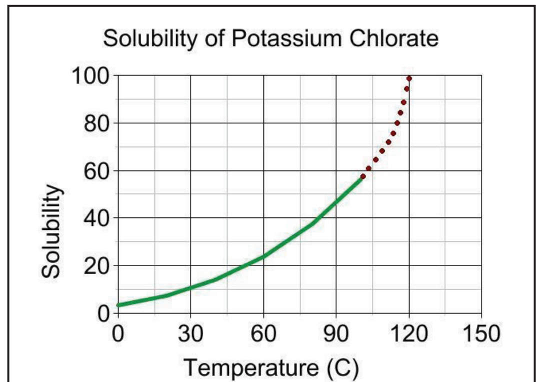
We can now reasonably certainly read data fr actually measured. If we wish to determine the solubility of $KClO_3$ at 70°C, we follow the vertical grid line for 70°C up to where it touches the graphed line and then follow the horizontal grid line to the axis to read the solubility. In this case, we would read the solubility to be 30. g/100 mL of H_2O at 70°C.

There are also occasions when scientists wish to determine data points from a graph that are not between actual data points but are beyond the ends of the actual data points. Creating data points beyond the end of the graph line, using the basic shape of the curve as a guide is called **extrapolation**.



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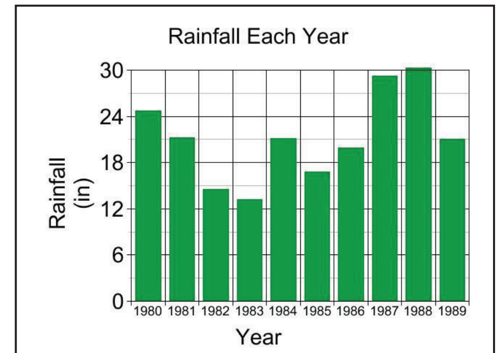
Suppose the graph for the solubility of potassium chlorate has been made from just three actual data points. If the actual data points for the curve were the solubility at 60°C, 80°C, and 100°C, the graph would be the solid line shown on the graph above. If the solubility at 30°C was desired, we could extrapolate (the dotted line) from the graph and suggest the solubility to be 5.0 g/100 mL of H_2O. If we check on the more complete graph above, you can see that the solubility at 30°C is close to 10 g/100 mL of H_2O. The reason the second graph produces such a poor answer is that the relationship that appears in the less complete graph does not hold beyond the ends of the graph. For this reason, extrapolation is only acceptable for graphs where there is evidence that the relationship shown in the graph will be true beyond the ends of the graph. Extrapolation is more dangerous that interpolation in terms of possibly producing incorrect data.

In situations in which both the independent and dependent variables are measured or counted quantities, an X-Y scatter plot is the most useful and appropriate type of graph. A line graph cannot be used for independent variables that are groups of data, or nonmeasured data. In these situations in which groups of data, rather than exact measurements, were recorded as the independent variable, a bar graph can typically be used. Consider the data in the following table.

For this data, a bar graph is more appropriate because independent variable is a group, not a measurement (for example, everything that happened in 1980). The concept of the average yearly rainfall halfway between the years 1980 and 1981 does not make sense, so a

line graph doesn't work. Additionally, each year represents a group that we are looking at, and not a measured quantity. A bar graph is better suited for this type of data. From this bar graph, you could very quickly answer questions like, "Which year was most likely a drought year for Trout Creek?", and "Which year was Trout Creek most likely to have suffered from a flood?"

Year	*Rainfall (inches)*
1980	24.7
1981	21.2
1982	14.5
1983	13.2
1984	21.1
1985	16.8
1986	19.9
1987	29.2
1988	31.6
1989	21.0



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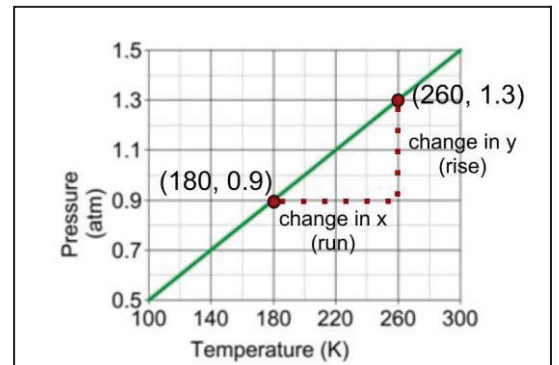
Finding the Slope of a Graph

As you may recall from algebra, the slope of the line may be determined from the graph. The slope represents the rate at which one variable is changing with respect to the other variable. For a straight-line graph, the slope is constant for the entire line but for a non-linear graph, the slope is different at different points along the line. For a straight-line graph, the slope for all points along the line can be determined from any section of the graph. For a non-linear graph, the must be determined for each point from data at that point. Consider the given data table and the linear graph that follows.

Temperature vs. Volume for a Gas

Temperature (°C)	Volume of Gas (mL)
20	60
40	65
60	70
80	75
100	80
120	85



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The relationship in this set of data is linear, that is, it produces a straight-line graph. The slope of this line is constant at all points on the line. The **slope** of a line is defined as the rise (change in vertical position) divided by the run (change in horizontal position).
Frequently in science, all of our data points do not fall exactly on a line. In this situation, we draw a **best fit line**, or a line that goes as close to all of our points as possible. When finding the slope, it is important to use two points that are on the best fit line itself, instead of our measured data points which may not be on our best fit line. For a pair of points on the line, the coordinates of the points are identified as (x_1, y_1) and (x_2, y_2). In this case, the points selected are (260, 1.3) and (180, 0.9). The slope can then be calculated in the manner:

$$slope = \frac{rise}{run} = \frac{(y_2 - y_1)}{(x_2 - x_1)} = \frac{(1.3 - 0.9)}{(260 - 180)} = 0.005\ atm/K$$

Therefore, the slope of the line is 0.005 atm/K. The fact that the slope is positive indicates that the line is rising as it moves from left to right and that the pressure increases by 0.005 atm for each 1 Kelvin increase in temperature. A negative slope would indicate that the line was falling as it moves from left to right.

Lesson Summary

- Two common methods of presenting data that aid in the search for regularities and trends are tables and graphs.
- When we draw a line graph from a set of data points, we are creating data points between known data points. This process is called interpolation.
- Creating data points beyond the end of the graph line, using the basic shape of the curve as a guide is called extrapolation.
- The slope of a graph represents the rate at which one variable is changing with respect to the other variable.

Vocabulary

- Graph: a pictorial representation of patterns using a coordinate system
- Interpolation: the process of estimating values between measured values
- Extrapolation: the process of creating data points beyond the end of the graph line, using the basic shape of the curve as a guide
- Slope: the ratio of the change in one variable with respect to the other variable.

Further Reading / Supplemental Links

- Use the following link to create both x-y and bar graphs: http://nces.ed.gov/nceskids/createagraph/default.aspx
- These websites offer more tips on graphing and interpreting data: http://staff.tuhsd.k12.az.us/gfoster/standard/bgraph2.htm and http://www.sciencebuddies.org/science-fair-projects/project_data_analysis.shtml

1.3: Review Questions

1) On a data table, where is the independent variable typically listed? What about the dependent variable?
2) On a graph, how do you identify the independent variable and dependent variable?
3) Andrew was completing his density lab for his chemistry lab exam. He collected the given data for volume and mass.
   a) Identify the independent and dependent variables in this experiment.
   b) Draw a graph to represent the data, including a best-fit-line.

#3 data	
Volume of Solution (mL)	**Mass of Solution (g)**
0.3	3.4
0.6	6.8
0.9	10.2
1.9	21.55
2.9	32.89
3.9	44.23
4.9	55.57

   c) If the graph is a straight line, calculate the slope, including units.
   d) What would you expect the mass of 2.5 mL of solution to have?
   e) What volume would you expect 60 g of the solution to occupy?

4) Donna is completing an experiment to find the effect of the concentration of ammonia on rate (or speed) of the reaction. She has collected the given data from her time trials and is ready for the analysis.
   a) Identify the independent and dependent variables in this experiment.
   b) Draw a graph to represent the data, including a best-fit-line
   c) If the concentration of ammonia was 0.30 mol/L, how much time has passed?
   d) After 8 seconds, what will be the approximate concentration of ammonia?

#4 data	
Time (s)	**Concentration of ammonia (mol/L)**
0.71	2.40
1.07	2.21
1.95	2.00
5.86	1.53
10.84	1.30
14.39	1.08
20.43	0.81
29.67	0.60
39.80	0.40
49.92	0.20

5) Consider the data table for an experiment on the behavior of gases.
   a) Identify the independent and dependent variables in this experiment.
   b) Draw a graph to represent the data.
   c) Calculate the slope, including units.
   d) What would be the pressure at 55°C?
   e) What would be the pressure at 120°C?

#5 data	
Temperature (°C)	**Pressure (mmHg)**
10	726
20	750
40	800
70	880
100	960

Chapter 2: The Structure of the Atom

2.1: Early Ideas of Atoms

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.

Introduction

You learned earlier how 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 and the Greek Philosophers

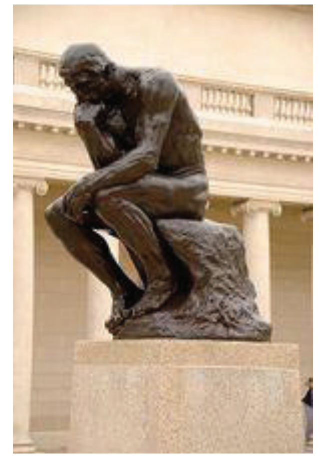
Before we discuss the experiments and evidence that have, over the years, convinced scientists that matter is made up of atoms, it's only fair to give credit to the man who proposed "atoms" in the first place. About 2,500 years ago, early Greek philosophers believed the entire universe was a single, 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."

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.

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.



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

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.

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 1700's, 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.

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?

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

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:

1. Matter is made of tiny particles called atoms.
2. 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.
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.

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=

2.1: 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?

2.2: 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.

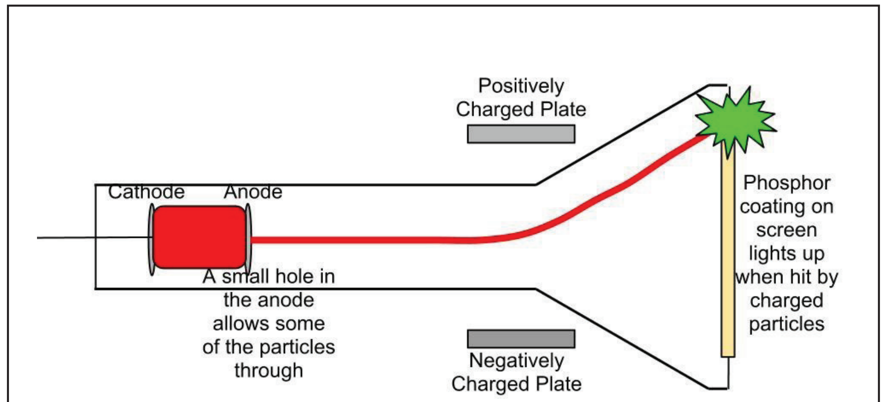
Thomson's Plum Pudding Model

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.

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!



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?
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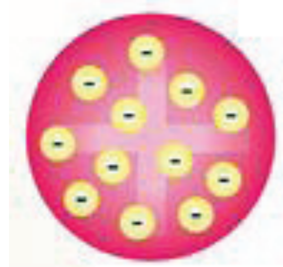
J. J. Thomson then did some rather complex experiments with magnets, and used his results to prove that cathode rays were not only negatively charged, but also had mass. Remember that anything with mass is part of what we call matter. In other words, these cathode rays must be the result of negatively 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.

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.)



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

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 1.6×10^{-19}C, the charge on a single electron.

Rutherford's Nuclear Model

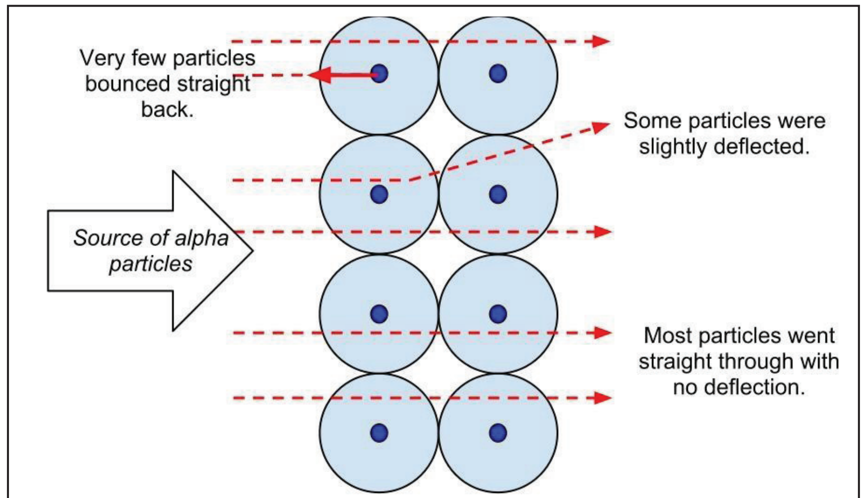
Everything about Thomson's experiments suggested the "plum pudding" model was correct – but according to the scientific method, any new theory or model should be tested by further experimentation and observation. In the case of the "plum pudding" model, it would take a man named Ernest Rutherford to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

Ernest Rutherford

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 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.



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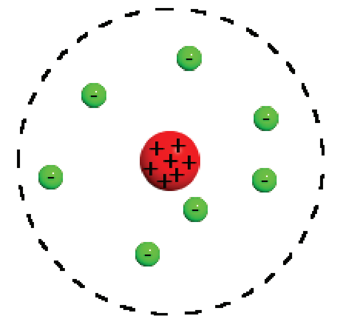
Every so often, though, one of the alpha particles would be deflected slightly as if it had bounced off of something hard. Even less often, Rutherford observed alpha particles bouncing straight back at the "gun" from which they had been fired! It was as if these alpha particles had hit a wall "head-on"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

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

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
5) All atoms have electrons	a. Most alpha particles shot at gold foil go straight through, without any change in their direction.
6) Atoms are mostly empty space.	b. A few alpha particles shot at gold foil bounce in the opposite direction.
7) Electrons have a negative charge	c. Some alpha particles (with positive charges) when shot through gold foil bend away from the gold.
8) 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.
9) 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.

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

11) Electrons are ______ negatively charged metals plates and ______ positively charged metal plates.

Consider the following two paragraphs for #12-14

Scientist 1: Although atoms were once regarded as the smallest part of nature, they are composed of even smaller particles. All atoms contain negatively charged particles, called electrons. However, the total charge of any atom is zero. Therefore, this means that there must also be positive charge in the atom. The electrons sit in a bed of positively charged mass.

Scientist 2: It is true that atoms contain smaller particles. However, the electrons are not floating in a bed of positive charge. The positive charge is located in the central part of the atom, in a very small, dense mass, called a nucleus. The electrons are found outside of the nucleus.

12) What is the main dispute between the two scientists' theories?
13) Another scientist was able to calculate the exact charge of an electron to be -1.6×10^{-19} C. What effect does this have on the claims of Scientist 1? (Pick one answer)
    a) Goes against his claim
    b) Supports his claim
    c) Has no effect on his claim.
14) If a positively charged particle was shot at a thin sheet of gold foil, what would the second scientist predict to happen?

2.3: 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.

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.

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

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.

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.

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.

$$\text{mass number } A = (\text{number of protons}) + (\text{number of neutrons})$$

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

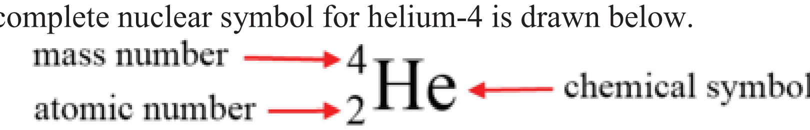
Example:
What is the mass number of an atom of helium that contains 2 neutrons?

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.

$^{59}_{28}Ni$ $^{238}_{92}U$

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)

2.3: Review Questions

Label each of the following statements as true or false.

1) The nucleus of an atom contains all of the protons in the atom.
2) The nucleus of an atom contains all of the electrons in the atom.
3) Neutral atoms must contain the same number of neutrons as protons.
4) Neutral atoms must contain the same number of electrons as protons.

Match the subatomic property with its description.

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

Indicate whether each statement is true or false.

8) An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
9) A neutral atom with 4 protons must have 4 electrons.
10) An atom with 7 protons and 7 neutrons will have a mass number of 14.
11) An atom with 7 protons and 7 neutrons will have an atomic number of 14.
12) A neutral atom with 7 electrons and 7 neutrons will have an atomic number of 14.

Use the periodic table to find the symbol for the element with:

13) 44 electrons in a neutral atom
14) 30 protons
15) An atomic number of 36

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
16) an atom with 2 protons and 1 neutron	a. a C (carbon) atom with 6 neutrons
17) a Be (beryllium) atom with 5 neutrons	b. an atom with 2 protons and 2 neutrons
18) 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
19) an atom with 1 proton and a mass number of 1	d. an atom with an atomic number of 1 and 1 neutron
20) an atom with an atomic number of 7 and 7 neutrons	e. an atom with an atomic number of 4 and 6 neutrons

Write the nuclear symbol for each element described:

21) 32 neutrons in an atom with mass number of 58

22) An atom with 10 neutrons and 9 protons.

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

23) ${}^{4}_{2}He$
24) Sodium-23
25) ${}^{1}_{1}H$
26) Iron-55
27) ${}^{37}_{17}Cl$
28) Boron-11
29) ${}^{238}_{92}U$
30) Uranium-235

2.4: Atomic Mass

Objectives:
- Explain what is meant by the atomic mass of an element.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Introduction

In chemistry we very rarely deal with only one isotope of an element. We use a mixture of the isotopes of an element in chemical reactions and other aspects of chemistry, because all of the isotopes of an element react in the same manner. That means that we rarely need to worry about the mass of a specific isotope, but instead we need to know the average mass of the atoms of an element. Using the masses of the different isotopes and how abundant each isotope is, we can find the average mass of the atoms of an element. The **atomic mass** of an element is the weighted average mass of the atoms in a naturally occurring sample of the element. Atomic mass is typically reported in atomic mass units.

Calculating Atomic Mass

You can calculate the atomic mass (or average mass) of an element provided you know the **relative abundances** (the fraction of an element that is a given isotope) the element's naturally occurring isotopes, and the masses of those different isotopes. We can calculate this by the following equation:

$$\textit{Atomic mass} = (\%_1)(mass_1) + (\%_2)(mass_2) + \ldots$$

Look carefully to see how this equation is used in the following examples.

Example: Boron has two naturally occurring isotopes. In a sample of boron, 20% of the atoms are B-10, which is an isotope of boron with 5 neutrons and a mass of 10 amu. The other 80% of the atoms are B-11, which is an isotope of boron with 6 neutrons and a mass of 11 amu. What is the atomic mass of boron?

Solution: Boron has two isotopes. We will use the equation:

$$\textit{Atomic mass} = (\%_1)(mass_1) + (\%_2)(mass_2) + \ldots$$

Isotope 1: $\%_1=0.20$ (write all percentages as decimals), $mass_1=10$
Isotope 2: $\%_2=0.80$, $mass_2=11$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.20)(10) + (0.80)(11)$$
$$\text{Atomic mass} = 10.8 \text{ amu}$$

The mass of an average boron atom, and thus boron's atomic mass, is 10.8 amu.

Example: Neon has three naturally occurring isotopes. In a sample of neon, 90.92% of the atoms are Ne-20, which is an isotope of neon with 10 neutrons and a mass of 19.99 amu. Another 0.3% of the atoms are Ne-21, which is an isotope of neon with 11 neutrons and a mass of 20.99 amu. The final 8.85% of the atoms are Ne-22, which is an isotope of neon with 12 neutrons and a mass of 21.99 amu. What is the atomic mass of neon?

Solution:
Neon has three isotopes. We will use the equation:

$$\textit{Atomic mass} = (\%_1)(mass_1) + (\%_2)(mass_2) + \ldots$$

Isotope 1: $\%_1=0.9092$ (write all percentages as decimals), $mass_1=19.99$
Isotope 2: $\%_2=0.003$, $mass_2=20.99$
Isotope 3: $\%_3=0.0885$, $mass_3=21.99$

Substitute these into the equation, and we get:

$$\text{Atomic mass} = (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99)$$
$$\text{Atomic mass} = 20.17 \text{ amu}$$

The mass of an average neon atom is **20.17 amu**

The periodic table gives the atomic mass of each element. The atomic mass is a number that usually appears below the element's symbol in each square. Notice that atomic mass of boron (symbol B) is 10.8, which is what we calculated in example 5, and the atomic mass of neon (symbol Ne) is 20.18, which is what we calculated in example 6. Take time to notice that not all periodic tables have the atomic number above the element's symbol and the mass number below it. If you are ever confused, remember that the atomic number should

always be the smaller of the two and will be a whole number, while the atomic mass should always be the larger of the two and will be a decimal number.

Lesson Summary

- An element's atomic mass is the average mass of one atom of that element. An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes, and the masses of those isotopes are known.
- The periodic table is a convenient way to summarize information about the different elements. In addition to the element's symbol, most periodic tables will also contain the element's atomic number, and element's atomic mass.

Vocabulary

- Atomic mass: the weighted average of the masses of the isotopes of an element

2.4: Review Questions

1) Copper has two naturally occurring isotopes. 69.15% of copper atoms are Cu-63 and have a mass of 62.93amu. The other 30.85% of copper atoms are Cu-65and have a mass of 64.93amu. What is the atomic mass of copper?

2) Chlorine has two isotopes, Cl-35 and Cl-37. Their abundances are 75.53% and 24.47% respectively. Calculate the atomic mass of chlorine.

2.5: The Nature of Light

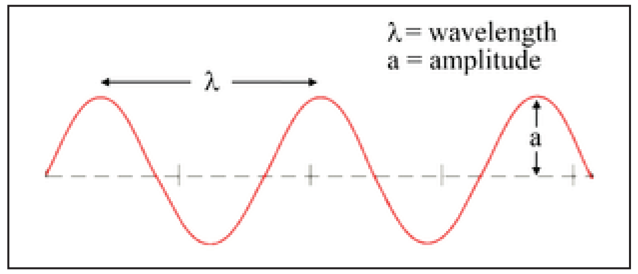
Objectives

- When given two comparative colors or areas in the electromagnetic spectrum, identify which area has the higher wavelength, the higher frequency, and the higher energy.
- Describe the relationship between wavelength, frequency, and energy of light waves (EMR)

Introduction

Most of us are familiar with waves, whether they are waves of water in the ocean, waves made by wiggling the end of a rope, or waves made when a guitar string is plucked. Light, also called **electromagnetic radiation**, is a special type of energy that travels as a wave.

Light Energy



Before we talk about the different forms of light or electromagnetic radiation (EMR), it is important to understand some of the general characteristics that waves share.

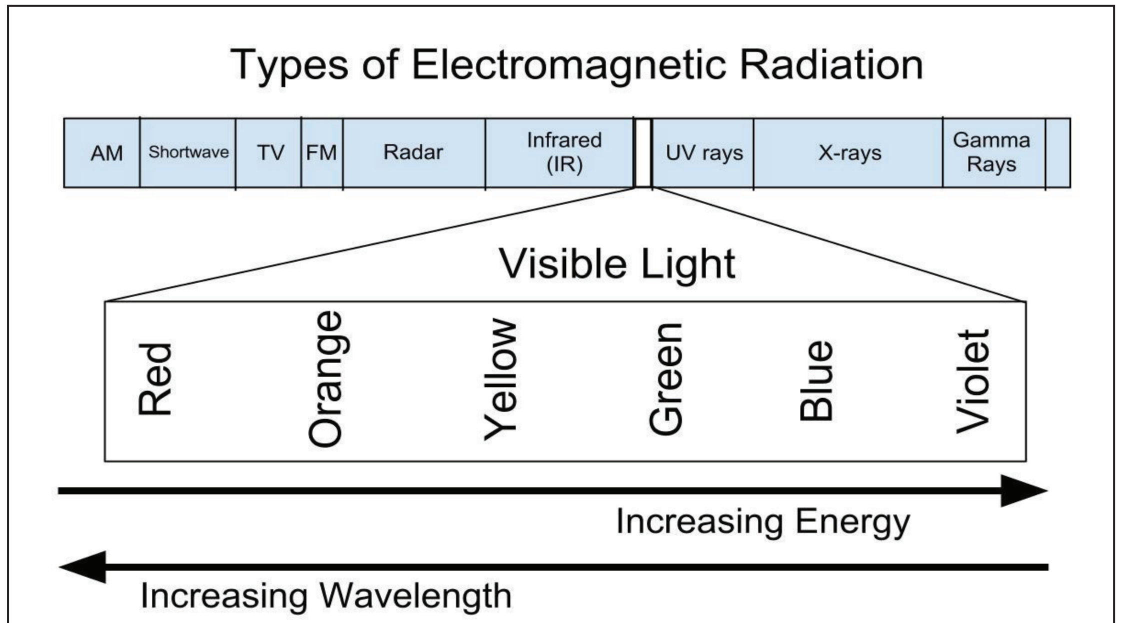
The high point of a wave is called the crest. The low point is called the trough. The distance from one point on a wave to the same point on the next wave is called the **wavelength** of the wave. You could determine the wavelength by measuring the distance from one trough to the next or from the top (crest) of one wave to the crest of the next wave. The symbol used for wavelength is the Greek letter lambda, λ.

Another important characteristic of waves is called frequency. The **frequency** of a wave is the number of waves that pass a given point each second. If we choose an exact position along the path of the wave and count how many waves pass the position each second, we would get a value for frequency. Frequency has the units of cycles/sec or waves/sec, but scientists usually just use units of 1/sec or Hertz (Hz).

All types of light (EMR) travels at the same speed, $3.00 \cdot 10^8$ m/s. Because of this, as the wavelength increases (the waves get longer), the frequency decreases (fewer waves pass). On the other hand, as the wavelength decreases (the waves get shorter), the frequency increases (more waves pass).

Electromagnetic waves (light waves) have an extremely wide range of wavelengths, frequencies, and energies. The **electromagnetic spectrum** is the range of all possible frequencies of electromagnetic radiation. The highest energy form of electromagnetic waves is gamma rays and the lowest energy form (that we have named) is radio waves.

On the far left of the figure above are the electromagnetic waves with the highest energy. These waves are called gamma rays and can be quite dangerous in large numbers to living systems. The next lowest energy form of electromagnetic waves is called x-rays. Most of you are familiar with the penetration abilities of these waves. They can also be dangerous to living systems. Next lower, in energy, are ultraviolet rays. These rays are part of sunshine and rays on the upper end of the ultraviolet range can cause sunburns and eventually skin



CC – Tracy Poulsen

cancer. The tiny section next in the spectrum is the visible range of light. These are the frequencies (energies) of the electromagnetic spectrum to which the human eye responds. The highest form of visible light energy is violet light, with red light having the lowest energy of all visible light. Even lower in the spectrum, too low in energy to see, are infrared rays and radio waves.

The light energies that are in the visible range are electromagnetic waves that cause the human eye to respond when those frequencies enter the eye. The eye sends signals to the brain and the individual “sees” various colors. The highest energy waves in the visible region cause the brain to see violet and as the energy of the waves decreases, the colors change to blue, green, then to yellow, orange, and red. When the energy of the wave is above or below the visible range, the eye does not respond to them. When the eye receives several different frequencies at the same time, the colors are blended by the brain. If all frequencies of light strike the eye together the brain sees white, and if there are no frequencies striking the eye the brain sees black.

All the objects that you see around you are light absorbers – that is, the chemicals on the surface of the objects absorb certain frequencies and not others. Your eyes detect the frequencies that strike your eye. Therefore, if your friend is wearing a red shirt, it means that the dye in that shirt absorbs every frequency except red and the red is reflected. When the red frequency from the shirt strikes your eye, your visual system sees red and you say the shirt is red. If your only light source was one exact frequency of blue light and you shined it on a shirt that absorbed every frequency of light except one exact frequency of red, then the shirt would look black to you because no light would be reflected to your eye. The light from many fluorescent types of light do not contain all the frequencies of sunlight and so clothes inside a store may appear to be a slightly different color than when you get them home.

Lesson Summary

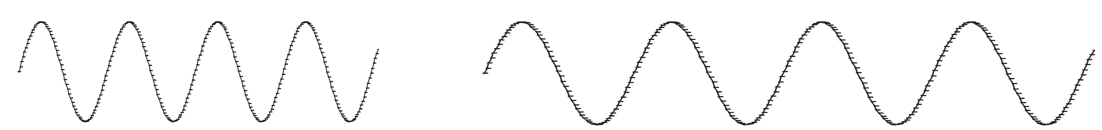
- Wave form energy is characterized by velocity, wavelength, and frequency.
- As the wavelength of a wave increases, its frequency decreases. Longer waves with lower frequencies have lower energy. Shorter waves with higher frequencies have higher energy.
- Electromagnetic radiation has a wide spectrum, including low energy radio waves and very high energy gamma rays.
- The different colors of light differ in their frequencies (or wavelengths).

Vocabulary

- Frequency of a wave: The number of waves passing a specific point each second.
- Wavelength: The distance between a point on one wave to the same point on the next wave (usually from crest to crest or trough to trough).
- Electromagnetic spectrum: A list of all the possible types of light in order of decreasing frequency, or increasing wavelength, or decreasing energy. The electromagnetic spectrum includes gamma rays, X-rays, UV rays, visible light, IR radiation, microwaves and radio waves.

2.5: Review Questions

1) Which color of visible light has the longer wavelength, red or blue?
2) What is the relationship between the energy of electromagnetic radiation and the frequency of that radiation?
3) Of the two waves drawn below, which one has the most energy? How do you know?



4) List the following parts of the electromagnetic spectrum in order of INCREASING energy: radio, gamma, UV, visible light, and infrared

5) List the visible colors of light in order of INCREASING energy.

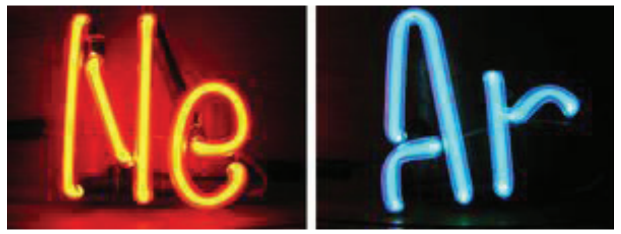
2.6: Atoms and Electromagnetic Spectra

Objectives

- Describe the appearance of an atomic emission spectrum.
- Explain that each element has a unique emission spectrum.
- Explain how an atomic (or emission) spectrum can be used to identify elements
- Describe an electron cloud that contains Bohr's energy levels.
- Explain the process through which an atomic spectrum is emitted according to Bohr's model of atoms.

Introduction

Electric light bulbs contain a very thin wire in them that emits light when heated. The wire is called a filament. The particular wire used in light bulbs is made of tungsten. A wire made of any metal would emit light under these circumstances but tungsten was chosen because the light it emits contains virtually every frequency and therefore, the light emitted by tungsten appears white. A wire made of some other element would emit light of some color that was not convenient for our uses. Every element emits light when energized by heating or passing electric current through it. Elements in solid form begin to glow when they are heated sufficiently and elements in gaseous form emit light when electricity passes through them. This is the source of light emitted by neon signs and is also the source of light in a fire.



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

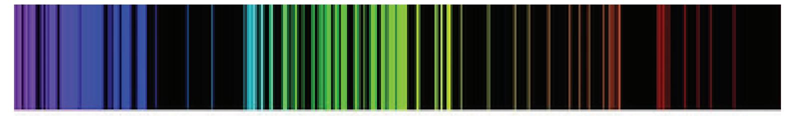
The light frequencies emitted by atoms are mixed together by our eyes so that we see a blended color. Several physicists, including Angstrom in 1868 and Balmer in 1875, passed the light from energized atoms through glass prisms in such a way that the light was spread out so they could see the individual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

The **emission spectrum** (or **atomic spectrum**) of a chemical element is the unique pattern of light obtained when the element is subjected to heat or electricity.


Atomic spectrum of hydrogen

When hydrogen gas is placed into a tube and electric current passed through it, the color of emitted light is pink. But when the color is spread out, we see that the hydrogen

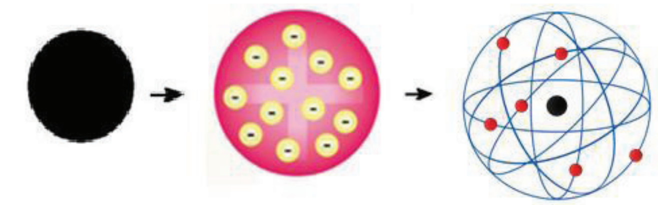
spectrum is composed of four individual frequencies. The pink color of the tube is the result of our eyes blending the four colors. Every atom has its own characteristic spectrum; no two atomic spectra are alike. The image below shows the emission spectrum of iron. Because each element has a unique emission spectrum, elements can be identified using them.


Atomic spectrum of iron.

You may have heard or read about scientists discussing what elements are present in the sun or some more distant star, and after hearing that, wondered how scientists could know what elements were 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 stars and finding the atomic spectrum of elements in that light. If the exact four lines that compose hydrogen's atomic spectrum are present in the light emitted from the star, that element contains hydrogen.

Bohr's Model of the Atom

By 1913, the evolution of our concept of the atom had proceeded from Dalton's indivisible spheres idea to J. J. Thomson's plum pudding model and then to Rutherford's nuclear atom theory.


Over time, our understanding of the atom has evolved. Dalton's model (on the left) was altered when Thomson discovered the electron and proposed the plum pudding model (in the middle). Rutherford discovered the nucleus and altered the model to the one on the right. Since then, Neils Bohr and other scientists discovered more about the location and energy of the electrons.

Rutherford, in addition to carrying out the brilliant experiment that demonstrated the presence of the atomic nucleus, also proposed that the electrons circled the nucleus in a planetary type motion. The solar system or planetary model of the atom was attractive to scientists because it was similar to something with which they were already familiar, namely the solar system.

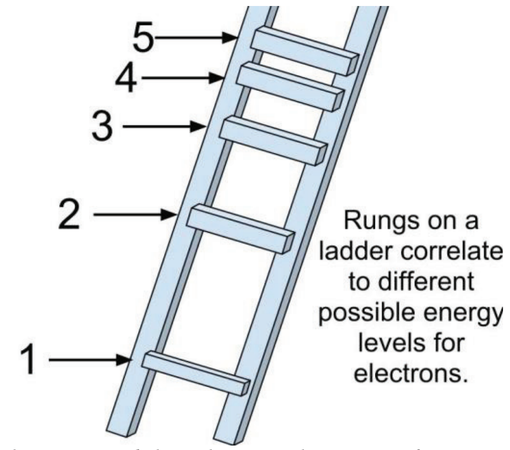
Unfortunately, there was a serious flaw in the planetary model. It was already known that when a charged particle (such as an electron) moves in a curved path, it gives off some form of light and loses energy in doing so. This is, after all, how we produce TV signals. If the electron circling the nucleus in an atom loses energy, it would necessarily have to move closer to the nucleus as it loses energy and would eventually crash into the nucleus. Furthermore, Rutherford's model was unable to describe how electrons give off light forming each element's unique atomic spectrum. These difficulties cast a shadow on the planetary model and indicated that, eventually, it would have to replaced.


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

In 1913, the Danish physicist Niels Bohr proposed a model of the electron cloud of an atom in which electrons orbit the nucleus and were able to produce atomic spectrum. Understanding Bohr's model requires some knowledge of electromagnetic radiation (or light).

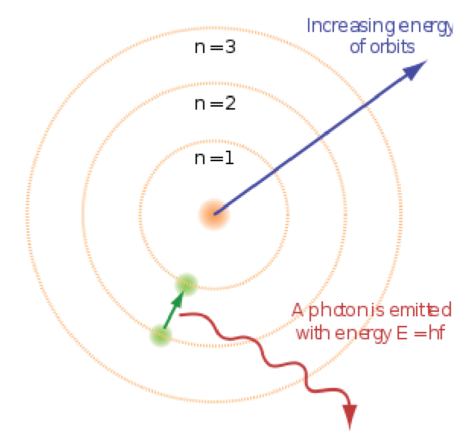
Energy Levels

Bohr's key idea in his 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 proposed that electrons have specific locations (or energy levels) around the nucleus, much like there are specific steps on a ladder. CC – Tracy Poulsen

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.



In Bohr's model of the atom, electrons absorb energy to move to higher energy levels and release energy to move to lower energy levels. Obtained from: http://en.wikipedia.org/wiki/File:Bohr_atom_model_English.svg

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 made by the electrons in an

atom moving 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**. The electrons then give off the energy in the form of a piece of light, called a **photon,** they had absorbed to fall back to a lower energy level. The energy emitted by electrons dropping back to lower energy levels would always be precise amounts of energy because the differences in energy levels were 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. This also explains why each element produces a different atomic spectrum. Because each element has different acceptable energy levels for their electrons, the possible steps each element's electrons can take differ from all other elements.

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 immediately dropping back to a lower energy level and emitting the energy difference between the two energy levels.
- The existence of the atomic spectra is support for Bohr's model of the atom.
- Bohr's model was only successful in calculating energy levels for the hydrogen atom.

Vocabulary

- Emission spectrum (or atomic spectrum): The unique pattern of light given off by an element when it is given energy
- Energy levels: Possible orbits that an electron can have in the electron cloud of an atom.
- Ground state: to be in the lowest energy level possible
- Excited state: to be in a higher energy level
- Photon: a piece of electromagnetic radiation, or light, with a specific amount of energy
- Quantized: having specific amounts

Further Reading / Supplemental Links

- *A short discussion of atomic spectra and some animation showing the spectra of elements you chose and an animation of electrons changing orbits with the absorbtion and emission of light can be viewed at* Spectral Lines (http://www.colorado.edu/physics/2000/quantumzone/index.html)
- Tutorial: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/linesp16.swf
- Tutorial: http://visionlearning.com/library/module_viewer.php?mid=51&l=&c3=
- Video: http://www.youtube.com/watch%3Fv%3DQI50GBUJ48s
- Fireworks & How Electrons Emit Photons video: http://www.youtube.com/watch%3Fv%3DncdmqhlTmGA

Review Questions

1) Bohr's model of the atom is frequently referred to as the "quantum model". Why? What does it mean to be quantized? How are electrons in atoms quantized?
2) Each element produces a unique pattern of light due to different energies within the atom. Why would this information be useful in analyzing a material?
3) It was known that an undiscovered element (later named helium) was on the sun before it was ever discovered on earth by looking at the sun's spectrum. How do scientists know that the sun contains helium atoms when no one has even taken a sample of material from the sun?
4) According to Bohr's theory, how can an electron gain or lose energy?
5) What happens when an electron in an excited atom returns to its ground level?
6) Why do electrons of an element release only a specific pattern of light? Why don't they produce all colors of light?
7) Use the following terms to explain how an electron releases a photon of light: electron, energy level, excited state, ground state, photon. Draw a picture if it is helpful.

2.6: Electron Arrangement in Atoms

Objectives

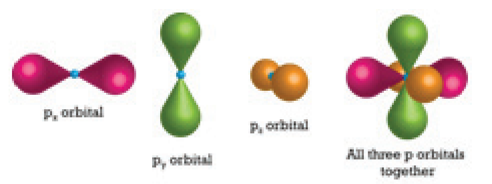
- List the order in which electron energy levels/sublevels will fill
- Write the electron configuration and abbreviated electron configuration for a given atom.

Introduction

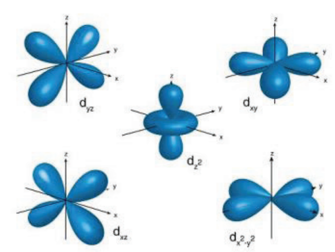
Chemists are particularly interested in the electrons in an atoms electron cloud. This is because the electrons determine the chemical properties of elements, such as what compounds the element will form and which reactions it will participate in. In this section, we will learn where the electrons are in atoms.

Electron Energy Levels

Although Bohr's model was particularly useful for hydrogen, it did not work well for elements larger than hydrogen. However, other physicists built on his model to create one that worked for all elements. It was found that the energy levels used for hydrogen were further composed of sublevels of different shapes. These sublevels were composed of orbitals in which the electrons were located.



The shape of p-orbitals



The shape of d-orbitals

The following table summarizes the possible energy levels and sublevels, including the number of orbitals that compose each sublevel and the number of electrons the sublevel can hold when completely filled.

Energy Level (related to the distance from the nucleus)	**Sublevel** (related to the shape)	**# of orbitals** in each sublevel	Maximum # of electrons possible
1	1s	1	2
2	2s	1	2
	2p	3	6
3	3s	1	2
	3p	3	6
	3d	5	10
4	4s	1	2
	4p	3	6
	4d	5	10
	4f	7	14
…	…	…	…

There are several patterns to notice when looking at the table of energy levels. Each energy level has one more sublevel than the level before it. Also, each new sublevel has two more orbitals. Can you predict what the 5th energy level would look like?

When determining where the electrons in an atom are located, a couple of rules must be followed:

1. Each added electron enters the orbitals of the lowest energy available.
2. No more than two electrons can be placed in any orbital.

The Electron Configuration

It would be convenient if the sublevels filled in the order listed in the table, such as 1s, 2s, 2p, 3s, 3p, 3d, 4s, 4p, etc. However, this is not the order the electrons fill the sublevels. Remember, the electrons will always go to the lowest energy available. When that is taken into account, the actual filling order is:

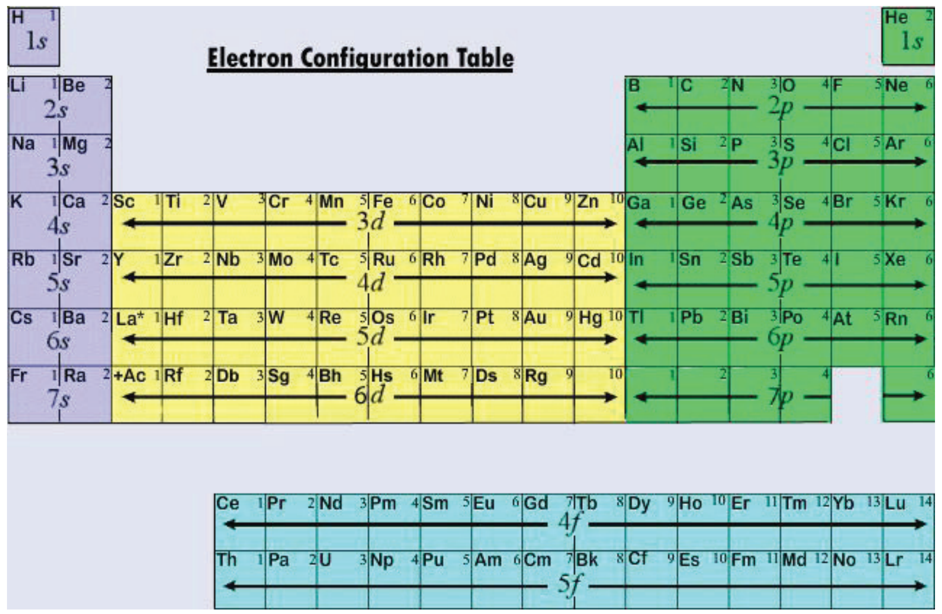
1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p…

Note that 4s has lower energy than 3d and, therefore, will fill first. The filling order gets more overlapped the higher you go.

An **electron configuration** lists the number of electrons in each used sublevel for an atom. For example, consider the element gallium, with 31 electrons. Its first two electrons would fit in the lowest energy possible, 1s. The next two would occupy 2s. 2p, with three orbitals, can hold its next 6 electrons. Gallium continues to fill up its orbitals, finally putting 1 electron in 4s. The electron configuration for gallium would be:

$${}_{31}\text{Ga: } 1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$$

Although you can choose to memorize the list and how many electrons fit in each sublevel for the purpose of writing electron configurations, there is a way for us to find this order by simply using our periodic table.



http://en.wikipedia.org/wiki/File:Electron_Configuration_Table.jpg

Look at the different sections of the periodic table. You may have noticed that there are several natural sections of the periodic table. The first 2 columns on the left make up the first section; the six columns on the right make up the next section; the middle ten columns make up another section; finally the bottom fourteen columns compose the last section. Note the significance of these numbers: 2 electrons fit in any *s* sublevel, 6 electrons fit in any *p* sublevel, 10 electrons fit in any *d* sublevel, and 14 electrons fit in any *f* sublevel. The four sections described previously are known as the *s*, *p*, *d*, and *f* blocks respectively.

If you move across the rows starting at the top left of the periodic table and move across each successive row, you can generate the same order of filling orbitals that was listed before and also how many electrons total fit in each orbital. Starting at the top left, you are filling 1s. Moving onto the second row, 2s is filled followed by 2p. Continuing with the filling order you generate the list:

1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p…

An electron configuration lists the sublevels the electrons occupy and the number of electrons in each of those sublevels, written as superscripts.

Example: Write the electron configurations for
(a) potassium, K
(b) arsenic, As
(c) phosphorus, P

Solution:
(a) Potassium atoms have 10 protons and, therefore, 19 electrons. Using our chart, we see that the first sublevel to fill is *1s*, which can hold 2 of those 19 electrons. Next to fill is *2s*, which also holds 2 electrons. Then comes *2p* which holes 6. We keep filling up the sublevels until all 19 of the electrons have been placed in the lowest energy level possible. *4s* only has 1 electron in it, although it can hold up to 2 electrons, because there are only 19 electrons total in potassium. Its electron configuration is written as: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$
(b) $_{33}$As: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$
(c) $_{15}$P: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^3$

Abbreviated Electron Configuration

As the electron configurations become longer and longer, it becomes tedious to write them out. A shortcut has been devised so that writing the configurations is less tedious. The electron configuration for potassium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^1$. The electron configuration for potassium is the same as the electron configuration for argon except that it has one more electron. The electron configuration for argon is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$ and in order to write the electron configuration for potassium, we need to add only $4s^1$. It is acceptable to use [Ar] to represent the electron configuration for argon and $[Ar]4s^1$ to represent the electron configuration for potassium. Using this shortcut, the abbreviated electron configuration for calcium would be $[Ar]\ 4s^2$ and the electron configuration for scandium would be $[Ar]4s^2\ 3d^1$.

Even though the periodic table was organized according to the chemical behavior of the elements, you can now see that the shape and design of the table is a perfect reflection of the electron configuration of the atoms. This is because the chemical behavior of the elements is also caused by the electron configuration of the atoms.

Example: Write the abbreviated electron configurations for
(a) potassium, K
(b) arsenic, As
(c) phosphorus, P

Solution:
(a) $_{19}$K: $[Ar]\ 4s^1$
(b) $_{33}$As: $[Ar]\ 4s^2\ 3d^{10}\ 4p^3$
(c) $_{15}$P: $[Ne]\ 3s^2\ 3p^3$

Lesson Summary

- Electrons are located in orbitals, in various sublevels and energy levels of atoms
- Electrons will occupy the lowest energy level possible.
- It is possible to write the electron configuration of en element using a periodic table.

Vocabulary

- Electron configuration: a list that represents the arrangement of electrons of an atom.

Further Reading / Supplemental Links

- http://www.ethbib.ethz.ch/exhibit/pauli/elektronenspin_e.html
- http://www.lorentz.leidenuniv.nl/history/spin/goudsmit.html
- http://en.wikipedia.org/wiki

2.7: Review Questions

1) Which principal energy level holds a maximum of eight electrons?
2) Which sub-energy level holds a maximum of six electrons?
3) Which sub-energy level holds a maximum of ten electrons?
4) If all the orbitals in the first two principal energy levels are filled, how many electrons are required?
5) In which energy level and sub-level of the carbon atom is the outermost electron located?
6) How many electrons are in the 2p sub-energy level of a neutral nitrogen atom?

7) Which element's neutral atoms will have the electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^1$?
8) What energy level and sub-level immediately follow 5s in the filling order?
9) What is the outermost energy level and sub-level used in the electron configuration of potassium?

Write electron configurations for each of the following neutral atoms:

10) Magnesium
11) Nitrogen
12) Yttrium
13) Tin
14) Krypton
15) Cesium
16) Uranium

Write the abbreviated electron configuration for each of the following neutral atoms:

17) Fluorine
18) Aluminum
19) Titanium
20) Arsenic
21) Rubidium
22) Carbon

Chapter 3: The Organization of the Elements

3.1: Mendeleev's Periodic Table

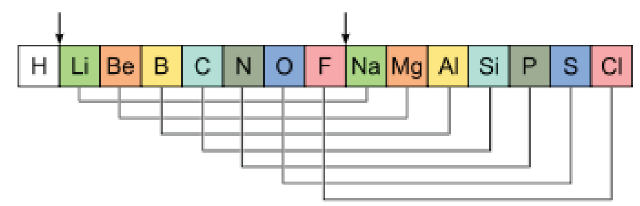
Objectives

- 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.

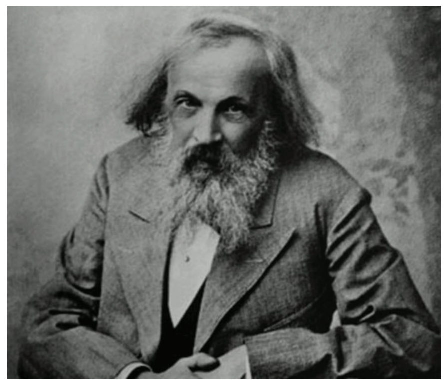
In 1864, John Newlands saw a connection between the chemical properties of elements and their atomic masses. He stated that if the known elements, beginning with lithium are arranged in order of increasing mass, the eighth element will have properties similar to the first, the ninth similar to the second, 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.



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

Mendeleev Organized His Table According to Chemical Behavior

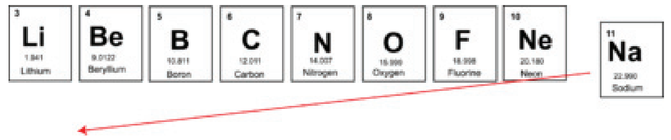
By 1869, a total of 63 elements had been discovered. As the number of known elements grew, scientists began to recognize patterns in the way chemicals reacted and began to devise ways to classify the elements. Dmitri Mendeleev, a Siberian-born Russian chemist, was the first scientist to make a periodic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



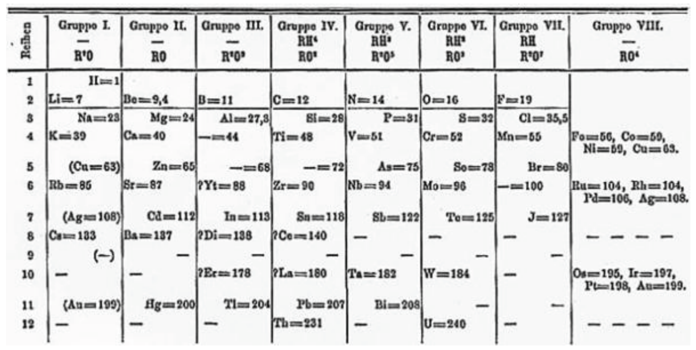
Dmitri Mendeleev created the first periodic table in 1869.

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.



Mendeleev avoided Newlands' mistake of trying to force elements into groups where their chemistry did not match, but still ran into a few problems as he constructed his table. Periodically, the atomic mass of elements would not be in the right order to put them in the correct group. For example, look at iodine and tellurium on your periodic table. Tellurium is heavier than iodine, but he put it before iodine in his table, because iodine has properties most similar to fluorine, chlorine, and bromine. Additionally, tellurium has properties more similar to the group with oxygen in it. On his table, he listed the element its place according to its properties and put a question mark (?) next to the symbol. The question mark indicated that he was unsure if the mass had been measured correctly.



Reihen	Gruppe I. — R_2O	Gruppe II. — RO	Gruppe III. — R_2O_3	Gruppe IV. RH_4 RO_2	Gruppe V. RH_3 R_2O_5	Gruppe VI. RH_2 RO_3	Gruppe VII. RH R_2O_7	Gruppe VIII. — RO_4
1	H=1							
2	Li=7	Be=9,4	B=11	C=12	N=14	O=16	F=19	
3	Na=23	Mg=24	Al=27,3	Si=28	P=31	S=32	Cl=35,5	
4	K=39	Ca=40	—=44	Ti=48	V=51	Cr=52	Mn=55	Fe=56, Co=59, Ni=59, Cu=63.
5	(Cu=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Yt=88	Zr=90	Nb=94	Mo=96	—=100	Ru=104, Rh=104, Pd=106, Ag=108.
7	(Ag=108)	Cd=112	In=113	Sn=118	Sb=122	Te=125	J=127	
8	Cs=133	Ba=137	?Di=138	?Ce=140	—	—	—	— — — —
9	(—)	—	—	—	—	—	—	
10	—	—	?Er=178	?La=180	Ta=182	W=184	—	Os=195, Ir=197, Pt=198, Au=199.
11	(Au=199)	Hg=200	Tl=204	Pb=207	Bi=208	—	—	
12	—	—	—	Th=231	—	U=240	—	— — — —

Mendeleev's 1869 periodic table

Another problem Mendeleev encountered was that sometimes the next heaviest element in his list did not fit the properties of the next available place on the table. He would skip places on the table, leaving holes, in order to put the element in a group with elements with similar properties. For example, at the time the elements Gallium and Germanium had not yet been discovered. After zinc, arsenic was the next heaviest element he knew about, but arsenic had properties most similar to nitrogen and phosphorus, not boron. He left two holes in his table for what he claimed were undiscovered elements. Note the dashes (-) with a mass listed after it in his original table. These indicate places in which he predicted elements would later be discovered to fit and his predicted mass for these elements.

Mendeleev went further with his missing elements by predicting the properties of elements in those spaces. In 1871 Mendeleev predicted the existence of a yet-undiscovered element he called eka-aluminium (because its location was directly under aluminum's on the table). The table below compares the qualities of the element predicted by Mendeleev with actual characteristics of Gallium (discovered in 1875).

Mendeleev made similar predictions for an element to fit in the place next to silicon. Germanium, isolated in 1882, provided the best confirmation of the theory up to that time, due to its contrasting more clearly with its neighboring elements than the two previously confirmed predictions of Mendeleev do with theirs.

How was Mendeleev able to make such accurate predictions? He understood the patterns that appeared between elements within a family, as well as patterns according to increasing mass, that he was able to fill in the missing pieces of the patterns. The ability to make accurate predictions is was put Mendeleev's table apart from other organization systems that were made at the same time and is what led to scientists accepting his table and periodic law.

Property	Mendeleev's prediction for Eka-aluminium	Actual properties of Gallium
atomic mass	68	69.72
density (g/cm³)	6.0	5.904
melting point (°C)	Low	29.78
oxide's formula	Ea_2O_3 (density - 5.5 g cm^{-3}) (soluble in both alkalis and acids)	Ga_2O_3 (density - 5.88 g cm^{-3}) (soluble in both alkalis and acids)
chloride's formula	Ea_2Cl_6 (volatile)	Ga_2Cl_6 (volatile)

Property	Mendeleev's predictions for Eka-silicon	Actual properties of Germanium
atomic mass	72	72.61
density (g/cm³)	5.5	5.35
melting point (°C)	high	947
color	grey	grey
oxide type	refractory dioxide	refractory dioxide
oxide density (g/cm³)	4.7	4.7
oxide activity	feebly basic	feebly basic
chloride boiling point	under 100°C	86°C ($GeCl_4$)
chloride density (g/cm³)	1.9	1.9

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)

3.1: 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.

3.2: Metals, Nonmetals, and Metalloids

Objectives

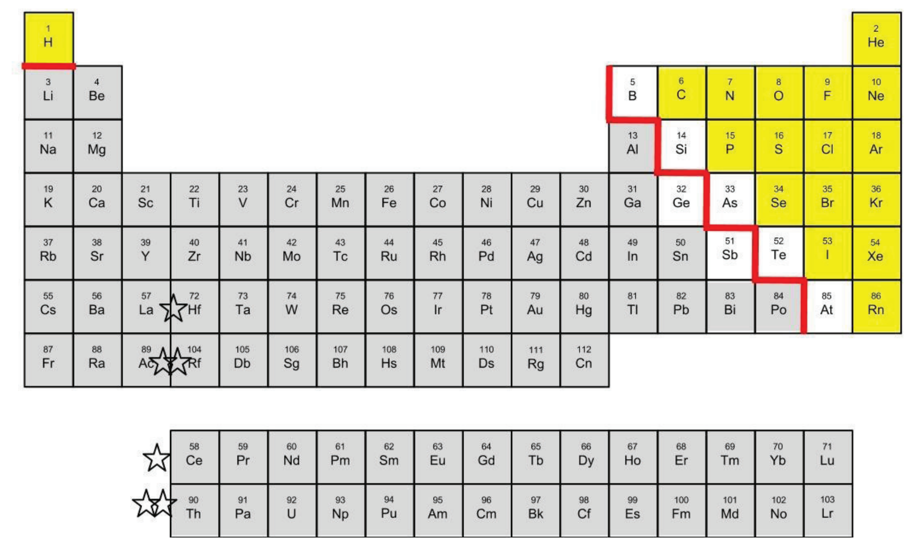
- Describe the differences among metals, nonmetals, and metalloids.
- Identify an element as a metal, nonmetal, or metalloid given a periodic table or its properties.

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



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.
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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.

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

3.2: Review Questions

Label each of the following elements as a metal, nonmetal, or metalloid.

1) Carbon
2) Bromine
3) Oxygen
4) Plutonium
5) Potassium
6) Helium

Given each of the following properties, label the property of as that of a metal, nonmetal, or metalloid.

7) Lustrous
8) Semiconductors
9) Brittle
10) Malleable
11) Insulators
12) Conductors
13) Along the staircase

14) 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?

3.3: Valence Electrons

Objectives

- Define valence electrons.
- Indicate the number of valence electrons for selected atoms.

Introduction

The electrons in the outermost shell are the **valence electrons** these are the electrons on an atom that can be gained or lost in a chemical reaction. Since filled *d* or *f* subshells are seldom disturbed in a chemical reaction, the valence electrons include those electrons in the outermost *s* and *p* sublevels.

Gallium has the following electron configuration.

Ga: $[Ar]\ 4s^2\ 3d^{10}\ 4p^1$

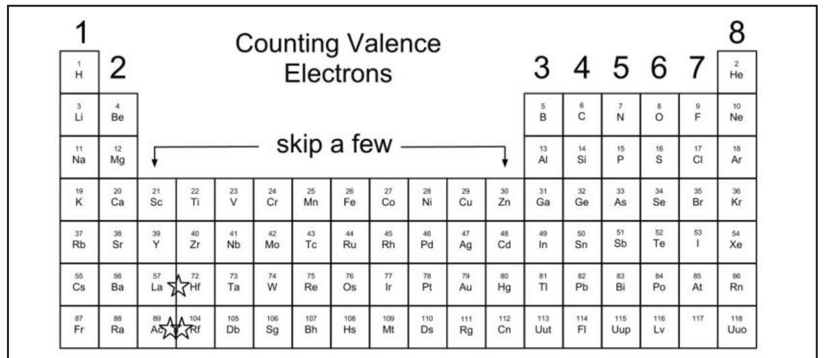
The electrons in the fourth energy level are further from the nucleus than the electrons in the third energy level. The *4s* and *4p* electrons can be lost in a chemical reaction, but not the electrons in the filled *3d* subshell. Gallium therefore has three valence electrons: the two in *4s* and one in *4p*.

Determining Valence Electrons

The number of valence electrons for an atom can be seen in the electron configuration. The electron configuration for magnesium is $1s^2\ 2s^2\ 2p^6\ 3s^2$. The outer energy level for this atom is $n=3$ and it has two electrons in this energy level. Therefore, magnesium has two valence electrons.

The electron configuration for sulfur is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^4$. The outer energy level in this atom is $n=3$ and it holds six electrons, so sulfur has six valence electrons.

The electron configuration for gallium is $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^1$. The outer energy level for this atom is $n=4$ and it contains three electrons. You must recognize that even though the 3d sub-level is mixed in among the 4s and 4p sub-levels, 3d is NOT in the outer energy level and therefore, the electrons in the 3d sub-level are NOT valence electrons. Gallium has three electrons in the outer energy level and therefore, it has three valence electrons. The identification of valence electrons is vital because the chemical behavior of an element is determined primarily by the arrangement of the electrons in the valence shell.



The number of valence electrons in an atom can be easily found by counting the s and p columns in the periodic table.
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This pattern can be summarized very easily, by merely counting the *s* and *p* blocks of the periodic table to find the total number of valence electrons. One system of numbering the groups on the periodic table numbers the *s* and *p* block groups from 1A to 8A. The number indicates the number of valence electrons.

Lesson Summary

- Valence electrons are the electrons in the outermost principal quantum level of an atom.
- The number of valence electrons is important, because the chemical behavior of an element depends primarily by the arrangement of the electrons in the valence shell.

Vocabulary

- Valence electrons: the electrons in the outermost energy level of an atom.

3.3: Review Questions

1) How many valence electrons are present in the following electron configuration: $1s^22s^22p^63s^23p^3$?
2) How many valence electrons are present in the following electron configuration: $1s^22s^22p^63s^23p^64s^23d^{10}4p^1$?

For each of the following atoms, indicate the total number of valence electrons in each atom:

3) Fluorine
4) Bromine
5) Sodium
6) Cesium
7) Aluminum
8) Gallium
9) Argon
10) Krypton

3.4: 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

The chemical behavior of atoms is controlled by their electron configuration. 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

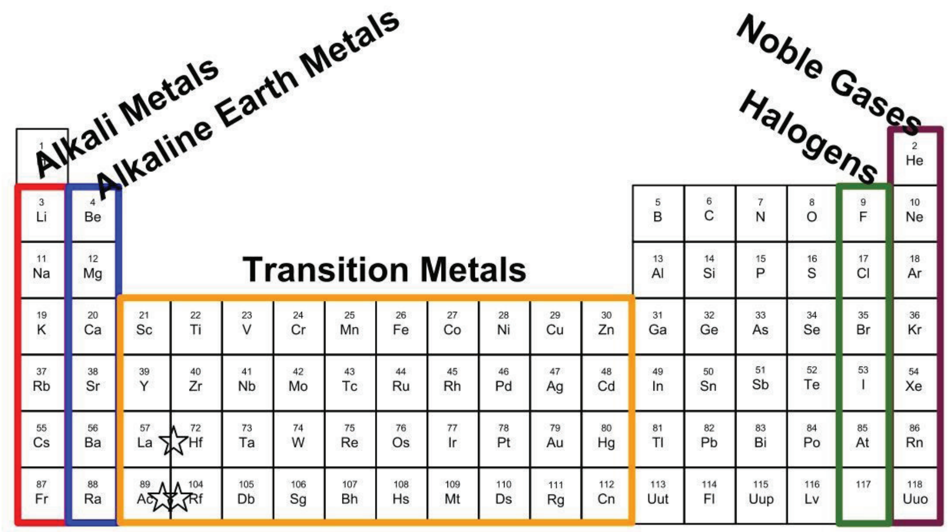
Remember that Mendeleev arranged the periodic table so that elements with the most similar properties were placed in the same group. A **group** is a vertical column of the periodic table. All of the 1A elements have one valence electron. This is what causes these elements to react in the same ways as 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.



Families of the periodic table.
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An alternate numbering system numbers all of the *s*, *p*, and *d* block elements 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.

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

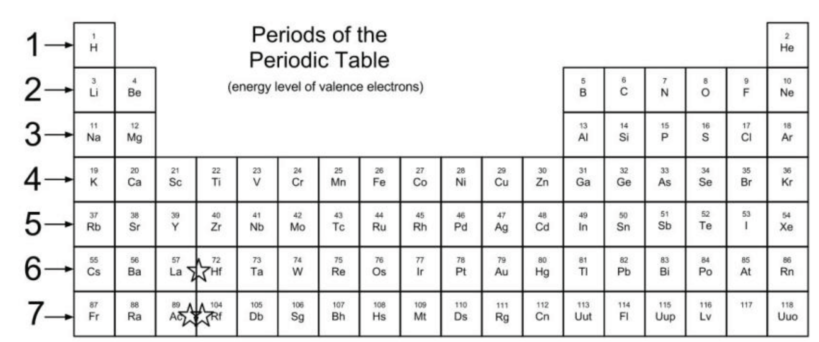

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table. For example, the elements sodium (Na) and magnesium (Mg) are both in period 3. The elements astatine (At) and radon (Rn) are both in period 6.

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

3.4: 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

Give the name of the family in which each of the following elements is located:

4) astatine
5) krypton
6) barium
7) francium

Which family is characterized by each of the following descriptions?

8) A very reactive family of nonmetals
9) Have 7 valence electrons
10) A nonreactive family of nonmetals
11) Forms colorful compounds
12) Have 2 valence electrons
13) A very reactive family of metals

3.5: Periodic Trends

Objectives

- Explain what is meant by the term periodic law
- Describe the general trend in atomic size for groups and periods.
- Describe the trends that exist in the periodic table for ionization energy.
- Describe the trends that exist in the periodic table for electronegativity.

Introduction

We have talked in great detail about how the periodic table was developed, but we have yet to talk about where the periodic table gets its name. To be periodic means to "have repeating cycles" or repeating patterns. In the periodic table, there are a number of physical properties that are "trend-like". This means is that as you move down a group or across a period, you will see the properties changing in a general direction.

The periodic table is a powerful tool that provides a way for chemists to organize the chemical elements. The word "periodic" means happening or recurring at regular intervals. The periodic law states that the properties of the elements recur periodically as their atomic numbers increase. The electron configurations of the atoms vary periodically with their atomic number. Because the physical and chemical properties of elements depend on their electron configurations, many of the physical and chemical properties of the elements tend to repeat in a pattern.

The actual repeating trends that are observed have to do with three factors. These factors are:

(1) The number of protons in the nucleus (called the nuclear charge).

(2) The number of energy levels holding electrons (and the number of electrons in the outer energy level).

(3) The number of electrons held between the nucleus and its outermost electrons (called the **shielding effect**). This affects the attraction between the valence electrons and the protons in the nucleus.

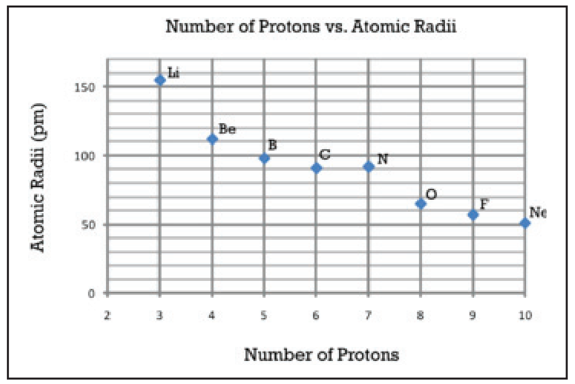
Trends in Atomic Radius

The atomic radius is a way of measuring the size of an atom. Although this is difficult to directly measure, we are, in essence, looking at the distance from the nucleus to the outermost electrons.

Let's look at the atomic radii or the size of the atom from the top of a family or group to the bottom. Take, for example, the Group 1 metals. Each atom in this family (and all other main group families) has the same number of electrons in the outer energy level as all the other atoms of that family. Each row (period) in the periodic table represents another added energy level. When we first learned about principal energy levels, we learned that each new energy level was larger than the one before. Energy level 2 is larger than energy level 1, energy level 3 is larger than energy level 2, and so on. Therefore, as we move down the periodic table from period to period, each successive period represents the addition of a larger energy level.

You can imagine that with the increase in the number of energy levels, the size of the atom must increase. The increase in the number of energy levels in the electron cloud takes up more space. Therefore the trend within a group or family on the periodic table is that *the atomic size increases with increased number of energy levels.*

In order to determine the trend for the periods, we need to look at the number of protons (nuclear charge), the number of energy levels, and the shielding effect. For a row in the periodic table, the atomic number still increases (as it did for the groups) and thus the number of protons would increase. When we examine the energy levels for period 2, we find that the outermost energy level does not change as we increase the number of electrons. In period 2, each additional electron goes into the second energy level. So the number of energy levels does not go up. As we move from left to right across a period, the number of electrons in the outer energy level increases but it is the same outer energy level.



Looking at the elements in period 2, the number of protons increases from lithium with three protons, to fluorine with nine protons. Therefore, the nuclear charge increases across a period. Meanwhile, the number of energy levels occupied by electrons remains the same. The numbers of electrons in the outermost energy level increases from left to right across a period, but how will this affect the radius? With an increase in nuclear charge, there is an increase in the pull between the protons and the outer level, pulling the outer electrons toward the nucleus. The net result is that *the atomic size decreases going across the row.*

Considering all the information about atomic size, you will recognize that the largest atom on the periodic table is all the way to the left and all the way to the bottom, francium, #87, and the smallest atom is all the way to the right and all the way to the top, helium, #2.

The fact that the atoms get larger as you move downward in a family is probably exactly what you expected before you even read this section, but the fact that the atoms get smaller as you move to the right across a period is most likely a big surprise. Make sure you understand this trend and the reasons for it.

Example: Which of the following has a greater radius?
(a) As or Sb
(b) Ca or K
(c) Polonium or Sulfur

Solution:
(a) Sb because it is below As in Group 15.
(b) K because it is further to the left on the periodic table.
(c) Polonium because it is below Sulfur in Group 16.

Periodic Trends in Ionization Energy

Lithium has an electron configuration of $1s^2\ 2s^1$. Lithium has one electron in its outermost energy level. In order to remove this electron, energy must be added. Look at the equation below:

$$Li(g) + \text{energy} \rightarrow Li^+(g) + e^-$$

With the addition of energy, a lithium ion can be formed from the lithium atom by losing one electron. This energy is known as the ionization energy. The **ionization energy** is the energy required to remove the most loosely held electron from a gaseous atom. The higher the value of the ionization energy, the harder it is to remove that electron.

Ionization Energies for some Group 1 Elements

Element	First Ionization Energy
Lithium, Li	520 kJ/mol
Sodium, Na	495.5 kJ/mol
Potassium, K	801 kJ/mol

Ionization Energies for Period 2 Elements

Element	Ionization Energy
Lithium, Li	520 kJ/mol
Beryllium, Be	899 kJ/mol
Boron, B	801 kJ/mol
Carbon, C	1086 kJ/mol
Nitrogen, N	1400 kJ/mol
Oxygen, O	1314 kJ/mol
Fluorine, F	1680 kJ/mol

Consider the ionization energies for the elements in group 1A of the periodic table, the alkali metals. Comparing the electron configurations of lithium to potassium, we know that the electron to be removed is further away from the nucleus, as the energy level of the valence electron increase. Because potassium's valence electron is further from the nucleus, there is less attraction between this electron and the protons and it requires less energy to remove this electron. As you move down a family (or group) on the periodic table, the ionization energy decreases.

We can see a trend when we look at the ionization energies for the elements in period 2. When we look closely at the data presented in the table above, we can see that as we move across the period from left to right, in general, the ionization energy increases. As we move across the period, the atoms become smaller which causes the nucleus to have greater attraction for the valence electrons. Therefore, as you move from left to right in a period on the periodic table, the ionization energy increases.

Example: Which of the following has a greater ionization energy?
(a) As or Sb
(b) Ca or K
(c) Polonium or Sulfur

Solution:
(a) As because it is above Sb in Group 15.
(b) Ca because it is further to the right on the periodic table.
(c) S because it is above Po in Group 16.

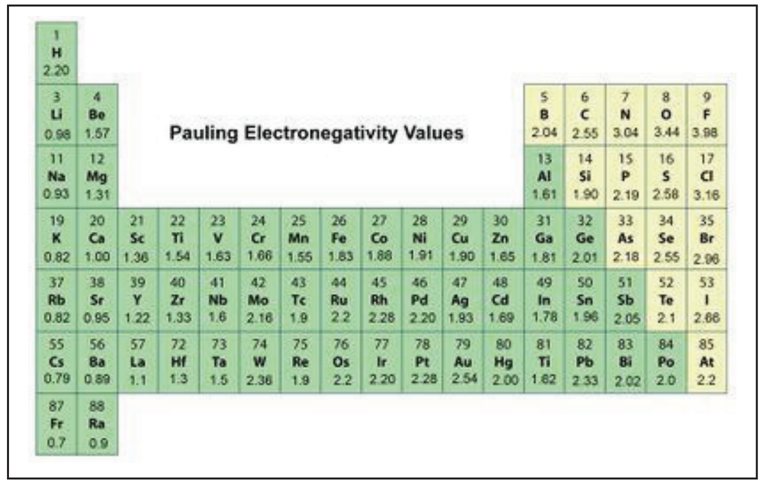
Periodic Trends in Electronegativity

Around 1935, the American chemist Linus Pauling developed a scale to describe the attraction an element has for electrons in a chemical bond. This is the **electronegativity**.

The values of electronegativity are higher for elements that more strongly attract electrons. On this Pauling scale fluorine, with an electronegativity of 4.0 is the most electronegative element, and cesium and francium, with electronegativities of 0.7, are the least electronegative.

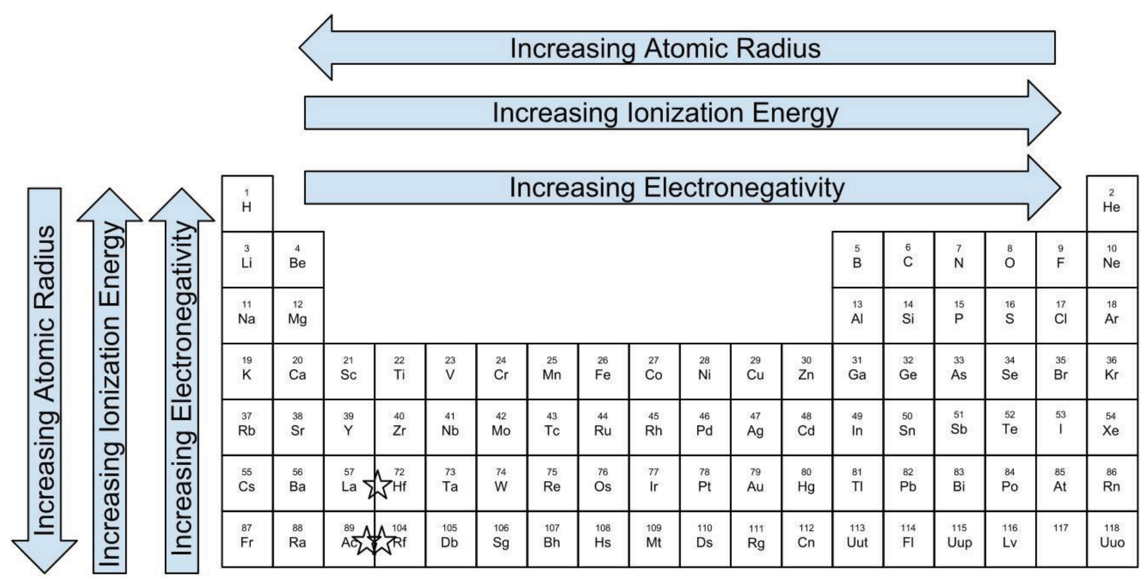
The electronegativity of atoms increases as you move from left to right across a period in the periodic table. This is because as you go from left to right across a period, the atoms of each element have the same number of energy levels. However, the nucleus charge increases, so the attraction that the atoms have for the valence electrons increases.

The electronegativity of atoms decreases as you move from top to bottom down a group in the periodic table. This is because as you go from top to bottom down a group, the atoms of each element have an increasing number of energy levels.



Atoms with low ionization energies have low electronegativities because their nuclei do not have a strong attraction for electrons. Atoms with high ionization energies have high electronegativities because the nucleus has a strong attraction for electrons.

Lesson Summary



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- Atomic size is the distance from the nucleus to the valence shell where the valence electrons are located.

- The atomic radius increases from the top to the bottom in any group and decreases from left to right across a period.
- Ionization energy is the energy required to remove the most loosely held electron from a gaseous atom or ion.
- Ionization energy generally increases across a period and decreases down a group.
- The higher the electronegativity of an atom, the greater its ability to attract shared electrons.
- The electronegativity of atoms increases as you move from left to right across a period in the periodic table and decreases as you move from top to bottom down a group in the periodic table.

Vocabulary

- Nuclear charge: the number of protons in the nucleus
- Shielding effect: the inner electrons help "shield" the outer electrons and the nucleus from each other.
- Ionization energy: the energy required to remove the most loosely held electron from a gaseous atom or ion
- Electronegativity: the ability of an atom in a molecule to attract shared electrons

3.5: Review Questions

Multiple choice

1) Why is the table of elements called "the periodic table"?
   a) it describes the periodic motion of celestial bodies.
   b) it describes the periodic recurrence of chemical properties.
   c) because the rows are called periods.
   d) because the elements are grouped as metals, metalloids, and non-metals.
   e) None of these.
2) Which of the following would have the largest ionization energy?
   a) Na
   b) Al
   c) H
   d) He
3) Which of the following would have the smallest ionization energy?
   a) K
   b) P
   c) S
   d) Ca

Short Answer

4) Which of the following would have a smaller radius: indium or gallium?
5) Which of the following would have a smaller radius: potassium or cesium?
6) Which of the following would have a smaller radius: titanium or polonium?
7) Explain why iodine is larger than bromine.
8) Arrange the following in order of increasing atomic radius: Tl, B, Ga, Al, In.
9) Arrange the following in order of increasing atomic radius; Ga, Sn, C.
10) Define ionization energy.

11) Place the following elements in order of increasing ionization energy: Na, S, Mg, Ar
12) Define electronegativity.

For each pair of elements, choose the element that has the lower electronegativity.
13) Li or N
14) Cl or Na
15) Na or K
16) Mg or F

Chapter 4: Describing Compounds

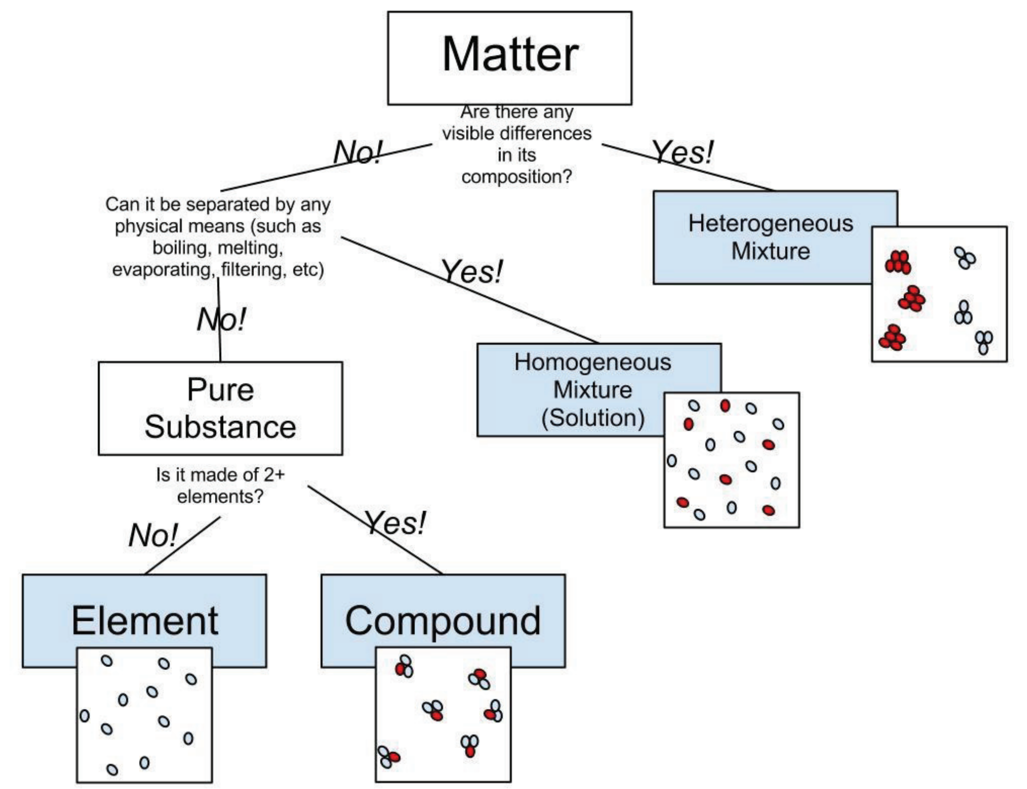
4.1: Introduction to Compounds

Objectives

- Explain the difference between an element, a compound and a mixture

Substances and Mixtures

Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning it's the same everywhere) and properties that are constant throughout the sample (meaning there is only one set of properties such as melting point, color, boiling point, etc throughout the matter). Elements and compounds are both example of pure substances.



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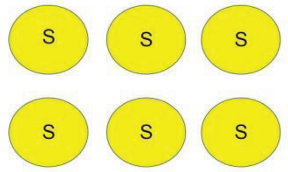
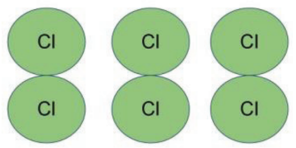
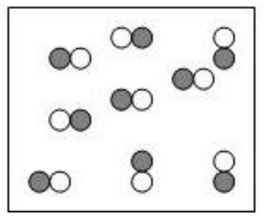
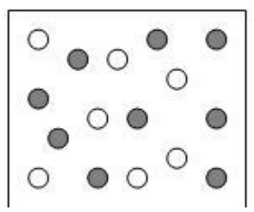
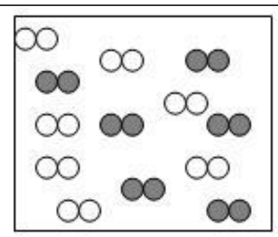
Mixtures are physical combinations of two or more elements and/or compounds. The term "physical combination" refers to mixing two different substances together where the substances do not chemically react. The physical appearance of the substances may change but the atoms and/or molecules in the substances do not change.

The chemical symbols are used not only to represent the elements; they are also used to write chemical formulas for the millions of compounds formed when elements chemically combine to form compounds. The **law of constant composition** states that the ratio by mass of the elements in a chemical compound is always the same, regardless of the source of the compound. The law of constant composition can be used to distinguish between compounds and mixtures. Compounds have a constant composition, and mixtures do not. For example, pure water is always 88.8% oxygen and 11.2% hydrogen by weight, regardless of the source of the water. Because water is a compound, it will always have this exact composition. Brass is an example of a mixture. Brass consists of two elements, copper and zinc, but it can contain as little 10% or as much as 45% zinc.

Consider the following examples including elements, compounds, and mixtures.

Pure substance (element)		*Matter with only one type of atom is called an element.*
Pure substance (element)		*Although the chlorine atoms are bonded in pairs, since there is only one type of atom, this is an element.*
Pure substance (compound)		*When two or more elements are bonded together, a compound is produced.*
Mixture		*When two or more pure substances (in this case, two elements) are combined, but not bonded together, a mixture is produced.*
Mixture		*When two or more pure substances are combined, but not bonded together, a mixture is produced.*
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The last couple of chapters have focused on elements and their properties. This unit will focus on compounds, including what compounds form and how elements combine to make compounds. Later chapters will cover mixtures.

Compounds and Chemical Formulas

The formula for a compound uses the symbols to indicate the type of atoms involved and uses subscripts to indicate the number of each atom in the formula. For example, aluminum combines with oxygen to form the compound aluminum oxide. To form aluminum

oxide requires two atoms of aluminum and three atoms of oxygen. Therefore, we write the formula for aluminum oxide as Al_2O_3. The symbol Al tells us that the compound contains aluminum, and the subscript 2 tells us that there are two atoms of aluminum in each molecule. The O tells us that the compound contains oxygen, and the subscript 3 tells us that there are three atoms of oxygen in each molecule. It was decided by chemists that when the subscript for an element is 1, no subscript would be used at all. Thus the chemical formula $MgCl_2$ tells us that one molecule of this substance contains one atom of magnesium and two atoms of chlorine. In formulas that contain parentheses, such as $Ca(OH)_2$, the subscript 2 applies to everything inside the parentheses. Therefore, this formula (calcium hydroxide) contains one atom of calcium and two atoms of oxygen and two atoms of hydrogen.

Lesson Summary

- Matter can be classified into two broad categories: pure substances and mixtures.
- A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample.
- Mixtures are physical combinations of two or more elements and/or compounds.
- Elements and compounds are both example of pure substances.
- Compounds are substances that are made up of more than one type of atom.
- Elements are the simplest substances made up of only one type of atom.

Vocabulary

- Element: a substance that is made up of only one type of atom
- Compound: a substance that is made up of more than one type of atom bonded together
- Mixture: a combination of two or more elements or compounds which have not reacted to bond together; each part in the mixture retains its own properties

Further Reading / Supplemental Links

- You may listen to Tom Lehrer's humorous song "The Elements" with animation at The Element Song (http://www.privatehand.com/flash/elements.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. Video on Demand – The World of Chemistry (http://www.learner.org/resources/series61.html?pop=yes&pid=793#)

4.1: Review Questions

Classify each of the following as an element, compound, or mixture.

1) Salt, NaCl
2) Oil
3) Gold
4) Sugar, $C_6H_{12}O_6$
5) Salad dressing
6) Salt water
7) Water
8) Copper
9) Air
10) Milk

4.2: Types of Compounds and Their Properties

Objectives:

- Distinguish between ionic, covalent, and metallic bonding in terms of electron behavior
- Given a formula, classify a compound as ionic, covalent, or metallic
- List properties of ionic, covalent, and metallic compounds

Introduction

Before students begin the study of chemistry, they usually think that the most stable form for an element is that of a neutral atom. As it happens, that particular idea is not true. If we were to consider the amount of sodium in the earth, we would find a rather large amount, approximately 190,000,000,000,000,000 kilotons. How much of this sodium would we find in the elemental form of sodium atoms? The answer is almost none. The only sodium metal that exists in the earth in the elemental form is that which has been man-made and is kept in chemistry labs and storerooms. Because sodium reacts readily with oxygen in the air and reacts explosively with water, it must be stored in chemistry storerooms under kerosene or mineral oil to keep it away from air and water. If those 1.9×10^{17} kilotons of sodium are not in the form of atoms, in what form are they? Virtually all the sodium in the earth is in the form of sodium ions, Na^+.

If all those tons of sodium ion can be found in nature and no sodium atoms can be found, it seems reasonable to suggest that, at least in the case of sodium, the ions are chemically more stable that the atoms. By *chemically stable*, we mean less likely to undergo chemical change. This is true not only for sodium but for many other elements as well.

The Octet Rule

Recall that the noble gas elements are the least reactive of all the elements on the periodic table – they almost never form any type of compound. Their electron configuration is the most stable of all of the elements, having their *s* and *p* sublevels filled. The noble gases have what is frequently referred to as an "octet", meaning they have eight valence electrons. The other elements are typically more stable if they have an octet, too. Other atoms will gain electrons, lose electrons, or share electrons in order to obtain an octet. The way in which an atom gets an octet determines the type of bond formed.

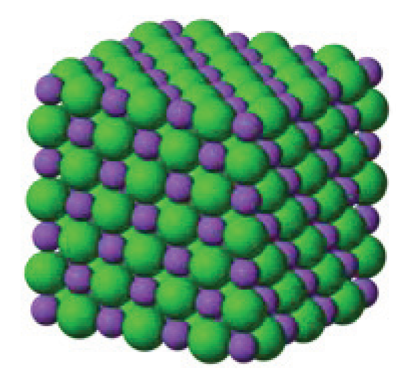
When an atom gains electrons, the atom will obtain a negative charge and is now called an **anion**. When an atom loses electrons, the atom will obtain a positive charge and is now called a **cation**. This may feel backwards, but remember that electrons themselves have a negative charge. When anions and cations are bonded together, the bond is said to be ionic. Metal atoms will lose electrons to obtain an octet and nonmetals will gain electrons. Therefore, in an ionic bond metals are typically bonded to nonmetals.

Some atoms are capable of obtaining an octet by sharing their valence electrons with another atom. This type of bonding is called a **covalent bond**. Only nonmetals are capable of forming covalent bonds with other nonmetals.

Properties of Ionic Compounds

When ionic compounds are formed, we are almost never dealing with just a single positive ion and a single negative ion. When ionic compounds are formed in laboratory conditions, many cations and anions are formed at the same time. The positive and negative

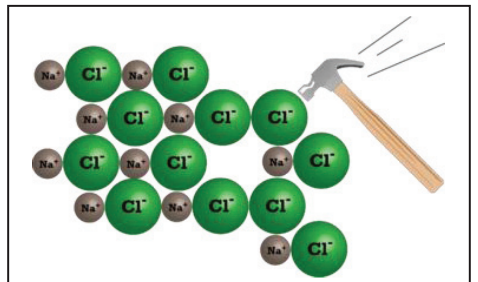
ions are not just attracted to a single oppositely charged ion. The ions are attracted to several of the oppositely charged ions. The ions arrange themselves into organized patterns where each ion is surrounded by several ions of the opposite charge. The organized patterns of positive and negative ions are called **lattice structures**. Because ionic compounds form these large lattice structures in the solid phase, they are not referred to as "molecules" but rather as lattice structures or crystals.



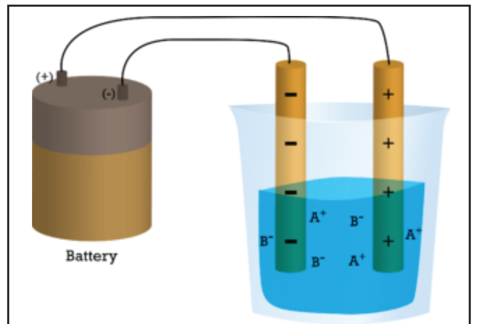
The three-dimensional crystal lattice structure of sodium chloride. (Source: http://en.wikipedia.org/wiki/Image:Sodium-chloride-3D-ionic.png. Public Domain)

The image shows the solid structure of sodium chloride. Each sodium ion is touching six chloride ions – the four surrounding ones and one above and one below. Each chloride ion is touching six sodium ions in the same way.

When electrons are transferred from metallic atoms to non-metallic atoms during the formation of an ionic bond, the electron transfer is permanent. That is, the electrons now belong to the non-metallic ion. This compound does not act like sodium and chlorine atoms did before they combined. This compound will act as sodium cations and chloride anions. If the compound is melted or dissolved, the particles come apart in the form of ions.



Ionic compounds are brittle, and break apart easily when struck with a hammer.



When dissolved in water, ionic compounds are able to conduct electricity.

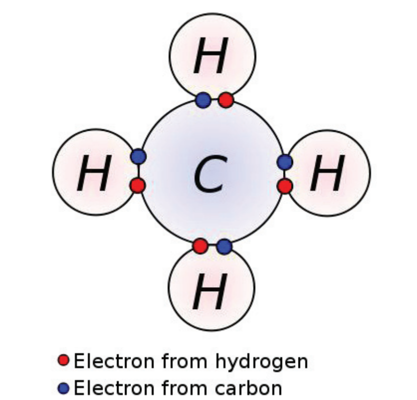
The electrostatic attraction between the oppositely charge ions is quite strong and therefore, ionic compounds have very high melting and boiling points. Sodium chloride (table salt), for example, must be heated to around 800°C to melt and around 1500°C to boil. There is only one type of solid that has higher melting points and boiling points, in general, than ionic compounds.

You can see that negative ions are surrounded by positive ions and vice versa. If part of the lattice is pushed downward, negative ions will then be next to negative ions and the structure will break up, therefore ionic compounds tend to be brittle solids. If you attempt to hammer on ionic substances, they will shatter. This is very different from metals which can be hammered into different shapes without the metal atoms separating from each other.

Ionic substances generally dissolve readily in water. In an ionic compound that has been melted or an ionic compound dissolved in water, ions are present that have the ability to move around in the liquid. The presence of the mobile ions in liquid or solution allow the solution to conduct electric current.

Properties of Covalent Compounds

In ionic compounds, we learned that atoms are able to achieve and octet through a metal giving away electrons (forming cations) and nonmetals taking electrons (forming anions). Some elements, however, can achieve an octet a different way, by sharing their valence electrons with other atoms instead. Typically, only nonmetals and sometimes metalloids are able to form covalent bonds. Metals, with their low numbers of valence electrons, are unable to achieve an octet through sharing valence electrons.



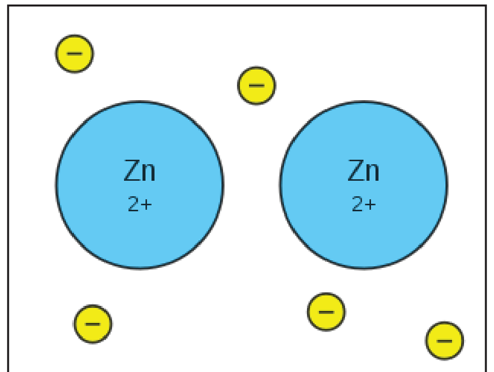
Carbon will share electrons with hydrogen, or other atoms, to get an octet.

The term *covalent bond* dates from 1939. The prefix *co-* means *jointly* (as in, coworker, cooperate, etc), etc.; " *valent"* is referring to an atom's valence electrons. Thus, a "co-valent bond", essentially, means that the atoms share valence electrons.

Covalent compounds have properties very different from ionic compounds. Ionic compounds have high melting points causing them to be solid at room temperature, and conduct electricity when dissolved in water. Covalent compounds have low melting points and many are liquids or gases at room temperature. Whereas most ionic compounds are capable of dissolving in water, many covalent compounds do not. Also unlike ionic compounds, when covalent compounds are dissolved in water, they are not conductors of electricity.

Properties of Metallic Bonds

There is a third type of bond that may be formed between two atoms. In **metallic bonding,** the electrons between neighboring metal atoms are delocalized, meaning that the electrons are not tied to one atoms specifically. The electrons, instead, are gathered in what we call an "electron sea". In an electron sea, the metal nuclei form the basis, and the electrons move around the nuclei.



Metallic Bonding has delocalized electrons, in which the electrons are free to move throughout the metal.

Because of this unique type of bonding structure, metallic bonding accounts for many physical properties of metals, such as strength, malleability (or bendability), ductility, conductivity (allows heat and electricity to go through), and luster (shine).

Properties of Compounds vs. the Elements of which they are composed

It is important to point out what when elements combine to form covalent compounds or ionic compounds the properties of the compound are different from the properties of the elements from which the compound is formed. Consider, for example, sugar, formed from the elements carbon, hydrogen, and oxygen: $C_{12}H_{22}O_{11}$. Forms of carbon you are probably familiar with include coal and graphite (pencil lead). Oxygen is a gas necessary for your

survival, and hydrogen is also a very flammable gas. You wouldn't want the sugar you put on your cereal to taste like coal or be as flammable as hydrogen and oxygen gases. When combined, a new compound is made with its own unique properties, different from the elements that formed the compound.

The process of gaining and/or losing electrons completely changes the chemical properties of the substances. The chemical and physical properties of an ionic compound will bear no resemblance to the properties of the elements which formed the ions. For example, sodium is a metal that is shiny, an excellent conductor of electric current, and reacts violently with water. Chlorine is a poisonous gas. When sodium and chlorine are chemically combined to form sodium chloride (table salt), the product has an entirely new set of properties. Sometimes, we sprinkle sodium chloride on our food. This is not something we would do if we expected it to explode when contacted by water or if we expected it to poison us.

What happens when these elements combine in different ratios, forming compounds such as isopropyl alcohol (commonly called rubbing alcohol), C_3H_7OH, or acetone (the main ingredient in most finger nail polish removers), C_3H_6O? Does rubbing alcohol have the same properties as finger nail polish remover or sugar? No! When elements combine in different ratios, different compounds are formed which have their own unique properties. Each compound will typically have its own melting point, boiling point, and density. Frequently, they will have a unique smell or taste. They will also have unique chemical properties and react differently from other compounds.

Summary:

- The octet rule is an expression of the tendency for atoms to gain or lose the appropriate number of electrons so that the resulting ion has either completely filled or completely empty outer energy levels.
- Ionic compounds form ionic crystal lattices rather than molecules, have very high melting and boiling points, and tend to be brittle solids. They are generally soluble in water and their water solutions will conduct electricity.
- Covalent compounds are formed from nonmetals sharing electrons. They tend to have low melting and boiling points. Although some are soluble in water, they do not conduct electricity when dissolved.
- Metallic bonds allow the electrons to move freely, resulting in materials that are very conductive, malleable, and lustrous.
- Compounds have chemical properties that are unrelated to the chemical properties of the elements from which they were formed.

Vocabulary:

- Octet Rule: The tendency of an atom to be more stable with eight valence electrons
- Ionic compound: a positively charged particle (typically a metal) bonded to a negatively charged particle (typically a nonmetal) held together by electrostatic attraction
- Covalent compound: two or more atoms (typically nonmetals) forming a molecule in which electrons are being shared between atoms.

Further Reading / Supplemental Links

- To see a video and clips discussing the types of bonds, go to http://www.uen.org/dms/. Go to the k-12 library. Search for covalent bond, metallic bond, or ionic bond. (you can get the username and password from your teacher)

4.2: Review Questions

1) What does the octet rule state?
2) Which elements are able to form covalent bonds in order to get an octet? Which are not?

Given the following chemical formulas, label each compound as ionic, metallic, or covalent.

3) H_2O 4) MgO 5) NO 6) Li_3PO_4 7) Rb_3N

8) CCl_4 9) $Ni_3(PO_4)_2$ 10) Cu Zn 11) CH_4 12) NH_3

Label each of the following properties as a property of an ionic, covalent, or metallic compound:

13) Low melting point
14) Conducts electricity in solid state
15) Conducts electricity when dissolved in water
16) Brittle crystal structure
17) Composed only nonmetallic atoms

18) Steve is given three substances in the lab to identify. He performs a conductivity test, a solubility test (determines whether the compound dissolve in water), and determines the melting point using a melting point apparatus. Some of the melting points, the teacher tells him are too high or low to measure using the laboratory melting point apparatus so she gives him the melting point. Help Steve match the properties of the unknowns (from the table below) to the substance names.

List of Unknown names:
Sodium chloride, NaCl
Zinc, Zn
Sucrose (table sugar, $C_{12}H_{22}O_{11}$)

Unknown Substance	Conducts when dissolved	Malleable	Conducts as a solid	Dissolves in water	Melting Point (°C)
1	No	No	No	Yes	164°C
2	Yes	No	No	Yes	~800°C
3	N/A	Yes	Yes	No	420°C

4.3: Names and Charges of Ions

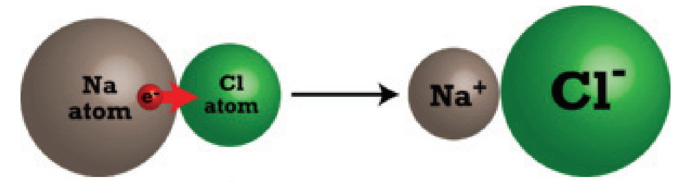
Objectives

- Describe how atoms form an ionic bond.
- Indicate the most likely number of electrons the atom will gain or lose when forming an ion
- Describe what polyatomic ions are.

Introduction

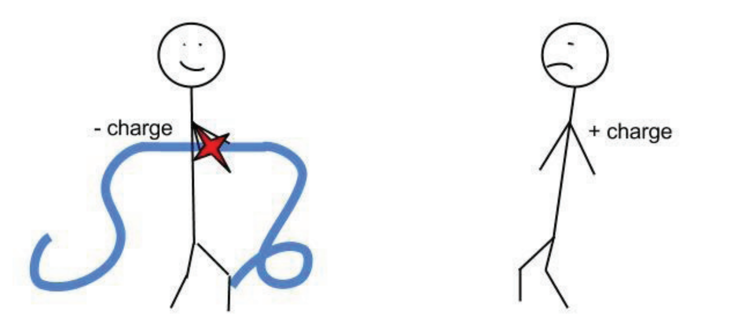
Molecular collisions between atoms that tend to lose electrons (metals) and atoms that tend to gain electrons (non-metals) are sometime sufficient to remove electrons from the metal atom and add them to the non-metal atom. This transference of electrons from metals to non-metals forms positive and negative ions, which in turn, attract each other due to opposite charges. The compounds formed by this electrostatic attraction are said to be ionically bonded.

Ionic Bonding



When an ionic bond forms, electrons are lost from a metal and given to a nonmetal.

The process of transferring an electron from a sodium atom to a chlorine atom as shown the diagram produces oppositely charged ions which then stick together because of **electrostatic attraction**. Electrostatic attraction is the attraction between opposite charges. The electrostatic attraction between oppositely charged ions is called an **ionic bond.** These ions are chemically more stable than the atoms were.



An ionic bond is similar to a tug-of-war in which the stronger (more electronegative) nonmetal atom pulls the electron away from the weaker (less electronegative) metal atom.
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If we had been using sodium and sulfur atoms for the transfer discussion, the process would be only slightly different. Sodium atoms have a single electron in their outermost energy level and therefore can lose only one electron. Sulfur atoms, however, require two electrons to complete their outer energy level. In such a case, two sodium atoms would be required to collide with one sulfur atom. Each sodium atom would contribute one electron for a total of two electrons and the sulfur atom would take on both electrons. The two Na atoms would become Na^+ ions and the sulfur atom would become a S^{2-} ion. Electrostatic attractions would cause all three ions to stick together.

The cation forming elements, metals, lose all valence electrons so the electron configuration for the ions formed will have the eight electrons of the previous noble gas. (Those whose electron dot formula matches helium, of course, will have only two.) The anion forming elements, nonmetals, will gain enough electrons so the electron dot formulas of their ions will match those of the following noble gas. In all cases, for the "A" groups elements, the ions will have eight electrons in their electron dot formula. The **octet rule** is an expression of this end result of eight electrons in the outer most energy level.

An atom becomes an ion when it gains or loses electrons. The ions that are formed when an atom loses electrons are positively charged because they have more protons in the nucleus than electrons in the electron cloud. Positively charged ions are called **cations** (pronounced CAT-ions). The ions that are formed when an atom gains electrons are negatively charged because they have more electrons in the electron cloud than protons in the nucleus. Negatively charged ions are called **anions** (pronounced AN-ions).

Predicting Charges of Main Group Ions

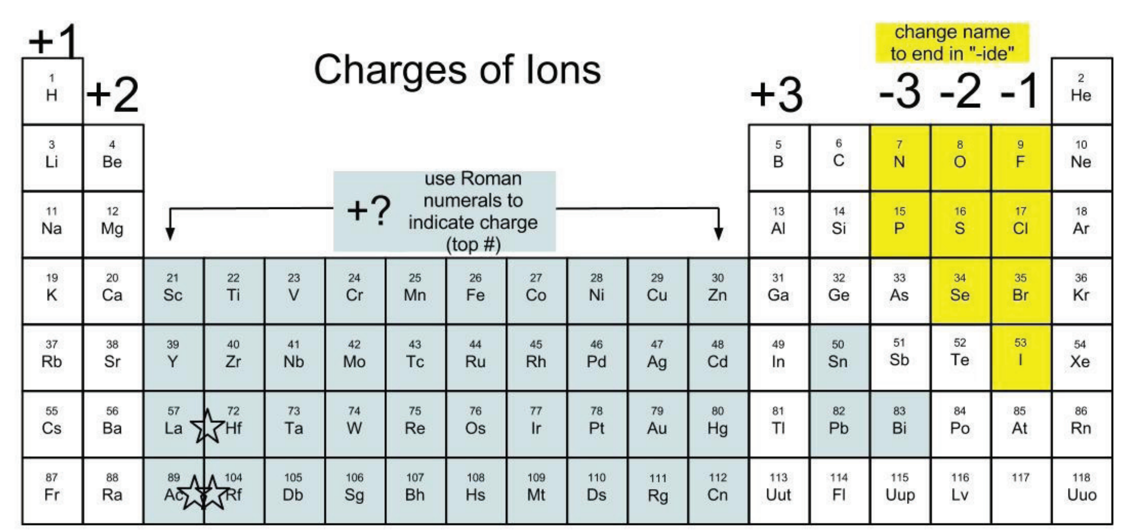
All the metals in family 1A have electron configurations ending with an s^1 electron in the outer energy level. For that reason, all family 1A members will tend to lose exactly one electron when they are ionized, obtaining an electron configuration like the closest noble gas. The entire family forms +1ions: Li^+, Na^+, K^+, etc. We need to note that while hydrogen is in this same column, it is not considered to be an metal. There are times that hydrogen acts as if it is a metal and forms +1ions; however, most of the time it bonds with other atoms as a nonmetal. In other words, hydrogen doesn't easily fit into any chemical family. All members of family 1A form ions with +1charge.

The metals in family 2A all have electron configurations ending with two electrons in an s^2 position in the outermost energy level. To have an electron configuration like the closest noble gas, each of the elements in this family will lost two valence electrons and form +2 ions; Be^{2+}, Mg^{2+}, etc. Other metal elements' charges can be predicted using the same patterns. Members of family 3A form ions with 3+ charge.

Family 5A nonmetals at the top will gain electrons to form negative ions. By gaining electrons, they are able to obtain the electron configuration of the noble gas closest to them on the periodic table. Family 6A non-metals will gain two electrons to obtain a octet thus forming a -2 ion. Family 7A will form -1 ions: F^-, Cl^-, etc.

Family 8A, of course, is the noble gases and has no tendency to either gain or lose electrons so they do not form ions.

The charges ions form can be summarized as in the following table. Many of the transition elements have variable oxidation states so they can form ions with different charges, and, therefore, are left off of this chart.



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When main group nonmetals gain electrons to form anions, their names are changed to end in "-ide". For example, fluorine atoms gain electrons to become fluoride ions.

Polyatomic Ions

Thus far, we have been dealing with ions made from single atoms. Such ions are called monatomic ions. There also exists a group of **polyatomic ions**, ions composed of a group of atoms that are covalently bonded and behave as if they were a single ion. Almost all the common polyatomic ions are negative ions.

A table of many common polyatomic ions is given. The more familiar you become with polyatomic ions, the better you will be able to write names and formulas of ionic compounds. It is also important to note that there are many polyatomic ions that are not on this chart.

Cations

+1
Ammonium, NH_4^+

Anions

-1	-2	-3
Hypochlorite, ClO^- Chlorite, ClO_2^- Chlorate, ClO_3^- Perchlorate, ClO_4^-	Sulfite, SO_3^{2-} Sulfate, SO_4^{2-}	Phosphate, PO_4^{3-}
Nitrite, NO_2^- Nitrate, NO_3^- Bicarbonate, HCO_3^-	Carbonate, CO_3^{2-}	
Hydroxide, OH^-	Peroxide, O_2^{2-}	
Acetate, $C_2H_3O_2^-$	Oxalate, $C_2O_4^{2-}$ Silicate, SiO_3^{2-} Thiosulfate, $S_2O_3^{2-}$	
Permanganate, MnO_4^-	Chromate, CrO_4^{2-} Dichromate, $Cr_2O_7^{2-}$	
Cyanide, CN^- Thiocyanate, SCN^-		

Naming Transition Metals

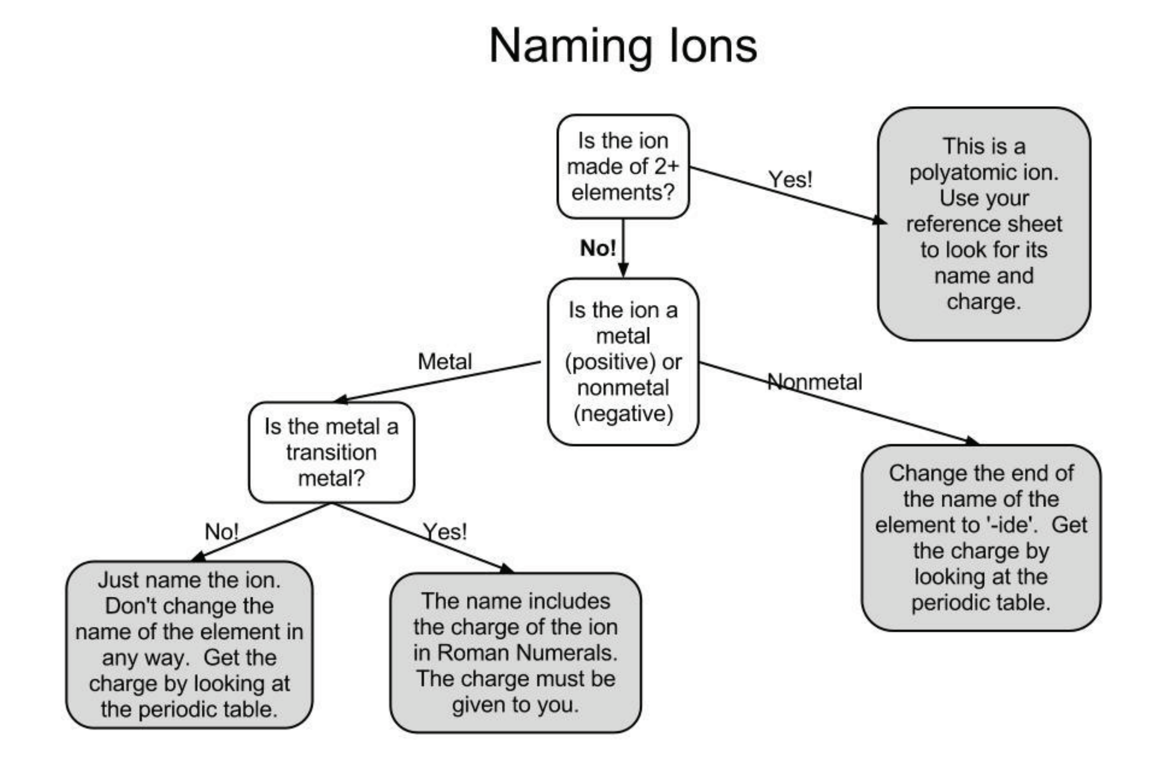
Some metals are capable of forming ions with various charges. These include most of the transition metals and many post transition metals. Iron, for example, may form Fe^{2+} ions by losing 2 electrons or Fe^{3+} ions by losing 3 electrons. The rule for naming these ions is to insert the charge (oxidation number) of the ion with Roman numerals in parentheses after the name. These two ions would be named iron (II) and iron (III). When you see that the compound involves any of the variable oxidation number metals (iron, copper, tin, lead, nickel, and gold), you must determine the charge (oxidation number) of the metal from the formula and insert Roman numerals indicating that charge.

Consider FeO and 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 is exhibiting a charge of 2+ and in

Fe_2O_3, it is exhibiting a charge of 3+. The first, FeO, is named iron (II) oxide. The second, Fe_2O_3, is named iron (III) oxide.

Lesson Summary

- When an atom gains one or more extra electrons, it becomes a negative ion, an anion
- When an atom loses one or more of its electrons, it becomes a positive ion, a cation.
- Polyatomic ions are ions composed of a group of atoms that are covalently bonded and behave as if they were a single ion.
- Some transition elements have fixed oxidation numbers and some have variable charges. When naming these charge variable ions, their charges are included in Roman numerals.



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Vocabulary

- Ion: An atom or group of atoms with an excess positive or negative charge.
- Cation: positive ion
- Anion: negative ion
- ionic bond: A bond between ions resulting from the transfer of electrons from one of the bonding atoms to the other and the resulting electrostatic attraction between the ions.
- electrostatic attraction: The force of attraction between opposite electric charges.

- octet rule: the tendency for atoms gain or lose the appropriate number of electrons so that the resulting ion has either completely filled or completely empty outer energy levels, or 8 valance electrons.

Further Reading / Supplemental Links

- Website with lessons, worksheets, and quizzes on various high school chemistry topics. Lesson 4-1 is on Electronegativity. Lesson 4-2 is on Types of Bonds.
- http://www.fordhamprep.org/gcurran/sho/sho/lessons/lesson31.htm

4.3: Review Questions

1) Define an ion.
2) Will an iron atom form a positive or negative ion? Why?
3) Will a bromine atom form a positive or negative ion? Why?

Predict the charge of each ion. Then give the name each ion would have.

4) Cl	5) Br	6) N
7) O	8) Ca	9) F
10) Mg	11) Li	12) I
13) Na	14) K	15) Al

16) How are transition metals that form ions named differently than other metals? Why is this important? What does the Roman numeral tell you?

Name the following ions.

17) Cu^{2+}	18) Co^{2+}	19) Co^{3+}
20) Cu^{+}	21) Ni^{2+}	22) Cr^{3+}
23) Fe^{2+}	24) Fe^{3+}	

25) What are polyatomic ions?

Name each of the following ions.

26) NO_3^-	27) $C_2H_3O_2^-$	28) OH^-
29) PO_4^{3-}	30) SO_3^{2-}	31) CO_3^{2-}

4.4: Writing Ionic Formulas

Objectives

- Write the correct formula for an ionic compound

Introduction

Ionic compounds do not exist as molecules. In the solid state, ionic compounds are in crystal lattices containing many ions each of the cation and anion. An ionic formula, like NaCl, is an empirical formula. This formula merely indicates that sodium chloride is made of an equal number of sodium and chloride ions. Sodium sulfide, another ionic compound, has the formula Na_2S. This formula indicates that this compound is made up of twice as many sodium ions as sulfide ions. This section will teach you how to find the correct ratio of ions, so that you can write a correct formula.

Ionic Formulas

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 combine with fluorine, which takes on one electron, then one calcium atom must combine with two fluorine atoms. The formula would be CaF_2.

To write the formula for an ionic compound:

1) Write the symbol and charge of the cation (first word)
   a) If the element is in group 1, 2, Al with a consistent charge, you can get the charge using your periodic table.
   b) If the metal is a transition metal with a variable charge, the charge will be given to you in Roman numerals.
2) Write the symbol and charge of the anion (second word).
   a) Look at your polyatomic ion chart first. If your anion is a polyatomic ion, write the ion in parentheses.
   b) If the anion is not on the polyatomic chart, it is a nonmetal anion from your periodic table. You can get its charge using your table.
3) Write the correct subscripts so that the total charge of the compound will be zero.
4) Write the final formula. Leave out all charges and all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.

Pay close attention to how these steps are followed in the given examples.

Example: Write the formula for aluminum chloride.

Solution:
Cation, aluminum: Al^{3+} (you can find this charge using your periodic table)
Anion, chloride: Cl^- (chloride is chlorine as an ion, get its charge from your periodic table)
To balance the charges you need 1·(+3) and 3·(-1). Giving: $Al_1^{3+}Cl_3^{-1}$
The final formula is: $AlCl_3$

Example: Write the formula for aluminum sulfide.

Solution:
Cation, aluminum: Al^{3+} (you can find this charge using your periodic table)
Anion, sulfide: S^{2-} (sulfide is sulfur as an ion, get its charge from your periodic table)
To balance the charges you need 2·(+3) and 3·(-2). Giving: $Al_2^{3+}S_3^{-2}$
The final formula is: Al_2S_3

Example: Write the formula for lead (IV) oxide.

Solution:
Cation, lead (IV): Pb^{4+} (the charge is given to you as Roman numerals, because this is a metal with a variable charge)
Anion, oxide: O^{2-} (oxide is oxygen as an ion, get its charge from your periodic table)
To balance the charges you need 1·(+4) and 2·(-2). Giving: $Pb_1^{4+}O_2^{-2}$
The final formula is: PbO_2

Example: Write the formula for calcium nitrate.

Solution:
Cation, calcium: Ca^{2+} (you can find this charge using your periodic table)
Anion, nitrate: $(NO_3)^-$ (this is a polyatomic ion)
To balance the charges you need 1·(+2) and 2·(-1). Giving: $Ca_1^{2+}(NO_3)_2^{1-}$
The final formula is: $Ca(NO_3)_2$
In this case you need to keep the parentheses. There are two of the group $(NO_3)^-$. Without the parentheses, you are merely changing the number of oxygen atoms.

Example: Write the formula for magnesium sulfate.

Solution:
Cation, magnesium: Mg^{2+} (you can find this charge using your periodic table)
Anion, sulfate: $(SO_4)^{2-}$ (this is a polyatomic ion)
To balance the charges you need 1·(+2) and 1·(-2). Giving: $Mg_1^{2+}(SO_4)_1^{2-}$
The final formula is: $MgSO_4$
In this case you do not need parentheses. They are only required if there is more than one of the polyatomic ion.

Example: Write the formula for copper (II) acetate.

Solution:
Cation, copper (II): Cu^{2+} (the charge is given to you in Roman numerals)
Anion, acetate: $(C_2H_3O_2)^-$ (this is a polyatomic ion)
To balance the charges you need 1·(+2) and 2·(-1). Giving: $Cu_1^{2+}(C_2H_3O_2)_2^{1-}$
The final formula is: $Cu(C_2H_3O_2)_2$
In this case you need to keep the parentheses. There are two of the group $(C_2H_3O_2)^-$. Without the parentheses, you are merely changing the number of oxygen atoms.

Lesson Summary

- Formulas for ionic compounds contain the lowest whole number ratio of subscripts such that the sum of the subscript of the more electropositive element times its oxidation number plus the subscripts of the more electronegative element times its oxidation number equals zero.

Vocabulary

- Ionic Formula: includes the symbols and number of each atom present in a compound in the lowest whole number ratio

Further Reading / Supplemental Links

- http://www.kanescience.com/_chemistry/5Ionic.htm
- http://visionlearning/library/module_viewer.php?mid=55

4.4: Review Questions

Copy and fill in the chart by writing formulas for the compounds that might form between the ions in the columns and rows. Some of these compounds don't exist but you can still write formulas for them.

	Na^+	Ca^{2+}	Fe^{3+}
NO_3^-	1)	2)	3)
SO_4^{2-}	4)	5)	6)
Cl^-	7)	8)	9)
PO_4^{3-}	10)	11)	12)
OH^-	13)	14)	15)
CO_3^{2-}	16)	17)	18)

Write the formulas from the names of the following compounds.

19) Magnesium sulfide
20) Lead(II) Nitrate
21) Sodium Oxide
22) Calcium hydroxide
23) Potassium Carbonate
24) Aluminum Bromide
25) Iron (III) nitrate
26) Iron(II) Chloride
27) Copper(II) Nitrate
28) Magnesium oxide
29) Calcium Oxide
30) Copper(I) Bromide
31) Aluminum sulfide
32) Hydrogen Carbonate
33) Potassium permanganate
34) Copper (I) dichromate
35) Iron(III) Chloride
36) Iron(II) Sulfate

4.5: Naming Ionic Compounds

Objectives

- Correctly name binary ionic compounds, compounds containing metals with variable oxidation numbers, and compounds containing polyatomic ions given the formulas.

Introduction:

We have already learned about naming individual ions, including main group ions, transition metal ions, and polyatomic ions. We have also learned how to put these into correct charge-balanced formulas. In this section, we will learn how to correctly naming a compound, given its formula.

Naming Ionic Compounds

To name ionic compounds, we will need to follow these steps:

1. Split the formula into the cation and anion. The first metal listed will be the cation and the remaining element(s) will form the anion.
2. Name the cation. We learned two types of cations:
   a. Main group cations in which the name of the ion is the same as that of the element (for example, K^+ is potassium).
   b. Transition metals with variable charges with Roman numerals indicating the charge of the ion (you will have to do a little bit of math to find this charge).
3. Name the anion. There are also two general types of anions:
   a. Main group anions in which the name of the anion ends in "-ide" (for example, F^- is fluoride)
   b. Polyatomic ions (as listed on the polyatomic ion chart)

When writing the name of an ionic compound, it is important to note that the name gives no information about the number of ions. The name only tells the types of ions present. The formula uses subscripts to indicate how many of each ion there are.

Example: What is the name of Na_2O?

Solution:
Split up the formula: Na_2 | O
Name the cation: Na is a group 1 metal with a consistent charge. It does not need Roman numerals. Its name is "sodium"
Name the anion: O is not polyatomic. When oxygen atoms get a -2 charge, the name changes to end in –ide, so the anion is "oxide"
Final answer: sodium oxide

Example: What is the name of $NaC_2H_3O_2$?

Solution:
Split up the formula: Na | $C_2H_3O_2$
Name the cation: Na is a group 1 metal with a consistent charge. It does not need Roman numerals. Its name is "sodium"
Name the anion: $C_2H_3O_2$ is polyatomic. Its name is "acetate".
Final answer: sodium acetate

Example: Write the name of $CuCl_2$?

Solution:
Split up the formula: Cu | Cl_2
Name the cation: Cu is a transition metal with a variable charge. It needs Roman numerals.
To find the charge, consider the charge of the other ion and the number of both ions:

$Cu_1^?Cl_2^{-1}$. The copper must have a charge of +2 to balance out the negatives: 1·(+2) to cancel out 2·(-1). Its name is "copper (II)"
Name the anion: Cl is not polyatomic. When chlorine atoms get a -1 charge, the name changes to end in –ide, so the anion is "chloride"
Final answer: copper (II) chloride

Example: Write the name of PbS_2?

Solution:
Split up the formula: Pb | S_2
Name the cation: Pb is a post-transition metal with a variable charge. It needs Roman numerals. To find the charge, consider the charge of the other ion and the number of both ions: $Pb_1^?S_2^{-2}$. The copper must have a charge of +4 to balance out the negatives: 1·(+4) to cancel out 2·(-2). Its name is "lead (IV)"
Name the anion: S is not polyatomic. When sulfur atoms get a -2 charge, the name changes to end in –ide, so the anion is "sulfide"
Final answer: Lead (IV) sulfide.

The most common error made by students in naming these compounds is to choose the Roman numeral based on the number of atoms of the metal instead of the charge of the metal. For example, in PbS_2, the oxidation state of lead Pb is +4 so the Roman numeral following the name lead is "IV." Notice that there is no four in the formula. As in previous examples, the formula is always the lowest whole number ratio of the ions involved. Think carefully when you encounter variable charge metals. Make note that the Roman numeral does not appear in the formula but does appear in the name.

Example: Write the name of $Mg_3(PO_4)_2$?

Solution:
Split up the formula: Mg_3 | $(PO_4)_2$
Name the cation: Mg is a group 2 metal with a consistent charge. It does not need Roman numerals. Its name is "magnesium"
Name the anion: PO_4 is polyatomic. Its name is "phosphate".
Final answer: sodium acetate

Example: Write the name of $Cr(NO_2)_3$?

Solution:
Split up the formula: Cr | $(NO_2)_3$
Name the cation: Cr is a transition metal with a variable charge. It needs Roman numerals. To find the charge, consider the charge of the other ion and the number of both ions: $Cr_1^?(NO_2)_3^{-1}$. The copper must have a charge of +3 to balance out the negatives: 1·(+3) to cancel out 3·(-1). Its name is "chromium (III)"
Name the anion: NO_2^- is polyatomic. Its name is "nitrite".
Final answer: chromium (III) nitrite

Remember, when writing the name of an ionic compound, it is important to note that the name gives no information about the number of ions. The name only tells the types of ions present. The formula uses subscripts to indicate how many of each ion there are.

Lesson Summary

- Ionic bonds are formed by transferring electrons from metals to non-metals after which the oppositely charged ions are attracted to each other.
- Ionic compounds form crystal lattice structures rather than molecules.
- Binary ionic compounds are named by naming the metal first followed by the non-metal with the ending of the non-metal changed to "ide."
- Compounds containing polyatomic ions are named with the name of the polyatomic ion in the place of the metal or non-metal or both with no changes in the name of the polyatomic ion.
- Compounds containing variable oxidation number metals are named with Roman numerals in parentheses following the name of the metal and indicating the oxidation number of the metal.

Vocabulary

- Anion: An ion with a negative charge.
- Cation: An ion with a positive charge.
- Chemical nomenclature: The system for naming chemical compounds.
- Ionic bond: The electrostatic attraction between ions of opposite charge.
- Polyatomic ion: A group of atoms bonded to each other covalently but possessing an overall charge.

Further Reading / Supplemental Links

- Matching Game: Naming Ionic Compounds with Polyatomic Ions: http://www.quia.com/mc/65767.html

4.5: Review Questions

Name the following compounds.

1) KCl	2) MgO	3) $CuSO_4$
4) NaCl	5) $CoBr_2$	6) MgF_2
7) $Ni(OH)_2$	8) $NaC_2H_3O_2$	9) CuO
10) $FeCl_2$	11) LiCl	12) $MgBr_2$
13) $NH_4(OH)$	14) Cu_2O	15) CaF_2
16) K_2CO_3	17) Na_2O	18) PbO
19) $Ca(NO_3)_2$	20) $Mg(OH)_2$	21) SnO_2

4.6: Covalent Compounds & Lewis Structures

Objectives

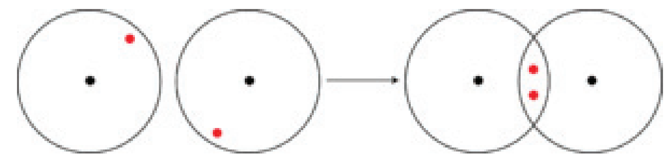
- Explain what covalent bonds are.
- Explain why covalent bonds are formed.
- Draw a Lewis structures for covalent compounds and polyatomic ions

Introduction

In ionic bonding, electrons leave metallic atoms and enter non-metallic atoms. This complete transfer of electrons changes both of the atoms into ions. Often, however, two atoms combine in a way that no complete transfer of electrons occurs. Instead, electrons are held in overlapping orbitals of the two atoms, so that the atoms are sharing the electrons. The shared electrons occupy the valence orbitals of both atoms at the same time. The nuclei of both atoms are attracted to this shared pair of electrons and the atoms are held together by this attractive force. The attractive force produced by sharing electrons is called a **covalent bond.**

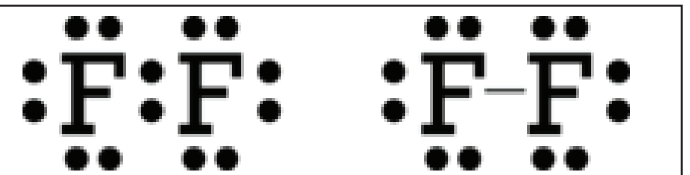
Covalent Bond Formation

In covalent bonding, the atoms acquire a stable octet of electrons by sharing electrons. The covalent bonding process produces molecular substances as opposed to the lattice structures of ionic bonding. There are far more covalently bonded substances than ionic substances.



In a covalent bond, electrons are shared by atoms in overlapping orbitals.

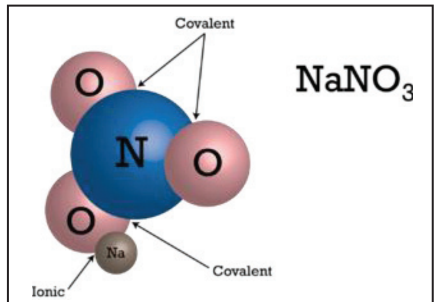
The diatomic hydrogen molecule, H_2, is one of the many molecules that are covalently bonded. Each hydrogen atom has a *1s* electron cloud containing one electron. These *1s* electron clouds overlap and produce a common volume which the two electrons occupy.



We can also show a covalent bond between atoms with an electron dot formula where the shared pair of electrons are the bonding electrons or with the bond represented by a dash.

Some Compounds Have Both Covalent and Ionic Bonds

If you recall the introduction of polyatomic ions, you will remember that the bonds that hold the polyatomic ions together are covalent bonds. Once the polyatomic ion is constructed with covalent bonds, it reacts with other substances as an ion. The bond between a polyatomic ion and another ion will be ionic. An example of this type of situation is in the compound sodium nitrate. Sodium nitrate is composed of a sodium ion and a nitrate ion. The nitrate ion is held together by covalent bonds and the nitrate ion is attached to the sodium ion by an ionic bond.



In this compound, the N and O atoms are covalently bonded, sharing electrons. However, the NO_3 anion is ionically bonded to the Na cation.

Lewis Structures

The **Lewis structure** of a molecule show how the valence electrons are arranged among the atoms of the molecule. These representations are named after G. N. Lewis. The rules for writing Lewis structures are based on observations of thousands of molecules. From experiment, chemists have learned that when a stable compound forms, the atoms usually have a noble gas electron configuration or eight valence electrons. Hydrogen forms stable molecules when it shares two electrons (sometimes called the duet rule). Other atoms involved in covalent bonding typically obey the octet rule. (Note: Of course, there will be exceptions.)

To draw a Lewis structure:

1. Determine the number of valence electrons that will be drawn in the Lewis structure.
   a. Use your periodic table to determine the number of valence electrons in each atom. Add these to get the total electrons in the structure.
   b. If you are drawing the structure for a polyatomic ion, you must add or subtract any electrons gained or lost. If an ion has a negative charge, electrons were gained. If the ion has a positive charge, electrons were lost.
2. Draw a skeleton
   a. Typically, the first element listed in the formula goes in the center, which the remaining atoms surrounding.
   b. Draw bonds to each of the surrounding atoms. Each bond is two valence electrons.
3. Use the remaining electrons to give each atom an octet (except hydrogen which only gets a duet)
   a. Place electrons left over after forming the bonds in the skeleton in unshared pairs around the atoms to give each an octet. *Remember, any bonds they have formed already count as two valence electrons each.
   b. If you run out of electrons, and there are still atoms without an octet, move some of the electrons that are not being shared to form double, sometimes triple bonds.

Example: Draw a Lewis structure for water, H_2O.

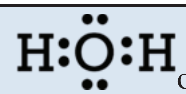
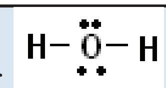
Solution:

1) add up all available valence electrons: each H atom has 1, each oxygen atom has 6, so 2(1)+6=8

2) Draw a skeleton. Although the first atom written typically goes in the middle, hydrogen can't, so O gets the middle spot. We need to draw bonds connecting atoms in the skeleton. We get:

H– O– H

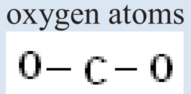
3) Use the remaining electrons to give each atom (except hydrogen) an octet. If we look at our skeleton, we drew two bonds, which uses 4 of our 8 available electrons. We are left with four more. Each H atom already has two valence electrons and O currently has 4 (each bond counts as two for each atom that it connects). We will give the remaining four electrons to O, in pairs. We get:

 or 

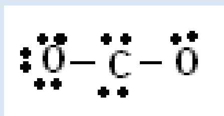
Check:
Is the total number of valence electrons correct? Yes. Our final picture has 8 valence e-.
Does each atom have the appropriate duet or octet of electrons? Yes

Example: Draw a Lewis structure for CO_2

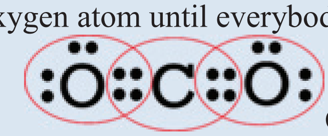
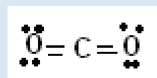
Solution:
1) add up all available valence electrons:
1(4) + 2(6) = 16
2) Draw a skeleton.
Carbon goes in the middle with the two oxygen atoms bonded to it:



3) Use the remaining electrons to give each atom (except hydrogen) an octet.
In this case, we have already used up four electrons to draw the two bonds in the skeleton, leaving 12 left. This is not enough to give everybody an octet. Our picture may look something like this with 16 electrons:

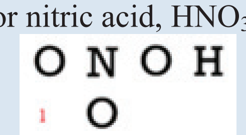


We have used up the 16 electrons, but neither O has an octet. The rules state that if you run out of electrons and still don't have octets, then you must use some of the unshared pairs of electrons as double or triple bonds instead. Move the electrons that are just on the carbon atom to share with the oxygen atom until everybody has an octet. We get:

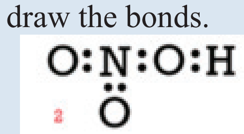
 OR 

Check:
Is the total number of valence electrons correct? Yes. Our final picture has 16 valence e-.
Does each atom have the appropriate duet or octet of electrons? Yes

Example: Draw a Lewis structure for nitric acid, HNO_3. The skeleton is given below:

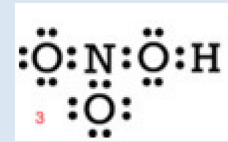


Solution:
1) add up all available valence electrons:
1(1) + 1(5) + 3(6)=24
2) Draw a skeleton.
This was given to us, but we need to draw the bonds.



3) Use the remaining electrons to give each atom (except hydrogen) an octet.
Each bond used up 2 electrons, so we have already used 8 electrons. If we use the reaining

16 electrons, we may get a picture such as:



But notice that the nitrogen atom still does not have an octet. We ran out of electrons so we must form a double bond. Use some of the electrons on an oxygen atom to share with the nitrogen. We get:

Check:
Is the total number of valence electrons correct? Yes. Our final picture has 24 valence e-.
Does each atom have the appropriate duet or octet of electrons? Yes

Lesson Summary
- Covalent bonds are formed by electrons being shared between two atoms.
- Half-filled orbitals of two atoms are overlapped and the valence electrons shared by the atoms.
- Covalent bonds are formed between atoms with relatively high electron affinity.

Vocabulary
- Covalent bond: A type of bond in which electrons are shared by atoms.

Further Reading / Supplemental Links
- Tutorial on bonding: http://visionlearning.org/library/module_viewer.php?mid=55&l=

4.6: Review Questions
Which of the following compounds would you expect to be ionically bonded and which covalently bonded?

1) CS_2
2) K_2S
3) FeF_3
4) PF_3
5) AlF_3
6) BaS

Draw a Lewis structure for each of the following compounds.

7) H_2O
8) CH_4
9) CO
10) PCl_3
11) C_2H_6
12) CO_3^{2-}
13) CO_2
14) NH_3
15) CH_2O
16) SO_3

4.7: 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.

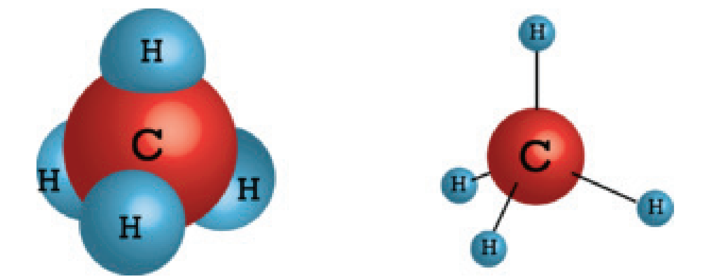
Introduction

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.

Predicting the Shape of Molecules

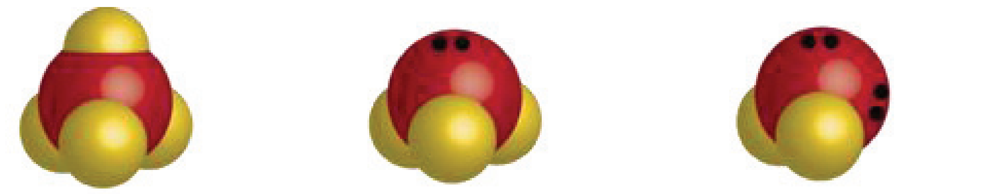
Consider, methane, commonly known as natural gas. In this molecule, carbon has four valence electrons and each hydrogen adds one more so the central atom in methane has four pairs of electrons in its valence shell. The 3d shape of this molecule is dictated by the repulsion of the electrons. Those four 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°.



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

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**.



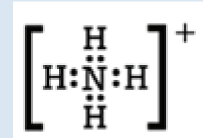
Various molecular geometries with four pairs of electrons around a central atom.

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.

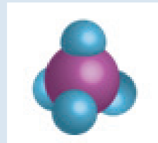
Summary of Molecular Geometry

# 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:

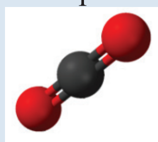


Solution: To answer this question, you need to count the number of atoms around the central atom and the number of unshared pairs. In this example, there are four atoms bonded to the N with zero unshared pairs of electrons. The shape must be tetrahedral.



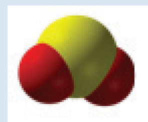
Example: Determine the shape of carbon dioxide, CO_2, given by the following Lewis structure:

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

4.7: Review Questions

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

1) CH_3Cl

```
    H
    |
Cl—C—H
    |
    H
```

2) Silicon tetrafluoride, SiF_4

```
   F
   |
F—Si—F
   |
   F
```

3) $CHCl_3$

```
    Cl
    |
H—C—Cl
    |
    Cl
```

4) Water, H_2O

```
  . .
H—O—H
  ` `
```

5) Ammonia, NH_3

```
    H
    |
H — N :
    |
    H
```

6) Methane, CH_4

```
    H
    |
H—C—H
    |
    H
```


4.8: Polarity & 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

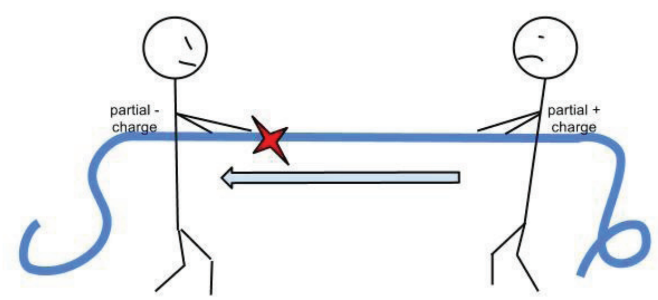
The ability of an atom in a molecule to attract shared electrons is called **electronegativity.** When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, 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.

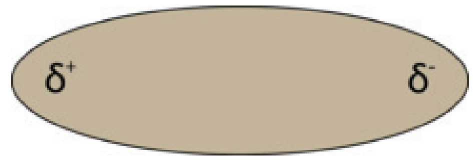
The other case is when the bonded atoms have a very large difference in their electronegativities. In this case, the more electronegative atom will take the electrons completely away from the other atom and an ionic bond forms.

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



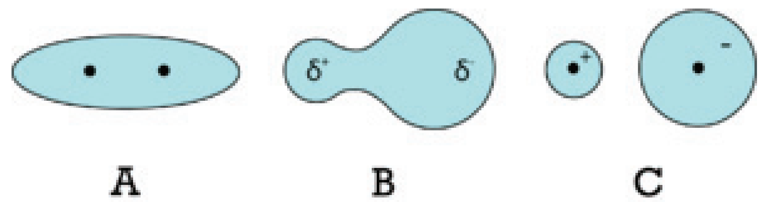
A polar covalent bond is similar to a tug-of-war in which one atom pulls more on the electrons and gains a partial negative charge. The weaker (less electronegative atom) has a partial positive charge.
CC – Tracy Poulsen

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.

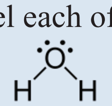


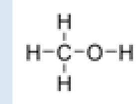
A) A nonpolar covalent bond in which two identical atoms are sharing electrons
B) a polar covalent bond in which the more electronegative atom pulls the electrons more toward itself (forming partial negative and positive sides)
C) an ionic bond in which an extremely electronegative atom is bonded to 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.

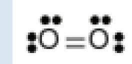
In order to determine if a molecule is polar or nonpolar, it is frequently useful to look a Lewis structures. Nonpolar compounds will be symmetric, meaning all of the sides around the central atom are identical – bonded to the same element with no unshared pairs of electrons. Polar molecules are assymetric,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

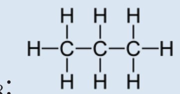
Example: Label each of the following as polar or nonpolar.

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.

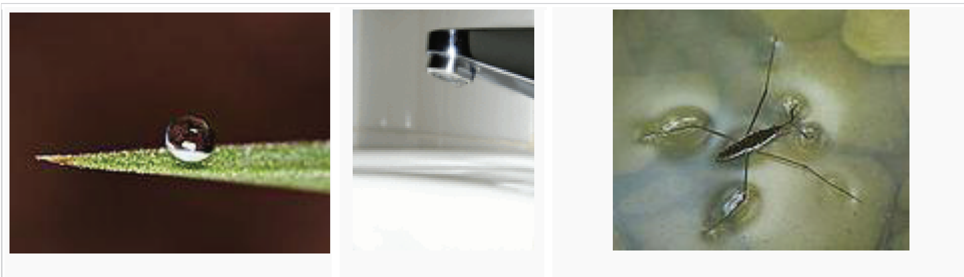
Hydrogen bonding strongly affects the crystal structure of ice, helping to create an open hexagonal lattice. The density of ice is less than water at the same temperature; thus, the solid phase of water floats on the liquid, unlike most other substances in which the solid form would sink in the liquid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

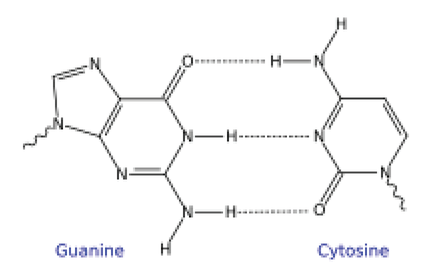
Have you ever experienced a belly flop? This is also due to the hydrogen bonding between water molecules, causing surface tension. On the surface of water, water molecules are even more attracted to their neighbors than in the rest of the water. This attraction makes it difficult to break through, causing 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 **B.** Water dripping from a tap **C.** Water striders stay atop the liquid due to surface tension

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.



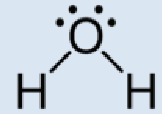
Hydrogen bonding between guanine (G) and cytosine(C), one of two types of base pairs in DNA. The hydrogen bonds are shown by dotted lines.

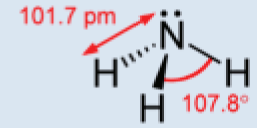
As a result of the strong attraction between molecules that occurs in a hydrogen bond, the following properties can be summarized. Molecules with hydrogen bonding tend to:

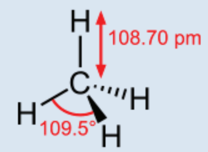
- have higher melting points than polar molecules

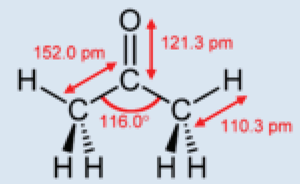
- have higher boiling points than polar molecules
- be more soluble in water (dissolve better) than polar molecules

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 CH_3 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) H_2O (polar, hydrogen bonding) vs. acetone (polar, no hydrogen bonding). H_2O will have a higher melting point because compounds with hydrogen bonding tend to have higher melting points than polar compounds.

b) CH_4 (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=

4.8: 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.

Given each of the following Lewis structures, indicate whether each is polar or nonpolar. Then indicate whether or not that compound exhibits hydrogen bonding.

7) CH_3Cl 8) Silicon tetrafluoride, SiF_4 9) $CHCl_3$

```
      H                 F                  Cl
      |                 |                  |
  Cl—C—H             F—Si—F             H—C—Cl
```



```
      |                 |                  |
      H                 F                  Cl
```

10) Water, H_2O 11) Ammonia, NH_3 12) Methane, CH_4

```
     . .              H                  H
   H—O—H              |                  |
     ` `           H — N :            H—C—H
                      |                  |
                      H                  H
```

For each of the following, indicate which of the compounds in the pair has the given property.

13) higher melting point: ammonia or methane
14) higher boiling point: water or CH_3Cl
15) more soluble in water: ammonia or $CHCl_3$
16) higher melting point: SiF_4 or ammonia

Chapter 5: Problem Solving & the Mole

5.1: Measurement Systems

Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit
- Explain the difference between mass and weight.
- Identify SI units of mass, distance (length), volume, temperature, and time.
- Describe absolute zero.

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. Mankind's first elementary efforts in measurement required convenient objects to be used as standards and the human body was certainly convenient. The names of several measurement units reflect these early efforts. 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.

It should be apparent that measurements of a foot by two people could differ by a few inches. To achieve more consistency, everyone could use the king's foot 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 was needed were objects that could be safely stored so they didn't change over time. Copies could be made of these objects and distributed so that everyone was using exactly the same units of measure. The requirements of science in the 1600s, 1700s, and 1800s necessitated even more 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. Let's consider the English system – that is, the one that is in everyday use in the US– to explain why the metric system is so much easier to manipulate. For instance, if you need to know how many inches are in a foot, you only need to remember what you at one time memorized: 12 inches = 1 foot. But now you need to know how many feet are in a mile. What happens if you never memorized this fact? Of course you can look it up online or elsewhere, but the point is that this fact must be given to you, as there is no way for you to derive it out yourself. This is true about all parts of the English 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 measure is now $\frac{1}{100}$ of a gram. When milli- is placed in front of meter, as in millimeter, the measure is now $\frac{1}{1000}$ of a meter. Common prefixes are shown in the table.

Common Prefixes in the International System		
Prefix	***Meaning***	***Symbol***
micro-	10^{-6}	µ
milli-	10^{-3}	m
centi-	10^{-2}	c
kilo-	10^{3}	k

These prefixes are used for all metric units of measurement, including units for volume, time, distance, etc. Common metric units, symbols, and relationships to a base unit are shown below.

$$1 \text{ micrometer} = 1\ \mu\text{m} = 1\text{x}10^{-6} \text{ m}$$
$$1 \text{ microliter} = 1\ \mu\text{L} = 1\text{x}10^{-6} \text{ L}$$
$$1 \text{ kilometer} = 1 \text{ km} = 1\text{x}10^{3} \text{ m}$$
$$1 \text{ kilogram} = 1 \text{ kg} = 1\text{x}10^{3} \text{ g}$$

SI Units

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 internationally agreed upon 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 added as advancements in science are made. Each SI base unit except the kilogram is described by stable properties of the universe.

Mass

Mass and weight are not the same thing. Although we often use the terms mass and weight interchangeably, each one has a specific definition and usage. The **mass** of an object is a measure of the amount of matter in it. The mass (amount of matter) of an object remains the same regardless of where the object is placed. For example, moving a brick to the moon does not cause any matter in it to disappear or be removed.

The balance on the left measures mass, a property that does not change based on location. The scale on the right measures weight, which differs depending on where in the universe an object is.

The **weight** of an object is the force of attraction between the object and the earth (or whatever large body it is resting on). We call this force of attraction 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 depending 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 gravity is only one-sixth that of Earth. The mass of this man, however, would be the same in each situation because the amount of matter in him is constant. Consistency requires that scientists use mass and not weight in its measurements of the amount of matter.

The basic unit of mass in the International System of Units is the kilogram. A **kilogram** is equal to 1000 grams. A gram is a relatively small amount of mass and so larger masses are often expressed in kilograms. When very tiny amounts of matter are measured, we often use milligrams which are equal to 0.001 gram. There are numerous larger, smaller, and intermediate mass units that may also be appropriate.

The standard kilogram copy stored and used in Denmark.

At the end of the 18th century, a kilogram was the mass of a liter of water. In 1889, a new international prototype of the kilogram was made of a platinum-iridium alloy. The kilogram is equal to the mass of this international prototype, which is held in Paris, France.

Length

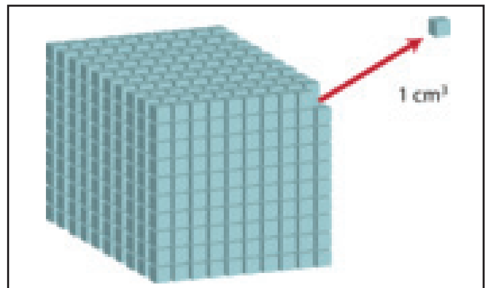
Length is the measurement of anything from end to end. In science, length usually refers to how long an object is. There are many units and sets of standards used in the world for measuring length. The ones familiar to you is probably inches, feet, yards, and miles. Most of the world, however, measures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

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 a bar 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.

Volume

The **volume** of an object is the amount of space it takes up. In the SI system, volume is a **derived unit**, that is, it is based on another SI unit. In the case of volume, a cube is created with each side of the cube measuring 1.00 meter. The volume of this cube is $1.0\text{m}\cdot 1.0\text{m}\cdot 1.0\text{m}=1.0\text{ m}^3$ or one cubic meter.



The cubic meter is the SI unit of volume. The cubic meter is a very large unit and is not very convenient for most measurements in chemistry. A more common unit is the liter (L) which $\frac{1}{1000}$ of a cubic meter. Another commonly used volume measurement is the milliliter, which is equal to $\frac{1}{1000}$ of a liter.

Temperature

When used in a scientific context, the words heat and temperature do NOT mean the same thing. Temperature represents the average kinetic energy of the particles that make up a material. Increasing the temperature of a material increases its thermal energy. Thermal energy is the sum of the kinetic and potential energy in the particles that make up a material. Objects do not "contain" heat; rather they contain thermal energy. Heat is the movement of thermal energy from a warmer object to a cooler object. When thermal energy moves from one object to another, the temperature of both objects change.

A thermometer is a device that measures temperature. The name is made up of "thermo" which means heat and "meter" which means to measure. The temperature of a substance is directly proportional to the average kinetic energy it contains. In order for the average kinetic energy and temperature of a substance to be directly proportional, it is necessary that when the temperature is zero, the average kinetic energy must also be zero. This is not true with either the Fahrenheit or Celsius temperature scales. Most of are familiar with temperatures that are below the freezing point of water. It should be apparent that even though the air temperature may be -5°C, the molecules of air are still moving. Substances like oxygen gas and nitrogen gas have already melted and boiled to vapor at temperatures below -150°C.

It was necessary for use in calculations in science for a third temperature scale in which zero degrees corresponds with zero kinetic energy, that is, the point where molecules cease to move. This temperature scale was designed by Lord Kelvin. Lord Kelvin stated that there is no upper limit of 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 at which molecules stop moving and therefore, have zero kinetic energy. The **Kelvin temperature scale** has its zero at absolute zero (determined to be -273.15°C), and uses the same size degree as the Celsius scale. Therefore, the mathematical relationship between the Celsius scale and the Kelvin scale is $K={}^\circ C + 273$. In the case of the Kelvin scale, the degree sign is not used. Temperatures are expressed, for example, simply as 450 K.

Time

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 the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom.

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.

- The International System of Units, abbreviated SI from the French *Le Système International d' Unites* is, since the 1960s, internationally agreed upon as the standard metric system.
- 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 SI unit of length is the meter.
- Temperature represents the average kinetic energy of the particles that make up a material.
- Absolute Zero is the temperature at which molecules stop moving and therefore, have zero kinetic energy.
- The Kelvin temperature scale has its zero at absolute zero (determined to be -273.15°C), and uses the same size degree as the Celsius scale.
- The mathematical relationship between the Celsius scale and the Kelvin scale is K=°C + 273.
- The SI unit for time is the second.

Vocabulary
- Metric system: an international decimal-based system of measurement.
- International System of Units (Le Système International d' Unites): the internationally agreed upon standard metric system
- Mass: a measure of the amount of matter in an object
- Weight: the force of attraction between the object and the earth (or whatever large body it is resting on)
- Temperature: the average kinetic energy of the particles that make up a material
- Absolute Zero: the temperature at which molecules stop moving and therefore, have zero kinetic energy

Further Reading / Supplemental Links
- http://en.wikipedia.org/wiki/Si_units

5.1: Review Questions
1) List three advantages to using the metric system over the English system or other measurement systems.

Identify which is bigger in each set of measurements:
2) 1 kg or 1 g
3) 10 mg or 10 g
4) 100 cg or 100 mg

Fill in the missing information in the following equivalencies:
5) ? g = 1 kg
6) 100 ? = 1 L
7) 1 m = ? cm

8) Why is it important for scientists to use the same system to make measurements?
9) What is the basic unit of measurement in the metric system for length?

10) What is the basic unit of measurement in the metric system for mass?
11) What unit is used in the metric system to measure volume? How is this unit related to the measurement of length?
12) Would it be comfortable to swim in a swimming pool whose water temperature was 275 K? Why or why not?

5.2: Scientific Notation

Objectives
- Identify when scientific notation is useful to record measurements
- Convert measurements to scientific notation.
- Convert quantities from scientific notation to their standard numerical form.

Introduction

Work in science frequently involves very large and very small numbers. The speed of light, for example, is 300,000,000 meters/second; 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 such numbers and even more inconvenient to attempt to carry out mathematical operations with them.

What is Scientific Notation?

Scientists and mathematicians have designed an easier method for dealing with such numbers. This more convenient system is called exponential notation by mathematicians and scientific notation by scientists.

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. The number 9,000,000 for example, can be written as the product of 9 times 1,000,000 and 1,000,000 can be written as 10^6. Therefore, 9,000,000 can be written as $9x10^6$. In a similar manner, 0.00000004 can be written as $4 \times \frac{1}{10^8}$ or $4x10^{-8}$.

Examples of Scientific Notation	
Decimal Notation	**Scientific Notation**
95,672	$9.5672x10^4$
8,340	$8.34x10^3$
100	$1x10^2$
7.21	$7.21x10^0$
0.014	$1.4x10^{-2}$
0.000000008	$8.0x10^{-9}$
0.00000000000975	$9.75x10^{-12}$

As you can see from the examples above, to convert a number from decimal form to scientific notation, you count the spaces that you need 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.

Calculators and Scientific Notation

Because you will be performing calculations using scientific notation, it is important that you understand how your calculator uses scientific notation. For this course, you will need a scientific or graphing calculator. These calculators have an "exponential" button. It is typically labeled "EXP" or "EE" and can be read "times ten to the…" Consider the following number: 6.02×10^{23}. If you were to read this out loud, you would say "6.02 times ten to the 23rd". To type this in your calculator, you would put "6.02EE23" or "6.02EXP23". Most calculators would print 6.02E23. By properly using the exponential button, you will avoid common mistakes made by students when multiplying or dividing these numbers.

Example: Perform the following calculation correctly using the exponential button. $1.20 \times 10^{24} / 6.02 \times 10^{23}$
Solution: To type this in my calculator, I would type: $1.20EE24 \div 6.02EE23$ OR $1.20EXP24 \div 6.02EXP23$ The answer is 2.0. *Be careful to look on the far right of your calculator screen. If you see E47 or a 47 typed offset, you typed it in wrong. Try again.

Lesson Summary

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

Vocabulary

- Scientific notation: a shorthand method of writing very large and very small numbers in terms of a decimal number between 1 and 10 multiplied by 10 to a power.

5.2: Review Questions

1) When is it useful to use scientific notation?

Write the following numbers in scientific notation.

2) 0.0000479
3) 4260
4) 251,000,000
5) 0.00206

Write each of the following numbers in standard notation.

6) 2.3×10^{4}
7) 9.156×10^{-4}
8) 7.2×10^{-3}
9) 8.255×10^{6}

Write the buttons you would push on your calculator to type in the following numbers using the exponential button.

10) 7.3×10^{14}
11) 6.01×10^{-6}
12) 7.98×10^{5}

Using the exponential button, perform each calculation on your calculator.

13) $2.0 \times 10^{3} \cdot 3.0 \times 10^{4}$
14) $2.0 \times 10^{3} / 3.0 \times 10^{4}$
15) $4.2 \times 10^{-4} / 3.0 \times 10^{-2}$
16) $7.3 \times 10^{-7} \cdot 8.0 \times 10^{-3}$

5.3: Math in Chemistry

Objectives

- Use the factor-label method to solve problems.
- Perform metric conversions using the factor label method for conversions.

Introduction

During your studies of chemistry (and physics also), you will note that mathematical equations are used in a number of different applications. Many of these equations have a number of different variables with which you will need to work. You should also note that these equations will often require you to use measurements with their units. Algebra skills become very important here!

Conversion Factors

A **conversion factor** is a factor used to convert one unit of measurement into another. 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 always must be remembered is that a conversion factor has to represent a fact; this fact can either be simple or much more complex. For instance, you already know that 12 eggs equal 1 dozen. A more complex fact is that the speed of light is 1.86×10^{5} miles/sec. Either one of these can be used as a conversion factor depending on what type of calculation you might be working with.

Factor-Label Method of Problem Solving

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 unit analysis or dimensional analysis. **The Factor-Label Method** is a technique that involves the study of the units of physical quantities. It affords a convenient means of checking mathematical equations. This method involves considering both the units you presently have (given measurement), the units you wish to end up with, and designing conversion factors than 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 or ratios of how units are related to each other.

In terms of making unit conversions, suppose you want to convert 0.0856 meters into millimeters. In this case, you need only one conversion factor and that conversion factor must cancel the meters unit and create the millimeters unit. The conversion factor will be created from the relationship1000 millimeters (mm) = 1 meter (m).

$$0.0856\ \cancel{m} \cdot \frac{1000\ mm}{1\ \cancel{m}} = 85.6\ mm$$

Remember that when you multiply fractions and you have the same number on top of one fraction and the bottom of another fraction, the numbers will cancel out leaving one. The same is true for units. When the above expression is multiplied as indicated, the meters units will cancel and only millimeters will remain. The unit analysis process involves creating conversion factors from equivalencies between various units.

The given table contains many useful conversion factors.

English Units	Metric Units
1 ounces (oz) (weight)	28.35 grams (g)
1 fluid ounce (oz) (volume)	29.6 mL
2.205 pounds (lb)	1 kilograms (kg)
1 inch (in)	2.54 centimeters (cm)
.6214miles (mi)	1 kilometer (km)
1 quart (qt)	0.95 liters (L)

Metric Prefix	Base unit equivalency
1000 milli (base unit)	1 base unit
100 centi (base unit)	1 base unit
1 kilo (base unit)	1000 base units

Of course, there are other ratios which are not listed in this table. They may include:

- Ratios embedded in the text of the problem (using words such as *per* or *in each*, or using symbols such as / or %)
- Conversions in the metric system, as covered earlier in this chapter.
- Common knowledge ratios (such as 60 seconds = 1 minute)

The general steps you must take in order to solve these problems include:

1. Identify the **"given"** information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to **"find"**? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

Look for each of these steps in the following examples.

Example: Convert 1.53 grams to centigrams.

Solution: The equivalency relationship is 100cg=1g (given in the second table), so the conversion factor is constructed from this equivalency to cancel *grams* and produce *centigrams*.

$$1.53\ \cancel{g} \cdot \frac{100\ cg}{1\ \cancel{g}} = \mathbf{153\ cg}$$

Example: Convert 1000 inches to feet.

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

$$1000\ \cancel{inches} \cdot \frac{1\ foot}{12\ \cancel{inches}} = \mathbf{83\ feet}$$

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

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 mile = 5280 feet*, *12 inches = 1 foot*, *2.54 cm = 1 inch*, *100 cm = 1m*, 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 miles to kilometers.

Solution: Although we have a ratio for miles to kilometers given in the table, we will solve this problem using other units to see what a longer process looks like. The answer would be the same.

$$12\ \cancel{miles} \cdot \frac{5280\ \cancel{feet}}{1\ \cancel{mile}} \cdot \frac{12\ \cancel{inches}}{1\ \cancel{foot}} \cdot \frac{2.54\ \cancel{cm}}{1\ \cancel{inch}} \cdot \frac{1\ \cancel{m}}{100\ \cancel{cm}} \cdot \frac{1\ km}{1000\ \cancel{m}} = \mathbf{19\ km}$$

In each step, the previous unit is cancelled and the next unit in the sequence is produced., each successive unit cancelling out until only the unit needed in the answer is left.

Conversion factors for area and volume can also be produced by this method.

Example: Convert 1500 cm^2 to m^2.

Solution:

$$1500\ \cancel{cm^2} \cdot \left(\frac{1\ m}{100\ \cancel{cm}}\right)^{\cancel{2}} = \mathbf{0.15\ m^2}$$

OR

$$1500\ \cancel{cm^2} \cdot \left(\frac{1\ m^2}{10{,}000\ \cancel{cm^2}}\right) = \mathbf{0.15\ m^2}$$

Lesson Summary

- Conversion factors are used to convert one unit of measurement into another.
- The factor-label method involves considering both the units you presently have, the units you wish to end up with and designing conversion factors than will cancel units you don't want and produce units you do want.

Vocabulary

- Conversion factor: a ratio used to convert one unit of measurement into another.

Further Reading / Supplemental Links

- Tutorial: Vision Learning: Unit Conversion & Dimensional Analysis http://visionlearning.com/library/module_viewer.php?mid=144&l=&c3=

5.3: Review Questions

For each of the following, first A) identify the given, find, and ratios within the problem. Then, B) solve the problem using the factor-label method. Show all work and unit cancellations.

1) What is the diameter of a 9" cake pan in centimeters?
2) It is approximately 52 miles from Spanish Fork to Salt Lake. If I drive 65 miles/hr, how many minutes will it take to drive there?

3) If there are 35 g of sugar in 8 oz of soda, what mass (in grams) of sugar is in an entire 2 liter bottle?
4) My car gets about 35 miles per gallon. Right now, gas costs $3.69 per gallon. How much does it cost me to drive to Salt Lake (52 miles away)?
5) What is your mass in grams? (Start with your weight in pounds)
6) Nervous Ned paced for 3 hours while his wife was in the delivery room. If he paces at 5 paces every 3 seconds, how far did he go, in miles? (In this case, one pace is 2.2 feet.) (There are 5280 feet per mile.)
7) If I drive 75 miles/hr, how long in minutes will it take me to drive 500 km?
8) A male elephant seal weighs about 4 tons. What is the mass of the seal in grams? (There are 2000 lbs in one ton)
9) My car gets about 37 miles per gallon. How many km/liter is this? (There are 4 quarts in a gallon)
10) 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.

5.4: 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 6.02×10^{23} particles and the name of this group is the **mole** (the abbreviation for ***mole*** is ***mol***). One mole of any object is 6.02×10^{23} 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 6.02×10^{23} particles. Eventually chemists decided to call that number of particles a mole. The number 6.02×10^{23} 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 (1×10^9) molecules of water?	
Solution: $1 \times 10^9 \cancel{molecules\ H_2O} \cdot \frac{1\ mol\ H_2O}{6.02 \times 10^{23} \cancel{molecules\ H_2O}} = \mathbf{1.7 \times 10^{-15} mol\ H_2O}$	

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: $0.00100\ \cancel{mol} \cdot \frac{6.02 \times 10^{23} molecules}{1\ \cancel{mol}} = \mathbf{6.02 \times 10^{20}\ molecules}$	

Converting Grams to Moles and Vice Versa

1.00 mol of carbon-12 atoms has a mass of 12.0 g and contains 6.02×10^{23} atoms. Likewise, 1.00 mol of water has a mass of 18.0 grams and contains 6.02×10^{23} molecules. 1.00 mole of any element or compound has a mass equal to its molecular mass in grams and contains 6.02×10^{23} 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 H_2O. 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 H_2O=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:	c) F_2 d) H_2SO_4

a) S b) H_2O	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: $108\ \cancel{g\ H_2O} \cdot \frac{1\ mol\ H_2O}{18.02\ \cancel{g\ H_2O}} = 5.99\ mol\ H_2O$

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: $7.50\ \cancel{mol\ CaO} \cdot \frac{56.0\ g\ CaO}{1\ \cancel{mol\ CaO}} = 420.\ g\ CaO$

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 6.02×10^{23} 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 6.02×10^{23}.
- 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=

5.4: Review Questions

How many molecules are present in the following quantities?
1) 0.250 mol H_2O
2) 0.0045 mol $Al_2(CO_3)_3$

How many moles are present in the following quantities?
3) 1.0×10^{20} molecules H_2O
4) 5 billion atoms of carbon

What is the molar mass of each of the following substances? Include units with your answer.
5) H_2O
6) NaOH
7) NH_4Cl
8) H_2SO_4
9) $Al_2(CO_3)_3$
10) PbO_2

Convert the following to moles.
11) 60.0 g NaOH
12) 5.70 g H_2SO_4
13) 2.73 g NH_4Cl
14) 10.0g PbO_2

Convert the following to grams.
15) 0.100 mol CO_2
16) 0.500 mol $(NH_4)_2CO_3$
17) 0.437 mol NaOH
18) 3.00 mol H_2O

How many molecules are present in the following masses?
19) 1.00 g Na_2CO_3
20) 1000. g H_2O

Convert the following to grams.
21) 2.0×10^{23} molecules H_2
22) 1.75×10^{24} molecules NaCl
23) 8.6×10^{22} molecules NaOH

Chapter 6: Mixtures & Their Properties

6.1: Solutions, Colloids, and Suspensions

Objectives

- Define a solution.
- Identify the solute and solvent in a solution.
- Explain the differences among solutions and heterogeneous mixtures, such as colloids and suspensions.

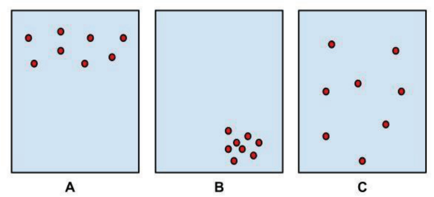
Introduction

In chapter 4, we distinguished between pure substances and mixtures. Remember, a mixture contains two or more pure substances that are not bonded together. These substances remain unbounded to each other, but are mixed within the same container. They also retain their own properties, such as color, boiling point, etc. There are two types of mixtures: mixtures in which the substances are evenly mixed together (called a **solution**) and a mixture in which the substances are not evenly mixed (called a **heterogeneous mixture**).

In this chapter we begin our study of solution chemistry. We all probably think we know what a solution is. We might be holding a can of soda or a cup of tea while reading this book and think … hey this is a solution. Well, you are right. But you might not realize that alloys, such as brass, are also classified as solutions, or that air is a solution. Why are these classified as solutions? Why wouldn't milk be classified as a solution? To answer these questions, we have to learn some specific properties of solutions. Let's begin with the definition of a solution and view some of the different types of mixtures.

Homogeneous Mixtures

A **solution** is an even (or homogeneous) mixture of substances. When you consider that the prefix "homo" means "same", this definition makes perfectly good sense. Solutions carry the same properties throughout. Take, for example, vinegar that is used in cooking is approximately 5% acetic acid in water. This means that every teaspoon of vinegar that is removed from the container contains 5% acetic acid and 95% water. This ratio of mixing is carried out throughout the entire container of vinegar.

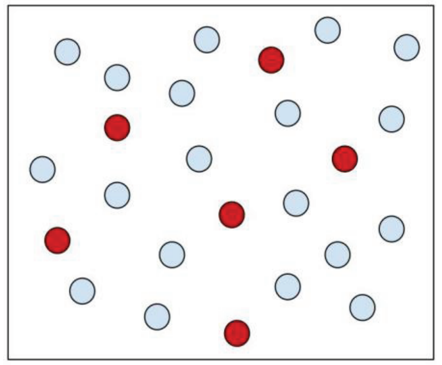


Solutions must have particles that are uniformly spread throughout. In this image, A and B are not solutions as the red particles are not evenly spread. Mixture C, however, is a solution.
CC – Tracy Poulsen

A point should be made here that when a solution is said to have uniform properties throughout, the definition is referring to properties at the particle level. Well, what does this mean? Let's consider brass as an example. The brass is an alloy made from copper and zinc. 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) and the copper and zinc atoms are evenly mixed at the atomic level. So the brass represents a homogeneous mixture. 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. The ratio of copper and zinc may be different. Additionally, you would see differences in the color at different places in the mixture (there are visible places in which there are more copper atoms

and visible places in which there are more zinc atoms). Thus the combination of zinc filings and copper pieces in a pile does not represent a homogeneous mixture, but is, instead a heterogeneous mixture. In a solution, the particles are so small that they cannot be distinguished by the naked eye. In a solution, the mixture would have the same appearance and properties in all places throughout the mixture.

The point should be made that because solutions have the same composition throughout does not mean you cannot vary the composition. If you were to take one cup of water and dissolve ¼ teaspoon of table salt in it, a solution would form. The solution would have the same properties throughout, the particles of salt would be so small that they would not be seen and the composition of every milliliter of the solution would be the same. But you can vary the composition of this solution to a point. If you were to add another ½ teaspoon of salt to the cup of water, you would make another solution, but this time there would be a different composition than the last. You still have a solution where the salt particles are so small that they would not be seen and the solution has the same properties throughout, thus it is homogeneous.



In this image, the dark particles are the solvent particles as there are fewer of them. The solute particles are the lighter colored particles.
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The solvent and solute are the two basic parts of a solution. The **solvent** is the substance present in the greatest amount, whichever substance there is more of in the mixture. The solvent is frequently, but not always, water. The **solute**, then, is the substance present in the least amount. Let's think for a minute that you are making a cup of hot chocolate. You take a teaspoon of cocoa powder and dissolve it in one cup of hot water. Since the cocoa powder is in the lesser amount it is said to be the solute; and the water is the solvent since it is in the greater amount.

Example: Name the solute and solvent in each of the following solutions.
(a) salt water
(b) air

Solution:
(a) solute = salt; solvent = water
(b) solute = oxygen; solvent = nitrogen

Colloids and Suspensions

Two other types of mixtures that we will compare to solutions include colloids and suspensions. These mixtures are frequently confused with solutions, but these are heterogeneous, not homogeneous, mixtures.

Recall that a solution is a mixture of substances in such a way that the final product has the same composition throughout. Remember the example of vinegar that is 5%, by mass, acetic acid in water. This clear liquid is a solution since light easily passes through it and it never separates. All liquid solutions have this shared property, in which the particles are so small that light goes straight through. In other words, the mixture is clear or see-through. It is important to note, however, that clear does not necessarily mean colorless.

On the other hand, **colloids** are mixtures in which the size of the particles is between 1×10^3 pm and 1×10^6 pm. In meters, these sizes translate to 1×10^{-9} m to 1×10^{-6} m – between 10 and 1000 times smaller than a small grain. These particles, although sounding small, are still much bigger than the particles in a solution.

A common example of a colloid is milk. One way to tell that milk is a colloid is by the **Tyndall effect**. The Tyndall effect is the scattering of light by particles. This involves shining a light through the mixture: when the light is shined through a colloid, the light is does not go straight through, but is instead scattered. Note that milk is *not* see-through, but has a cloudy appearance. Because light not allowed to pass through the mixture, the mixture is considered a colloid. When light is passed through a solution, the particles are so small that they do not obstruct the light. However, when light is passed through a colloid, since the particles are larger, they will act as an obstruction to the light and the light is scattered. The particles in a colloid, while able to scatter light, are still small enough so that they do not settle out of solution.

It is amazing just how common colloids are to us in our everyday lives. Some common colloids you may have seen include milk of magnesia, mayonnaise, jell-o, and marshmallows.

Suspensions are mixtures which contain even bigger particles than solutions or colloids do. In suspensions, particles settle into layers within a container if they are left standing. This means that the particles in a suspension are large enough so that gravity pulls them out of solution. With suspensions, filtration can usually be used to separate the excess particles from the solution. If a suspension is passed through a piece of filter paper (or a coffee filter) some particles will go through and others will be stopped in the filter paper. A common example of a suspension is muddy water. If you had a beaker of water and added a handful of fine dirt, even if you stirred it, when you let it stand, dirt would settle to the bottom.

Neither colloids nor suspensions are classified as solutions, but are special types of heterogeneous mixtures instead. In order to be as solution, you must have very small particles evenly distributed, so that the mixture has the same properties throughout. Colloids and suspensions have particles that are too big to be considered a solution.

Example: Label each of the following mixtures as a solution, colloid, or suspension.
a) Italian salad dressing
b) Mustard
c) Apple juice

Solution:
a) suspension – when left to sit, it separates into layers
b) colloid – although it does not separate into layers like suspensions do, mustard does not let light go through
c) solution – apple juice doesn't separate into layers like suspensions do, but apple juice will let light through so it is a solution and not a colloid.

Lesson Summary

- Generally speaking, 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.
- Common examples of colloids include milk, butter, Jell-O, and clouds.
- Suspensions are mixtures in which the particles are large enough so that they settle to the bottom of the container and can be filtered using filter paper.

Vocabulary

- Solution: a homogeneous mixture of substances
- Solvent: the substance in a solution present in the greatest amount
- Solute: the substance in a solution present in the least amount
- Colloid: type of mixture in which the size of the particles is between 1×10^3 pm and 1×10^8 pm
- Suspension: type of mixture in which the particles settle to the bottom of the container and can be separated by filtration

6.1: Review Questions

1) Distinguish between a solution, a colloid, and a suspension.
2) What is one way to tell you have a colloid and not a solution?

Multiple choice

3) The biggest difference between a colloid and a suspension is that:
   a) In colloids, the solute is permanently dissolved in the solvent.
   b) In colloids the particles eventually settle to the bottom.
   c) In suspensions the particles eventually settle to the bottom.
   d) None of these are correct
4) Karen was working in the lab with an unknown solution. She noticed that there was no precipitate in the bottom of the beaker even after it had been on the lab bench for several days. She tested it with a light and saw that light scattered as it passed through the solution. Karen concluded that the liquid was what type of a mixture?
   a) colloid
   b) suspension
   c) solution

6.2: Solution Formation

Objectives

- Explain why solutions form.
- Explain the significance of the statement "like dissolves like."
- Discuss the idea of water as the "universal solvent".
- Explain how water molecules attract ionic solids when they dissolve in water.

Introduction

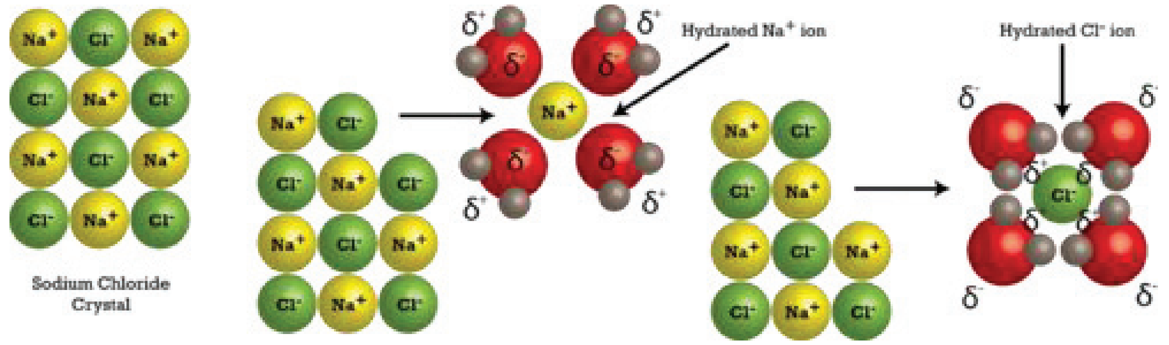
We have learned that solutions can be formed in a variety of combinations using solids, liquids, and gases. We also know that solutions have constant composition and we can also vary this composition up to a point to maintain the homogeneous nature of the solution. But how exactly do solutions form? Why is it that oil and water will not form a solution and yet vinegar and water will? Why could we dissolve table salt in water but not in vegetable oil? The reasons why solutions form will be explored in this section, along with a discussion of why water is used most frequently to dissolve substances of various types.

Ionic Compounds in Solution

Recall that metals form positive ions by losing electrons and nonmetals form negative ions by gaining electrons. In ionic compounds, the ions in the solid are held together by the attraction of these oppositely charges particles. Since ionic compounds can dissolve in polar solutions, specifically water, we can extend this concept to say that ions themselves are attracted to the water molecules because the ions of the ionic solid are attracted to the polar water molecule. When you dissolve table salt in a cup of water, the table salt dissociates into sodium ions and chloride ions:

$$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$$

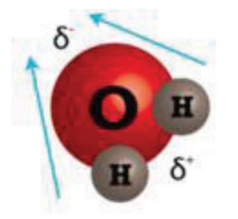
How does salt dissolve, though? Dissolving is based on electrostatic attraction, that is, the attraction between positive and negative charges. The sodium ions get attracted to the partially negative ends of the water molecule and the chloride ions get attracted to the partially positive end of the water molecule.



Most ionic compounds dissolve in water as the positive ions are attracted to the negative side of water molecules and the negative ions are attracted to the positive sides of water molecules.

To understand why salt will dissolve in water, we first must remember what it means for water to be polar. The more electronegative oxygen atom pulls the shared electrons away from the hydrogen atoms in a water molecule causing an unequal distribution of electrons. The hydrogen end of the water molecule will be slightly positive and the oxygen end of the water molecule will be slightly negative. These partial charges allow water to be attracted to the various ions in salt, which pulls the salt crystal apart.

The same is process true for any ionic compound dissolving in water. The ionic compound will separate into the positive and negative ions and the positive ion will be attracted to the partially negative end of the water molecules (oxygen) while the negative ion will be attracted to the partially positive end of the water molecules (hydrogen).



Water is a polar molecule, meaning it has a partial positive and negative side. These partial charges allow water to dissolve other substances with charges.

Covalent Compounds in Solution

Some other covalent compounds, aside from water, are also polar. Having these partial positive and negative charges within the molecule gives polar compounds the ability

to be attracted to water as well. Because of these partial charges, polar molecules are able to dissolve in other polar compounds.

If you mix a nonpolar compound with a polar compound, they will not form an even mixture. The polar compound is more attracted to the other molecules of the same compound than they are attracted to the nonpolar compound. If you have tried to mix oil and water together you may have witnessed this. Water is much more polar than oil, so the oil does not dissolve in the water. Instead, you will see two different layers form.

However, when a nonpolar compound is mixed with another nonpolar compound, neither of them have partial charges to be attracted to. They are instead attracted by London dispersion forces and are able to dissolve together, forming a solution. The similarity in type and strength of intermolecular forces allows two nonpolar compounds such as CO_2 and benzene, C_6H_6.

When we studied how ionic solids dissolve, we said that as they dissolve in solution, these solids separate into ions. More specifically, ionic solids separate into their positive ions and negative ions in solution. This is not true for molecular compounds. Molecular compounds are held together with covalent bonds meaning they share electrons. When they share electrons, their bonds do not easily break apart, thus the molecules stay together even in solution. For example, when you dissolve a spoonful of sugar into a glass of water, the intermolecular forces are broken but not the bonds. You can write the following equation for the dissolution of sugar in water.

$$C_{12}H_{22}O_{11}(s) \rightarrow C_{12}H_{22}O_{11}(aq)$$

Notice how the molecules of sugar are now separated by water molecules (aq). In other words, sugar molecules are separated from neighboring sugar molecules due to attraction for the water, but the molecules themselves have not. The bonds within the molecules have not broken.

Example: Which compounds will dissolve in solution to separate into ions?
(a) LiF
(b) P_2F_5
(c) C_2H_5OH

Solution:
LiF will separate into ions when dissolved in solution, because it is an ionic compound. P_2F_5 and C_2H_5OH are both covalent and will stay as molecules in a solution.

A simple way to predict which compounds will dissolve in other compounds is the phrase "like dissolves like". What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve is water but only to a limited degree. Have you ever wondered why fish are able to breathe? Oxygen gas, a nonpolar molecule, does dissolve in water and it is this oxygen that the fish take in through their gills. Or, one more example of a nonpolar compound that dissolves in water is the reason we can enjoy carbonated sodas. Pepsi-cola and all the other sodas have carbon dioxide gas, CO_2, a nonpolar compound, dissolved in a sugar-water solution. In this case, to keep as much gas in solution as possible, the sodas are kept under pressure.

This general trend of "like dissolves like" is summarized in the following table:

Combination	Solution Formed?
Polar substance in a polar substance.	Yes
Non-polar substance in a non-polar substance.	Yes
Polar substance in a non-polar substance.	No
Non-polar substance in a polar substance.	No
Ionic substance in a polar substance.	Yes
Ionic substance in a non-polar substance.	No

Note that every time charged particles (ionic compounds or polar substances) are mixed, a solution is formed. When particles with no charges (nonpolar compounds) are mixed, they will form a solution. However, if substances with charges are mixed with other substances without charges a solution does not form.

Lesson Summary

- Whether or not solutions are formed depends on the similarity of polarity or the "like dissolves like" rule.
- Polar molecules dissolve in polar solvents, non-polar molecules dissolve in non-polar solvents.
- Ionic compounds dissolve in polar solvents, especially water. This occurs when the positive cation from the ionic solid is attracted to the negative end of the water molecule (oxygen) and the negative anion of the ionic solid is attracted to the positive end of the water molecule (hydrogen).
- Water is considered as the universal solvent since it can dissolve both ionic and polar solutes, as well as some non-polar solutes (in very limited amounts).

Vocabulary

- Miscible: liquids that have the ability to dissolve in each other
- Immiscible: liquids that do not have the ability to dissolve in each other
- electrostatic attraction: the attraction of oppositely charged particles

6.2: Review Questions

1) What does the phrase "like dissolves like" mean? Give an example.
2) Why will LiCl not dissolve in CCl_4?
3) In which compound will you expect benzene, C_6H_6, to dissolve?
   a) Carbon tetrachloride, CCl_4
   b) water
   c) none of the above
4) Thomas is making a salad dressing for supper using balsamic vinegar and oil. He shakes and shakes the mixture but cannot seem to get the two to dissolve. Explain to Thomas why they will not dissolve.

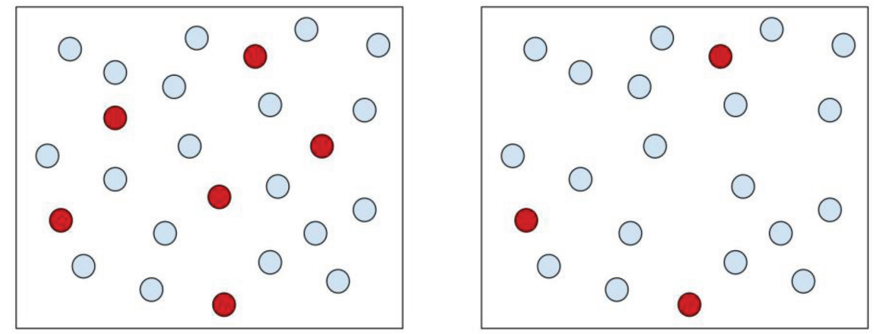
6.3: Concentration

Objectives

- Define the terms "concentrated" and "dilute".

- Define concentration, and list the common units used to express the concentration of solutions.
- Calculate concentration in units of molarity or molality.
- Calculate the amount of solute needed to make a given amount of solution with a given concentration.

Introduction



The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red) to solvent (blue) particles. The solute particles are closer together. The solution on the right is more dilute (less concentrated)
CC – Tracy Poulsen

Concentration is the measure of how much of a given substance is mixed with another substance. Solutions can be said to be dilute or concentrated. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution.

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. What you have to do is take the frozen juice from inside these containers and usually empty 3 or 4 times the container size full of water to mix with the juice concentrate and make your container of juice. Therefore, you are diluting the concentrated juice. When we talk about solute and solvent, the concentrated solution has a lot of solute verses the dilute solution that would have a smaller amount of solute.

The terms "concentrated" and "dilute" provide qualitative methods of describing concentration. Although qualitative observations are necessary and have their place in every part of science, including chemistry, we have seen throughout our study of science that there is a definite need for quantitative measurements in science. This is particularly true in solution chemistry. In this section, we will explore some quantitative methods of expressing solution concentration.

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 used square brackets to indicate a reference to the molarity of a substance. For example, the expression $[Ag^+]$ refers to the molarity of the silver ion. Solution concentrations expressed in molarity are the easiest to calculate with but the most difficult to make in the lab.

$$molarity\ (M) = \frac{mol\ solute}{L\ solution}$$

To solve these problems, we will set them using the factor-label method. To review these steps:

1. Identify the **"given"** information in the problem. Look for a number with units to start this problem with.
2. What is the problem asking you to **"find"**? In other words, what unit will your answer have?
3. Use **ratios** and conversion factors to cancel out the units that aren't part of your answer, and leave you with units that are part of your answer.
4. When your units cancel out correctly, you are ready to do the **math**. You are multiplying fractions, so you multiply the top numbers and divide by the bottom numbers in the fractions.

Example: What is the concentration, in mol/L, where 137 g of NaCl has been dissolved in enough water to make 500. mL of solution?

Solution:
Given: 137 g NaCl, 500. mL solution
Find: $molarity\ (M) = \frac{mol\ solute}{L\ solution}$

$$\frac{137\ \cancel{g\ NaCl}}{500.\ \cancel{mL\ solution}} \cdot \frac{1\ mol\ NaCl}{58.42\ \cancel{g\ NaCl}} \cdot \frac{1000\ \cancel{mL\ solution}}{1\ L\ solution} = \frac{4.69\ mol\ NaCl}{1\ L\ solution} = \mathbf{4.69\ M\ NaCl}$$

Example: What mass of potassium sulfate is in 250. mL of 2.50 M potassium sulfate, K_2SO_4, solution?

Solution:
Given: 250 mL solution
Find: g K_2SO_4
Ratios: 2.50 M or 2.50 mol K_2SO_4/1 L solution

$$250.\ \cancel{mL\ solution} \cdot \frac{1\ \cancel{L\ solution}}{1000\ \cancel{mL\ solution}} \cdot \frac{2.50\ \cancel{mol\ K_2SO_4}}{1\ \cancel{L\ solution}} \cdot \frac{174.3\ g}{1\ \cancel{mol\ K_2SO_4}} = \mathbf{109\ g\ K_2SO_4}$$

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*.

$$molality\ (m) = \frac{mol\ solute}{kg\ solvent}$$

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, HCl, has been dissolved in 115 g of water.

Solution:
Given: 12.5 g HCl, 115 g H_2O
Find: $molality\ (m) = \frac{mol\ HCl}{kg\ H_2O}$

$$\frac{12.5\ \cancel{g\ HCl}}{115\ \cancel{g\ H_2O}} \cdot \frac{1\ mol\ HCl}{36.46\ \cancel{g\ HCl}} \cdot \frac{1000\ \cancel{g\ H_2O}}{1\ kg\ H_2O} = \frac{2.98\ mol\ HCl}{1\ kg\ H_2O} = 2.98\ m\ HCl$$

Although these units of concentration are those which chemists most frequently use, they are not the ones you are most familiar with. Most commercial items you buy at the grocery store have concentrations reported as percentages. For example, hydrogen peroxide you buy is approximately 3% hydrogen peroxide in water; a fruit drink may be 5% real fruit juice. This unit is convenient for these purposes, but not very useful for many chemistry problems. Molarity and molality are preferred because these units involve moles, or how many solute particles there are in a given amount of solution. This comes in handy when performing calculations involving reactions between solutions.

Another common unit of concentration is parts per million (ppm) or parts per billion (ppb). If you have ever looked at the annual water quality report for your area, contaminants in water are typically reported in these units. These units are very useful for concentrations that are really low. A concentration of 1 ppm says that there is 1 gram of the solute for every million grams of the mixture. Because we will not deal with concentrations this low throughout most of this course, we will not use this unit in our calculations. However, you should be aware of it and understand it when you see it.

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.

Vocabulary

- Concentration: the measure of how much of a given substance is mixed with another substance
- Concentrated: a solution in which there is a large amount of solute in a given amount of solvent
- Dilute: a solution in which there is a small amount of solute in a given amount of solvent
- Molarity: the number of moles of solute per liter of solution
- Molality: the number of moles of solute per kilograms of solution

6.3: 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 2.5g of $LiNO_3$ in sufficient water to make 60.0 mL of solution?
3) Calculate the molality of a solution of copper(II) sulfate, $CuSO_4$, where 11.25g of the crystals has been dissolved in 325 g of water.
4) What is the molarity of a solution made by mixing 3.50g of potassium chromate, K_2CrO_4, in enough water to make 100. mL of solution?
5) What is the molarity of a solution made by mixing 50.0g of magnesium nitrate, $Mg(NO_3)_2$, in enough water to make 250. mL of solution?
6) Find the mass of aluminum nitrate, $Al(NO_3)_3$, required to mix with 750g of water to make a 1.5*m* solution.
7) The Dead Sea contains approximately 332 g of salt per kilogram of seawater. Assume this salt is all NaCl. What is the molality of the solution?
8) What is the molarity of a solution prepared by mixing 12.5 grams $FeCl_3$ in enough water to make 300 mL of solution?
9) If 5 grams of NaCl are mixed in enough water to make .5L of solution. What is the molarity of the solution?
10) What is the molality of a solution made by mixing 15 grams of $Ba(OH)_2$ in 250 grams of water?
11) A solution is made by mixing 10.2 grams of $CaCl_2$ in 250 grams of water. What is the molality of the solution?

6.4: Colligative Properties

Objectives

- Explain what the term "colligative" means, and list the colligative properties.
- Indicate what happens to the boiling point and the freezing point of a solvent when a solute is added to it.
- Calculate boiling point elevations and freezing point depressions for a solution

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? As a result of the information you will explore in this section 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

The example given in the introduction is an example of a colligative property. **Colligative properties** are properties that differ based on the concentration of solute in a solvent, but not on the type of solute. What this means for the example above is that people in colder climate don't necessary need salt to get the same effect on the roads – any solute will work. However, the higher the concentration of solute, the more these properties will change.

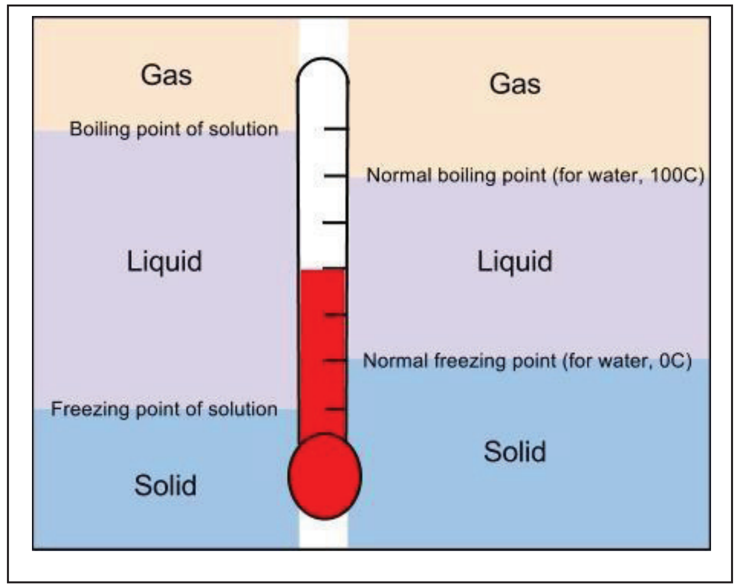
Boiling Point Elevation

Water boils at 100°C at 1 atm of pressure but a solution of salt water does not. When table salt is added to water the resulting solution has a higher boiling point than the water did by itself. The ions form an attraction with the solvent particles that then prevent the water molecules from going into the gas phase. Therefore, the salt-water solution will not boil at 100°C. In order to cause the salt-water solution to boil, the temperature must be raised above 100°C in order to allow the solution to boil. This is true for any solute added to a solvent; the boiling point of the solution will be higher than the boiling point of the pure solvent (without the solute). In other words, when anything is dissolved in water the solution will boil at a higher temperature than pure water would.

The boiling point elevation due to the presence of a solute is also a colligative property. That is, the amount of change in the boiling point is related to number of particles of solute in a solution and is not related to chemical composition of the solute. A 0.20 m solution of table salt and a 0.20 m solution of hydrochloric acid would have the same effect on the boiling point.

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. A solution will have a lower freezing point than a pure solvent. 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. The solute-solvent interactions require the temperature to decrease further in order to solidify the solution. 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°C but at a lower temperature, as low as -9°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** is the difference in the freezing points of the solution from the pure solvent. This is true for any solute added to a solvent; the freezing point of the solution will be lower than the freezing point of the pure solvent (without the solute). In other words, when anything is dissolved in water the solution will freeze at a lower temperature than pure water would.



The freezing point depression elevation due to the presence of a solute is also a colligative property. That is, the amount of change in the freezing point is related to number of particles of solute in a solution and is not related to chemical composition of the solute. A 0.20 m solution of table salt and a 0.20 m solution of hydrochloric acid would have the same effect on the freezing point.

Comparing the Freezing and Boiling Point of Solutions

Recall that covalent and ionic compounds do not dissolve in the same way. Ionic compounds break up into cations and anions when they dissolve. Covalent compounds do not break up. For example a sugar/water solution stays as sugar + water with the sugar molecules staying as molecules. 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. When the solute is ionic, such as NaCl however, adding 10 formulas 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. Colligative properties depend on the number of solute particles in the solution.

"*i*" 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 for NaCl $i = 2$, for lithium nitrate, $LiNO_3$, $i = 2$, and for calcium chloride, $CaCl_2$, $i = 3$. For covalent compounds, *i* is always equal to 1.

By knowing the molality of a solution and the number of particles a compound will dissolve to form, it is possible to predict which solution in a group will have the lowest freezing point.

To compare the boiling or freezing points of solutions, follow these general steps:

1. Label each solute as ionic or covalent.
2. If the solute is ionic, determine the number of *ions* in the formula. Be careful to look for polyatomic ions.
3. Multiply the original molality (*m*) of the solution by the number of particles formed when the solution dissolves. This will give you the total concentration of particles dissolved.
4. Compare these values. The higher total concentration will result in a higher boiling point and a lower freezing point.

Example: Rank the following solutions in water in order of *increasing* (lowest to highest) freezing point:

0.1 *m* NaCl   0.1 *m* $C_6H_{12}O_6$ 0.1 *m* CaI_2

Solution:

To compare freezing points, we need to know the total concentration of all particles when the solute has been dissolved.

- 0.1*m* NaCl: this compound is ionic (metal with nonmetal), and will dissolve into 2 parts. The total final concentration is: $(0.1m)(2) = 0.2m$
- 0.1*m* $C_6H_{12}O_6$: this compound is covalent (nonmetal with nonmetal), and will stay as 1 part. The total final concentration is: $(0.1m)(1) = 0.1m$
- 0.1*m* CaI_2: this compound is ionic (metal with nonmetal), and will dissolve into 3 parts. The total final concentration is: $(0.1m)(3) = 0.3m$

Remember, the greater the concentration of particles, the lower the freezing point will be.

0.1*m* CaI_2 will have the lowest freezing point, followed by 0.1*m* NaCl, and the highest of the three solutions 0.1*m* $C_6H_{12}O_6$, but all of them will have a lower freezing point than pure water.

The Mathematics of Boiling Point and Freezing Point Changes

The boiling point of a solution is higher than the boiling point of a pure solvent and the freezing point of a solution is lower than the freezing point of a pure solvent. However, the amount to which the boiling point increases or the freezing point decreases depends on the amount solute that is added to the solvent. A mathematical equation is used to calculate the boiling point elevation or the freezing point depression.

The boiling point elevation is the amount the boiling temperature *increases* compared to the original solvent. For example, the boiling point of pure water at 1.0 atm is 100°C while the boiling point of a 2% salt-water solution is about 102°C. Therefore, the boiling point elevation would be 2°C. The freezing point depression is amount the freezing temperature *decreases*.

Both the boiling point elevation and the freezing point depression are related to the molality of the solutions. Looking at the formulas for the boiling point elevation and freezing point depression, we can see similarities between the two. The equation used to calculate the increase in the boiling point is:

$$\Delta T_b = k_b \cdot m \cdot i$$

Where:

- ΔT_b = the amount the boiling temperature increased
- k_b = the boiling point elevation constant which depends on the solvent (for water, this number is 0.515ºC/m)
- *m* = the molality of the solution
- *i* = the number of particles formed when that compound dissolves. (for covalent compounds, this number is always 1)

The following equation is used to calculate the decrease in the freezing point:

$$\Delta T_f = k_f \cdot m \cdot i$$

Where:

- ΔT_f = the amount the freezing temperature decreased
- k_f = the freezing point depression constant which depends on the solvent (for water, this number is 1.86ºC/m)
- *m* = the molality of the solution
- *i* = the number of particles formed when that compound dissolves. (for covalent compounds, this number is always 1)

Example: Antifreeze is used in automobile radiators to keep the coolant from freezing. In geographical areas where winter temperatures go below the freezing point of water, using pure water as the coolant could allow the water to freeze. Since water expands when it freezes, freezing coolant could crack engine blocks, radiators, and coolant lines. The main component in antifreeze is ethylene glycol, $C_2H_4(OH)_2$. What is the concentration of ethylene glycol in a solution of water, in molality, if the freezing point dropped by 2.64°C? The freezing point constant, k_f for water is -1.86°C/*m*.

Solution:

$\Delta T_f = k_f \cdot m \cdot i$

Substituting in the appropriate values we get:

$2.64°C = (1.86°C/m)(m)(1)$

Solve for *m* by dividing both sides by 1.86°C/*m*.

***m*=1.42**

Example: A solution of 10.0g of sodium chloride is added to 100.0g of water in an attempt to elevate the boiling point. What is the boiling point of the solution? k_b for water is 0.52°C/*m*.

Solution:

$\Delta T_b = k_b \cdot m \cdot i$

W need to be able to substitute each variable into this equation.

- k_b=0.52°C/*m*.
- *m*: We must solve for this using stoichiometry. Given: 10.0 g NaCl and 100. g H_2O. Find: mol NaCl/kg H_2O. Ratios: molar mass of NaCl, 1000 g = 1 kg

$$\frac{10.0\ \cancel{g\ NaCl}}{100.\ \cancel{g\ H_2O}} \cdot \frac{1\ mol\ NaCl}{58.45\ \cancel{g\ NaCl}} \cdot \frac{1000.\ \cancel{g\ H_2O}}{1\ kg H_2O} = 1.71\ m$$

- For NaCl, i = 2

Substitute these values into the equation $\Delta T_b = k_b \cdot m \cdot i$. We get:

$$\Delta T_b = \left(0.52°\frac{C}{\cancel{m}}\right)(1.71\cancel{m})(2) = 1.78°C$$

Water normally boils at 100°C, but our calculation shows that the boiling point increased by 1.78°C. Our new boiling point is **101.78°C**.

Lesson Summary

- Colligative properties are properties that are due only to the number of particles in solution and not related to the chemical properties of the solute.
- Boiling points of solutions are higher that the boiling points of the pure solvents.
- Freezing points of solutions are lower than the freezing points of the pure solvents.
- Ionic compounds split into ions when they dissolve, forming more particles. Covalent compounds stay as complete molecules when they dissolve.

Vocabulary

- Colligative property: a property that is due only to the number of particles in solution and not the type of the solute
- Boiling point elevation: the amount the boiling point of a solution increases from the boiling point of a pure solvent
- Freezing point depression: the amount the freezing point of a solution decreases from the boiling point of a pure solvent

6.4: Review Questions

1) Which of the following statements are true when a solute is added to a solvent: (you may choose more than 1)
   a) the boiling point increases.
   b) the boiling point decreases.
   c) the freezing point increases.

   d) the freezing point decreases.
2) Why do we put salt on ice on the roads in the winter? What effect does it have on the ice? (Do NOT say that it melts the ice. What does it REALLY do?)
3) Besides adding flavor, what effect does adding salt to water that you cook spaghetti in?
4) How do covalent and ionic compounds differ in how they dissolve? How does this change the molality of the *particles* in the solution?

Label each of the following compounds as ionic or covalent. Then indicate the number of particles formed when dissolved (i) for each compound.

5) Salt, NaCl
6) Acetone, C_2H_2O
7) Benzene, C_6H_6
8) Copper(II) nitrate, $Cu(NO_3)_2$
9) $AlCl_3$
10) Potassium hydroxide, KOH

For each pair of solutions, indicate which would have a lower freezing point:

11) 0.2 m KI or 0.2 m $CaCl_2$
12) 0.1 m KI or 0.1 m $C_6H_{12}O_6$
13) 0.2 m NaCl or 0.3 m $C_6H_{12}O_6$

14) If 25.0g of sucrose, $C_{12}H_{22}O_{11}$, is added to 500.g of water, the boiling point is increased by what amount? (K_b for water is 0.52°C/*m*)

15) For a sample of seawater (an aqueous solution of NaCl), the concentration of salt is approximately 0.50*m*. Calculate the freezing point of seawater. (K_f for water is 1.86°C/m)

16) Calcium chloride is known to melt ice faster than sodium chloride but is not used on roads, because the salt itself attracts water. If 15g of $CaCl_2$ was added to 250g of water, what would be the new freezing point of the solution? K_f for water is 1.86°C/m

Chapter 7: Describing Chemical Reactions

7.1: Chemical & Physical Change

Objectives:

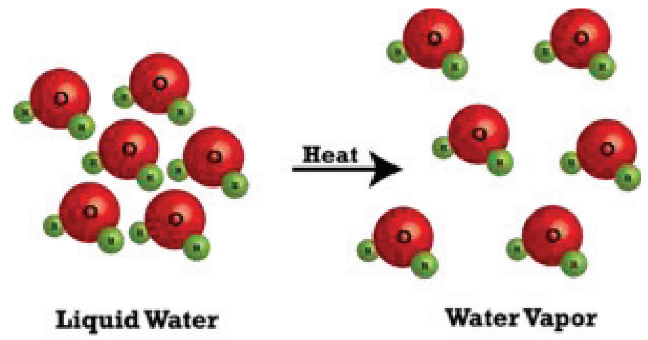
- Label a change as chemical or physical
- List evidence that can indicate a chemical change occurred

Introduction

Change is happening all around us all of the time. Just as chemists have classified elements and compounds, they have also classified types of changes. Changes are either classified as physical or chemical changes.

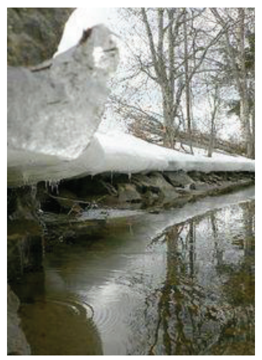
Physical & Chemical Change

Chemists learn a lot about the nature of matter by studying the changes that matter can undergo. Chemists make a distinction between two different types of changes that they study – physical changes and chemical changes. **Physical changes** are changes in which no bonds are broken or formed. This means that the same types of compounds or elements that were there at the beginning of the change are there at the end of the change. Because the ending materials are the same as the beginning materials, the properties (such as color, boiling point, etc) will also be the same. Physical changes involve moving molecules around, but not changing them. Some types of physical changes include:



Water changing from a liquid to gas (boiling) is a physical change, because H_2O molecules are present at the beginning and end of the change.

- Changes of state (changes from a solid to a liquid or a gas and vice versa)
- Separation of a mixture
- Physical deformation (cutting, denting, stretching)
- Making solutions (special kinds of mixtures)

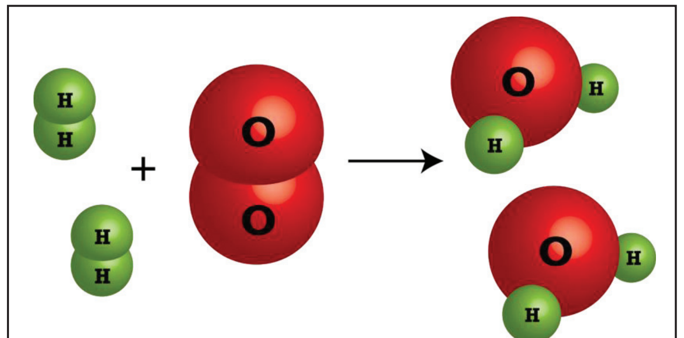


Melting snow is an example of a physical change.

When we heat the liquid water, it changes to water vapor. But even though the physical properties have changed, the molecules are exactly the same as before. We still have each water molecule containing two hydrogen atoms and one oxygen atom covalently bonded. When you have a jar containing a mixture of pennies and nickels and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of either the pennies or the nickels – you've merely separated them into two groups. This would be an example of a physical change. Similarly, if you have a piece of paper, you don't change it into something other than a piece of paper by ripping it up. What was paper before you starting tearing is still paper when you're done. Again, this is an example of a physical change.

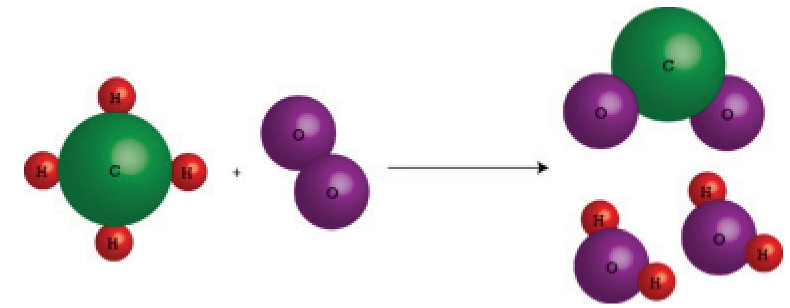
For the most part, physical changes tend to be reversible – in other words, they can occur in both directions. You can turn liquid water into solid water through cooling; you can also turn solid water into liquid water through heating. However, as we will later learn, some chemical changes can also be reversed.

Chemical changes occur when bonds are broken and/or formed between molecules or atoms. This means that one substance with a certain set of properties (such as melting point, color, taste, etc) is turned into a different substance with difference properties. Chemical changes are frequently harder to reverse than physical changes.



Water forming from the elements hydrogen and oxygen is a chemical change, because different substances are present at the end than the beginning.

One good example of a chemical change is burning paper. In contrast to the act of ripping paper, the act of burning paper actually results in the formation of new chemicals (carbon dioxide and water, to be exact). Another example of chemical change occurs when water is formed. Each molecule contains two atoms of hydrogen and one atom of oxygen chemically bonded.



Natural gas burning is an example of a chemical change, because bonds are broken and formed to make different molecules.

Firework displays are an example of a chemical change.

Another example of a chemical change is what occurs when natural gas is burned in your furnace. This time, on the left we have a molecule of methane, CH_4, and two molecules of oxygen, O_2, while on the right we have two molecules of water, H_2O, and one molecule of carbon dioxide, CO_2. In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.

Evidence of Chemical Change

We can't actually see molecules breaking and forming bonds, although that's what defines chemical changes. We have to make other observations to indicate that a chemical change has happened. Some of the evidence for chemical change will involve the energy changes that occur in chemical changes, but some evidence involves the fact that new substances with different properties are formed in a chemical change.

Observations that help to indicate chemical change include:

- Temperature changes (either the temperature increases or decreases)

- Light is given off
- Unexpected color changes (a substance with a different color is made, rather than just mixing the original colors together)
- Bubbles are formed (but the substance is not boiling – you made a substance that is a gas at the temperature as the beginning materials, instead of a liquid)
- Different smell or taste (do not taste your chemistry experiments, though!)
- A solid forms if two clear liquids are mixed (look for *floaties* – technically called a precipitate)

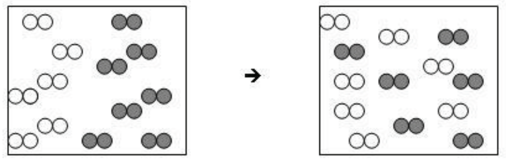
Example: Label each of the following changes as a physical or chemical change. Given evidence to support your answer.
a) boiling water
b) a nail rusts
c) a green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
d) two colorless solutions are mixed. The resulting mixture has a yellow precipitate

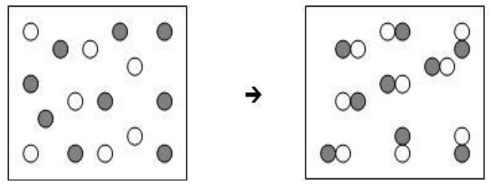
Solution:
a) physical: boiling and melting are physical changes. When water boils no bonds are broken or formed. The change could be written: $H_2O(l) \rightarrow H_2O(g)$
b) chemical: because the dark gray metal nail changes color to form an orange flakey substance (the rust) this must be a chemical change. Color changes indicate color change. The following reaction occurs: $Fe + O_2 \rightarrow Fe_2O_3$
c) physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
d) chemical: the formation of a precipitate and the color change from colorless to yellow indicates a chemical change.

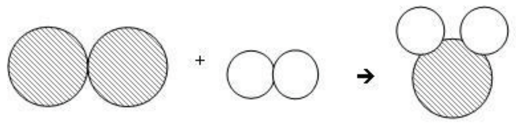
Lesson Summary
- Chemists make a distinction between two different types of changes that they study – physical changes and chemical changes.
- Physical changes are changes that do not alter the identity of a substance
- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes.
- Observations that indicate a chemical change occurred include color change, temperature change, light given off, formation of bubbles, formation of a precipitate, etc

Vocabulary
- Physical changes: changes that do not alter the identity of a substance, the same types of molecules are present at the beginning and end of the change.
- Chemical changes: changes that occur when one substance is turned into another substance; different types of molecules are present at the beginning and end of the change.

7.1: Review Questions
Label each of the following as a physical or chemical change.
1) Water boils at 100°C.
2) Water is separated by electrolysis (running electricity through it) into hydrogen gas and oxygen gas.
3) Sugar dissolves in water.
4) Vinegar and baking soda react to produce a gas.
5) Yeast acts on sugar to form carbon dioxide and ethanol.
6) Wood burns producing several new substances.
7) A cake is baked.

8) 

9) 

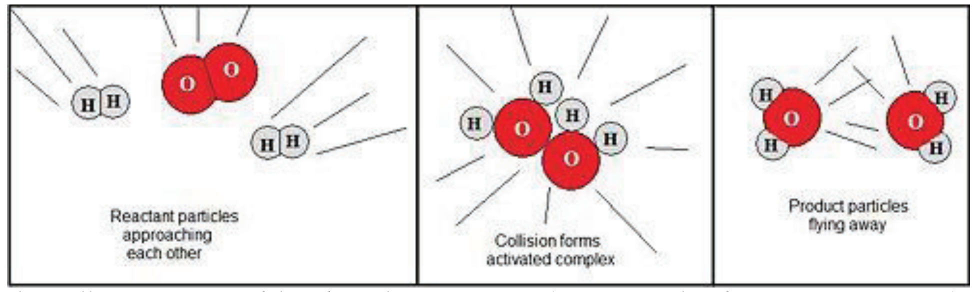
10) 

7.2: Reaction Rate
Objectives
- Describe the conditions for successful collisions that cause reactions
- Describe the rate in terms of the conditions of successful collisions.
- Describe how changing the temperature, concentration of a reactant, or surface area of a reaction affects the rate of a reaction
- Define a catalyst and how a catalyst affects the rate of a reaction

Introduction
We know that a chemical system can be made up of atoms (H_2, N_2, K, etc), ions (NO_3^-, Cl^-, Na^+, etc), or molecules (H_2O, $C_{12}H_{22}O_{11}$, etc). We also know that in a chemical system, these particles are moving around in a random motion. The **collision theory** explains why reactions occur at this particle level between these atoms, ions, and/or molecules. It also explains how it is possible to speed up or slow down reactions that are occurring.

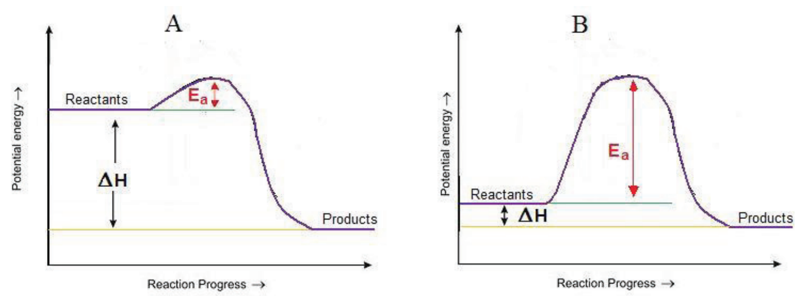
Collision Theory
The collision theory provides us with the ability to predict what conditions are necessary for a successful reaction to take place. These conditions include:
1. The particles must collide with each other.
2. The particles must collide with sufficient energy to break the old bonds.
3. The particles must have proper orientation.

A chemical reaction involves breaking bonds in the reactants, re-arranging the atoms into new groupings (the products), and the formation of new bonds in the products.



This collision is successful and results in reaction. (Source: Richard Parsons. CC-BY-SA)

Therefore, not only must a collision occur between reactant particles, but the collision has to have sufficient energy to break all the reactant bonds that need to be broken in order to form the products. Some reactions need less collision energy than others. The amount of energy the reactant particles must have in order to break the old bonds for a reaction to occur is called the **activation energy**, abbreviated E_a. Another way to think of this is to look at an energy diagram, as shown in the figure. Particles must be able to get over the "bump", the activation energy, if they are going to react. If the reactant particles collide with less than the activation energy, the particles will rebound (bounce off each other), and no reaction will occur.



Each reaction has its own activation energy, E_a. The smaller the "bump", the less energy particles must have to react.

Reaction Rate

Chemists use reactions to generate a product for which they have use. For the most part, the reactions that produce some desired compound are only useful if the reaction occurs at a reasonable rate. For example, using a reaction to produce brake fluid would not be useful if the reaction required 8,000 years complete the product. Such a reaction would also not be useful if the reaction was so fast that it was explosive. For these reasons, chemists wish to be able to control reaction rates. In some cases, chemists wish to speed up reactions that are too slow and slow down reactions that are too fast. In order to gain any control over reaction rates, we must know the factors that affect reaction rates. Chemists have identified many factors that affect the rate of a reaction.

The rate, or speed, at which a reaction occurs depends on the frequency of successful collisions. Remember, a successful collision occurs when two reactants collide with enough energy and with the right orientation. That means if we can do things that will increase the number of collisions, increase the number of particles that have enough energy to react and/or increase the number of particles with the correct orientation we will increase the rate of a reaction.

Effect of Temperature on Rate of Reaction

The rate of reaction was discussed in terms of three factors: collision frequency, the collision energy, and the geometric orientation. Remember that the collision frequency is the number of collisions per second. The collision frequency is dependent, among other factors, on the temperature of the reaction.

When the temperature is increased, the average velocity of the particles is increased. 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. However, 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. In other words, more particles will have the necessary activation energy.

At room temperature, the hydrogen and oxygen in the atmosphere do not have sufficient energy to attain the activation energy needed to produce water:

$$O_2(g) + H_2(g) \rightarrow \text{No reaction}$$

At any one moment in the atmosphere, there are many collisions occurring between these two reactants. But what we find is that water is not formed from the oxygen and hydrogen molecules colliding in the atmosphere because the activation energy barrier is just too high and all the collisions are resulting in rebound. When we increase the temperature of the reactants or give them energy in some other way, the molecules have the necessary activation energy and are able to react to produce water:

$$O_2(g) + H_2(g) \rightarrow H_2O(l)$$

There are times when the rate of a reaction needs to be slowed down. Lowering the temperature could also be used to decrease the number of collisions that would occur and lowering the temperature would also reduce the kinetic energy available for activation energy. If the particles have insufficient activation energy, the collisions will result in rebound rather than reaction. Using this idea, when the rate of a reaction needs to be lower, keeping the particles from having sufficient activation energy will definitely keep the reaction at a lower rate.

Society uses the effect of temperature on rate every day. Food storage is a prime example of how the temperature effect on reaction rate is utilized by society. Consumers store food in freezers and refrigerators to slow down the processes that cause it to spoil. The decrease in temperature decreases the rate at which the food will break down or be broken down by bacteria. In the early years of the 20th century, explorers were fascinated with trying to be the first one to reach the South Pole. In order to attempt such a difficult task at a time without most of the technology we take for granted today, they devised a variety of ways of surviving. One method was to store their food in the snow to be used later during their advances to the pole. On some explorations, they buried so much food, that they didn't need to use all of it and it was left. Many years later, when this food was located and thawed, it was found to still be edible.

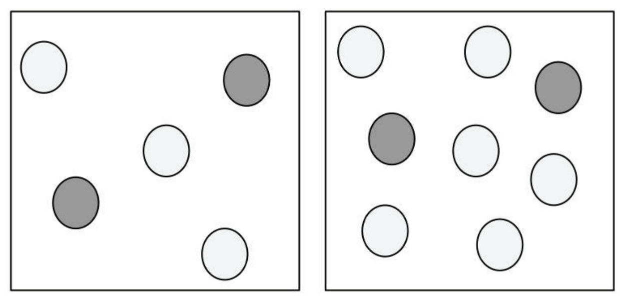
When milk, for instance, is stored in the refrigerator, the molecules in the milk have less energy. This means that while molecules will still collide with other molecules, few of them will react (which means in this case "spoil") because the molecules do not have sufficient energy to overcome the activation energy barrier. The molecules do have energy and are colliding, however, and so, over time, even in the refrigerator, the milk will spoil. Eventually the higher energy molecules will gain the energy needed to react and when enough of these reactions occur, the milk becomes “soured”.

However, if that same carton of milk was at room temperature, the milk would react (in other words “spoil’) much more quickly. Now most of the molecules will have sufficient energy to overcome the energy barrier and at room temperature many more collisions will be occurring. This allows for the milk to spoil in a fairly short amount of time. This is also the reason why most fruits and vegetables ripen in the summer when the temperature is much warmer. You may have experienced this first hand if you have ever bitten into an unripe banana – it was probably sour tasting and might even have felt like biting into a piece of wood! When a banana ripens, numerous reactions occur that produce all the compounds that we expect to taste in a banana. But this can only happen if the temperature is high enough to allow these reactions to make those products.

Effect of Concentration on Rate of Reaction

If you had an enclosed space, like a classroom, and there was one red ball and one green ball flying around the room with random motion 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 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. If one concentration of one of the reactants is doubled, the number of collisions will also double. Assuming that the percent of collision that are successful 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 over time and increasing the concentration of either reactant increases the number of collisions and therefore, increases the number of successful collisions and the reaction rate.



The reaction mixture on the left is less concentrated, so particles will not collide as often. The reaction will be slower.
Because the reacting particles on the right image have a greater change of colliding, the reaction will go faster.
CC – Tracy Poulsen

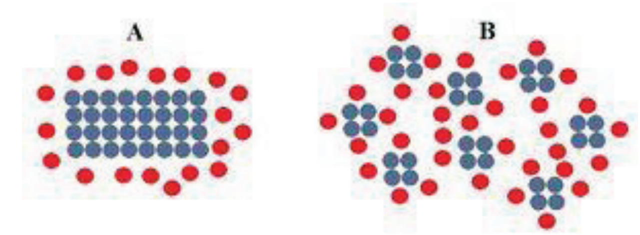
For example, the chemical test used to identify a gas as oxygen or not relies on the fact that increasing the concentration of a reactant increases reaction rate. The reaction we call combustion refers to a reaction in which a flammable substance reacts with oxygen. If we light a wooden splint (a thin splinter of wood) on fire and then blow the fire out, the splint

will continue to glow in air for a period of time. If we insert that glowing splint into any gas that does not contain oxygen, the splint will immediately cease to glow - that is the reaction stops. Oxygen is the only gas that will support combustion. Air is approximately, 20% oxygen gas. If we take that glowing splint and insert it into pure oxygen gas, the reaction will increase its rate by a factor of five - since pure oxygen has 5 times the concentration of oxygen that is in air. When the reaction occurring on the glowing splint increases its rate by a factor of five, the glowing splint will suddenly burst back into full flame. This test, of thrusting a glowing splint into a gas, is used to identify the gas as oxygen. Only a greater concentration of oxygen than that found in air will cause the glowing splint to burst into flame.

Effect of Surface Area on Rate of Reaction

The very first requirement for a reaction to occur between reactant particles is that the particles must collide with each other. The previous section pointed out how increasing the concentration of the reactants increases reaction rate because it increased the frequency of collisions between reactant particles. It can be shown that the number of collisions that occur between reactant particles is also dependent on the surface area of solid reactants. Consider a reaction between reactant RED and reactant BLUE in which reactant blue is in the form of a single lump. Then compare this to the same reaction where reactant blue has been broken up into many smaller pieces.

In the diagram, 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. In Figure A, 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 A, there are a number of blue particles on the interior of the lump that cannot be struck. 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. In diagram B, more collisions between blue and red will occur, and therefore, 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.



In these figures, only the particles on the outside of the solid blue reactant have a chance to collide with the red reactant. In figure B, the same amount of solid reactant as used in A was crushed into smaller particles. This means that more particles on the outside of the reactant have an opportunity to collide with the red reactant and speeds up the reaction.

Several smaller particles have more surface area than one large particle. The more surface area that is available for particles to collide, the faster the reaction will occur. You can see an example of this in everyday life if you have ever tried to start a fire in the fireplace. If you hold a match up against a large log in an attempt to start the log burning, you will find it to be an unsuccessful effort. Holding a match against a large log will not cause enough reactions to occur in order to keep the fire going by providing sufficient activation energy for further reactions. In order to start a wood fire, it is common to break a log up into many small, thin sticks called kindling. These thinner sticks of wood provide many times the surface area of a single log. The match will successfully cause enough reactions in the kindling so that sufficient heat is given off to provide activation energy for further reactions.

There have been, unfortunately, cases where serious accidents were caused by the failure to understand the relationship between surface area and reaction rate. One such example occurred in flour mills. A grain of wheat is not very flammable. It takes a significant effort to get a grain of wheat to burn. If the grain of wheat, however, is pulverized and scattered through the air, only a spark is necessary to cause an explosion. When the wheat is ground to make flour, it is pulverized into a fine powder and some of the powder gets scattered around in the air. A small spark then, is sufficient to start a very rapid reaction which can destroy the entire flour mill. In a 10-year period from 1988 to 1998, there were 129 grain dust explosions in mills in the United States. Efforts are now made in flour mills to have huge fans circulate the air in the mill through filters to remove the majority of the flour dust particles.

Another example is in the operation of coal mines. Coal, of course, will burn but it takes an effort to get the coal started and once it is burning, it burns slowly because only the surface particles are available to collide with oxygen particles. The interior particles of coal have to wait until the outer surface of the coal lump burns off before they can collide with oxygen. In coal mines, huge blocks of coal must be broken up before the coal can be brought out of the mine. In the process of breaking up the huge blocks of coal, drills are used to drill into the walls of coal. This drilling produces fine coal dust that mixes into the air and then a spark from a tool can cause a massive explosion in the mine. There are explosions in coal mines for other reasons but coal dust explosions contributed to the death of many miners. In modern coal mines, lawn sprinklers are used to spray water through the air in the mine and this reduces the coal dust in the air and eliminates coal dust explosions.

Effect of a Catalyst on Rate of Reaction

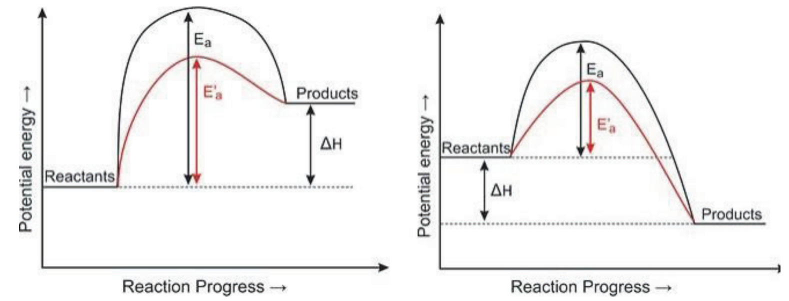
The final factor that affects the rate of the reaction is the effect of the catalyst. A **catalyst** is a substance that speeds up the rate of the reaction without itself being consumed by the reaction.

In the reaction of potassium chlorate breaking down to potassium chloride and oxygen, a catalyst is available to make this reaction occur much faster than it would occur by itself under room conditions. The reaction is:

$$2KClO_3(s) \xrightarrow{MnO_2(s)} 2KCl(s) + 3O_2(g)$$

The catalyst is manganese dioxide and its presence causes the reaction shown above to run many times faster than it occurs without the catalyst. When the reaction has reached completion, the MnO_2 can be removed from the reaction vessel and its condition is exactly the same as it was before the reaction. This is part of the definition of a catalyst . . . that it is not consumed by the reaction. You should note that the catalyst is not written into the equation as a reactant or product but is noted above the yields arrow. This is standard notation for the use of a catalyst.

Some reactions occur very slowly without the presence of a catalyst. In other words the activation energy for these reactions is very high. When the catalyst is added, the activation energy is lowered because the catalyst provides a new reaction pathway with lower activation energy.



A catalyst speeds up a reaction by lowering the activation energy, Ea,. Because less energy is required to react, more particles have the necessary energy.

In the figure on the left, the endothermic reaction shows the catalyst reaction in red with the lower activation energy, designated E'_a. The new reaction pathway has lower activation energy but has no effect on the energy of the reactants, the products, or the value of ΔH. The same is true for the exothermic reaction in Right Figure. The activation energy of the catalyzed reaction is lower than that of the uncatalyzed reaction. The new reaction pathway provided by the catalyst affects the energy required for reactant bonds to break and product bonds to form.

While many reactions in the laboratory can be increased by increasing the temperature, that is not possible for all the reactions that occur in our bodies throughout our entire lives. In fact, the body needs to be maintained at a very specific temperature: 98.6°F or 37°C. Of course there are times, for instance, when the body is fighting an infection, when the body temperature may be increased. But generally, in a healthy person, the temperature is quite consistent. However, many of the reactions that a healthy body depends on could never occur at body temperature. The answer to this dilemma is catalysts or what are also referred to as enzymes. Many of these enzymes are made in your cells since your DNA carries the directions to make them. However, there are some enzymes that your body must have but are not made in your cells. These catalysts must be supplied to your body in the food you eat and are called vitamins.

Lesson Summary

- The collision theory explains why reactions occur between atoms, ions, and/or molecules
- In order for a reaction to be effective, particles must collide with enough energy and having the correct orientation.
- With an increase in temperature, there is an increase in energy that can be converted into activation energy in a collision and therefore there will be an increase in the reaction rate. A decrease in temperature would have the opposite effect.
- With an increase in temperature there is an increase in the number of collisions.
- Increasing the concentration of a reactant increases the frequency of collisions between reactants and will, therefore, increase the reaction rate.
- Increasing the surface area of a reactant (by breaking a solid reactant into smaller particles) increases the number of particles available for collision and will increase the number of collisions between reactants per unit time.
- The catalyst is a substance that speeds up the rate of the reaction without itself being consumed by the reaction. When the catalyst is added, the activation energy is lowered because the catalyst provides a new reaction pathway with lower activation energy.

Vocabulary

- Catalyst: A substance that speeds up the rate of the reaction without itself being consumed by the reaction
- Surface area to volume ratio: The comparison of the volume inside a solid to the area exposed on the surface.

Further Readings / Supplemental Links

- Activation Energy: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/activa2.swf
- 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 ***Molecules in Action***.
- http://www.vitamins-guide.net
- http://en.wikipedia.org/wiki
- Observing molecules during chemical reactions helps explain the role of catalysts. Dynamic equilibrium is also demonstrated. Molecules in Action (http://www.learner.org/vod/vod_window.html?pid=806)
- Surface science examines how surfaces react with each other at the molecular level. On the Surface (http://www.learner.org/vod/vod_window.html?pid=812)
- http://en.wikipedia.org/wiki

7.2: Review Questions

Multiple Choice

1) According to the collision theory, what must happen in order for a reaction to be successful? (Select all that apply.)
   a) particles must collide
   b) particles must have proper geometric orientation
   c) particles must have collisions with enough energy
2) What would happen in a collision between two particles if particles did not have enough energy or had the incorrect 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
3) Why does higher temperature increase the reaction rate?
   a) more of the reacting molecules will have higher kinetic energy
   b) increasing the temperature causes the reactant molecules to heat up
   c) the activation energy will decrease
4) When the temperature is increased, what does not change?
   a) number of collisions
   b) activation energy
   c) number of successful collisions
   d) all of the above change
5) Why is the increase in concentration directly proportional to the rate of the reaction?

   a) The kinetic energy increases.
   b) The activation energy increases.
   c) The number of successful collisions increases.
   d) All of the above.

Choose the substance with the greatest surface in the following groupings:

6) a block of ice or crushed ice
7) sugar crystals or sugar cubes
8) a piece of wood or wood shavings

9) Why, using the collision theory, do reactions with higher surface area have faster reaction rates?

Limestone (calcium carbonate) reacts with hydrochloric acid in an irreversible reaction, to form carbon dioxide and water as described by the following equation:

$$CaCO_{3(s)} + 2\,HCl_{(aq)} \rightarrow CaCl_{2(aq)} + CO_{2(g)} + H_2O_{(l)}$$

What is the effect on the rate if:

10) The temperature is lowered?
11) Limestone chips are used instead of a block of limestone?
12) A more dilute solution of HCl is used?

7.3: Chemical Reactions and Equations

Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, (s), (l), (g), (aq), and →, appropriately when writing a chemical reaction
- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Introduction

In a chemical change, new substances are formed. In order for this to occur, the chemical bonds of the substances break, and the atoms that compose them separate and re-arrange themselves into new substances with new chemical bonds. When this process occurs, we call it a chemical reaction. A **chemical reaction** is the process in which one or more substances are changed into one or more new substances.

Reactants and Products

In order to describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

Sometimes when reactants are put into a reaction vessel, a reaction will take place to produce products. Reactants are the starting materials, that is, whatever we have as our initial ingredients. The products are just that, what is produced or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, some baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

Flour + Butter + Sugar + Vanilla + Baking Soda + Eggs + Chocolate Chips → Cookies

Writing Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $SO_2 + O_2$, are reactants and sulfur trioxide, SO_3, is the product.

$$2\,SO_{2(g)} + O_{2(g)} \longrightarrow 2\,SO_{3(g)}$$

Reactants → Products

In chemical reactions, the reactants are found before the symbol "→" and the products and found after the symbol "→". The general equation for a reaction is:

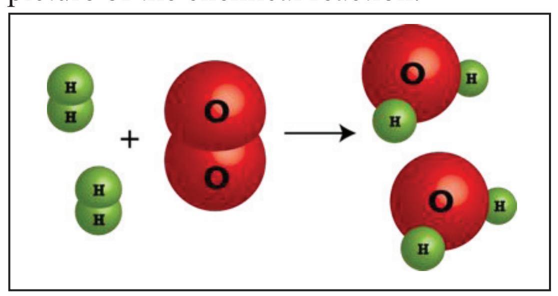
Reactants → Products

There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. You will find there are others but these are the main ones that we need to know.

Common Symbols in Chemical Reactions		
Symbol	**Meaning**	**Example**
→	Used to separate reactants from products; can be read as "to produce" or "yields.	$2\,H_2 + O_2 \rightarrow 2\,H_2O$
+	Used to separate reactants from each other or products from each other; can be read as "is added to" or "also forms".	$AgNO_3 + NaCl \rightarrow AgCl + NaNO_3$
(s)	in the solid state	sodium in the solid state = $Na(s)$
(l)	in the liquid state	water in the liquid state = $H_2O(l)$
(g)	in the gaseous state	carbon dioxide in the gaseous state = $CO_2(g)$
(aq)	in the aqueous state, dissolved in water	sodium chloride solution = $NaCl(aq)$

Chemists have a choice of methods for describing a chemical reaction.

1. They could draw a picture of the chemical reaction.



2. They could write a word equation for the chemical reaction:

"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor. "

3. They could write the equation in chemical shorthand.

$$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$$

In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products and symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

We could write that an aqueous solution of calcium nitrate is added to an aqueous solution of sodium hydroxide to produce solid calcium hydroxide and an aqueous solution of sodium nitrate. Or in shorthand we could write:

$$Ca(NO_3)_2(aq) + 2\,NaOH(aq) \rightarrow Ca(OH)_2(s) + 2\,NaNO_3(aq)$$

How much easier is that to read? Let's try it in reverse? Look at the following reaction in shorthand notation and write the word equation for the reaction.

$$Cu(s) + AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + Ag(s)$$

The word equation for this reaction might read something like "solid copper reacts with an aqueous solution of silver nitrate to produce a solution of copper (II) nitrate with solid silver".

In order to turn word equations into symbolic equations, we need to follow the given steps:

1. Identify the reactants and products. This will help you know what symbols go on each side of the arrow and where the + signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in chapter 4 (including making all ionic compounds charge balanced).
3. Write the correct formulas for all elements. Usually, this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning they are always found in pairs in nature. They include those elements listed in the table.

Diatomic Elements	
Element name	**Formula**
Hydrogen	H_2
Nitrogen	N_2
Oxygen	O_2
Fluorine	F_2
Chlorine	Cl_2
Bromine	Br_2
Iodine	I_2

Example: Transfer the following symbolic equations into word equations or word equations into symbolic equations.

(a) $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

(b) Gaseous propane, C_3H_8, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.

(c) Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution:

(a) An aqueous solution of hydrochloric acid reacts with an aqueous solution of sodium hydroxide to produce an aqueous solution of sodium chloride and liquid water.

(b) Reactants: propane (C_3H_8) and oxygen (O2)

Products: carbon dioxide (CO_2) and water (H_2O)

$$C_3H_8(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$$

(c) Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$HF(g) + K_2CO_3(aq) \rightarrow KF(aq) + H_2O(l) + CO_2(g)$

7.3: Review Questions

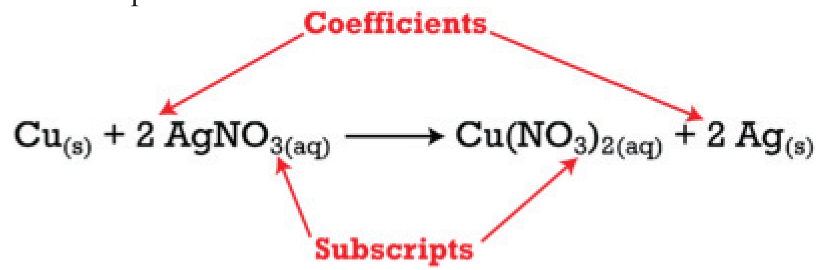
Convert the following equations from word equations into symbolic equations. Be sure to look up charges of ionic compounds to write the correct formula for the compound.

1) Solid calcium metal is placed in liquid water to produce aqueous calcium hydroxide and hydrogen gas.
2) Gaseous sodium hydroxide is mixed with gaseous chlorine to produce aqueous solutions of sodium chloride and sodium hypochlorite plus liquid water.
3) Iron reacts with sulfur when heated to form iron(II) sulfide.
4) When aluminum is added to sulfuric acid (H_2SO_4), the solution reacts to form hydrogen gas and aluminum sulfate.
5) When aluminum is mixed with iron(III) oxide, they react to produce aluminum oxide and iron.
6) Fluorine is mixed with sodium hydroxide to form sodium fluoride, oxygen, and water.
7) A solid chunk of iron is dropped into a solution of copper(I) nitrate forming iron(II) nitrate and solid copper.

7.4: Balancing Chemical Equations

Even though chemical compounds are broken up and new compounds are formed 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 – they 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 complete chemical equation, the same number of each atom must be present on the reactants and the products sides of 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 is used or produced.



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 number of each substance involved in the reaction and may be changed in order to balance the equation. 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 atoms of solid silver.

Once you have written a symbolic equation from words, it is important to balance the equation. It is very important to note that these steps must be carried out *in the correct order*. You must have the correct formulas for your reactants and products **before** you can use coefficients to balance the equation.

When you learned how to write formulas, it was made clear that when only one atom of an element is present, the subscript of "1" is not written - so that when no subscript appears for an atom in a formula, you read that as one atom. The same is true in writing balanced chemical equations. If only one atom or molecule is present, the coefficient of "1" is omitted. Coefficients are inserted into the chemical equation in order to balance it; that is, to make the total number of each atom on the two sides of the equation equal. Equation balancing is accomplished by changing coefficients, never by changing subscripts.

Example: Balance the following skeletal equation. (The term "skeletal equation" refers to an equation that has the correct formulas but has not yet had the proper coefficients added.)

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

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

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

Then we can 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: Write a balanced equation for the reaction that occurs between chlorine gas and aqueous sodium bromide to produce liquid bromine and aqueous sodium chloride.

Solution:

Step 1: Write the word equation (keeping in mind that chlorine and bromine refer to the diatomic molecules).

Chlorine + sodium bromide → bromine + sodium chloride

Step 2: Substitute the correct formulas into the equation.

$Cl_2 + NaBl \rightarrow Br_2 + NaCl$

Step 3: Insert coefficients where necessary to balance the equation.

By placing a coefficient of 2 in front of the NaBr, we can balance the bromine atoms and by placing a coefficient of 2 in front of the NaCl, we can balance the chlorine atoms.

$Cl_2 + 2\ NaBl \rightarrow Br_2 + 2\ NaCl$

A final check (always do this) shows that we have the same number of each atom on the two sides of the equation and we do not have a multiple set of 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. (You may need to refer to a chart of polyatomic ions.)

Solution:

Step 1: Write the word equation.

Aluminum sulfate + calcium bromide → aluminum bromide + calcium sulfate
Step 2: Replace the names of the substances in the word equation with formulas.

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

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.

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

In order to balance the sulfate ions, we must insert a coefficient of 3 in front of the $CaSO_4$ in the products.

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

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

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

The insertion of the 3 in front of the $CaBr_2$ in the reactants also balances the calcium atoms in the $CaSO_4$ in the products. A final check shows 2 aluminum atoms on each side, 3 sulfur atoms on each side, 12 oxygen atoms on each side, 3 calcium atoms on each side, and 6 bromine atoms on each side. This equation is balanced.

Chemical equations should be balanced with the simplest whole number coefficients that balance the equation. Here is the properly balanced equation from the previous section.

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

Note that the equation in the previous section would have the same number of atoms of each type on each side of the equation with the following set of coefficients.

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

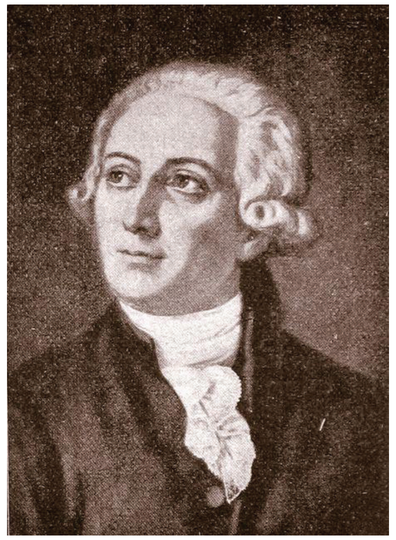
Count the number of each type of atom on each side of the equation to confirm that this equation is "balanced". While this set of coefficients does "balance" the equation, they are not the lowest set of coefficients possible that balance the equation. We could divide each of the coefficients in this equation by 2 and get another set of coefficients that are whole numbers and also balance the equation. Since it is required that an equation be balanced with the lowest whole number coefficients, the last equation 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 each of the following reactions.
(a) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$
(b) $H_2SO_4(aq) + Al(OH)_3(aq) \rightarrow Al_2(SO_4)_3(aq) + H_2O(l)$
(c) $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(aq) + NaNO_3(aq)$
(d) $C_2H_4(g) + O_2 \rightarrow CO_2(g) + H_2O(l)$

Solutions
(a) $CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$ (In this case, the equation balances with all coefficients being 1)
(b) $3\ H_2SO_4(aq) + 2\ Al(OH)_3(aq) \rightarrow Al_2(SO_4)_3(aq) + 6\ H_2O(l)$
(c) $Ba(NO_3)_2(aq) + Na_2CO_3(aq) \rightarrow BaCO_3(aq) + 2\ NaNO_3(aq)$
(d) $C_2H_4(g) + 3\ O_2 \rightarrow 2\ CO_2(g) + 2\ H_2O(l)$

Conservation of Mass in Chemical Reactions

In chapter 2 we discussed the development of the atomic theory, or the idea that everything is made of atoms. A strong piece of evidence for this theory was experimentally determined by Antoine Lavoisier, a French chemist. The **Law of Conservation of Mass**, as he states it, says that mass is conserved in chemical reactions. In other words, the mass of the starting materials (reactants) is always equal to the mass of the ending materials (products).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But what does this really mean? Dalton used this finding to support the idea of atoms. If the mass isn't changing, then the particles that carry the mass (atoms) aren't created or destroyed, but are only rearranged in a chemical reaction. Both the numbers of each type of atom and the mass are conserved during chemical reactions. An examination of a properly balanced equation will demonstrate that mass is conserved. Consider the following reaction.

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

You should check that this equation is balanced by counting the number of each type of atom on each side of the equation.

We can also demonstrate that mass is conserved in this reaction by determining the total mass on the two sides of the equation. We will use the molar masses to add up the masses of the atoms on the reactant side and compare this to the mass of the atoms on the product side of the reaction:

Reactant Side Mass

1 moles of $Fe(NO_3)_3$ x molar mass = (1mol)(241.9 g/mol) = 241.9 g
3 moles of NaOH x molar mass = (3mol)(40.0 g/mol) = 120. g
Total mass for reactants = 241.9 g + 120. g = 361.9 g

Product Side Mass

1 moles of $Fe(OH)_3$ x molar mass= (1mol)(106.9g/mol) = 106.9g
3 moles of $NaNO_3$ x molar mass = (3mol)(85.0 g/mol) = 255 g
Total mass for products = 106.9 g + 255 g = 361.9 g

As you can see, both number of atom types and mass are conserved during chemical reactions. A group of 20 objects stacked in different ways will still have the same total mass no matter how you stack them.

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 is read as "yields," and the products on the right.
- To be useful, chemical equations must always be balanced.

- Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio.
- Mass is always conserved in chemical reactions.

Vocabulary

- Chemical reaction: the process in which one or more substances are changed into one or more new substances
- Reactants: the starting materials in a reaction
- Products: materials present at the end of a reaction
- Balanced chemical equation: a chemical equation in which the number of each type of atom is equal on the two sides of the equation
- Subscripts: part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element
- Coefficient: a small whole number that appears in front of a formula in a balanced chemical equation

Further Reading / Supplemental Links

- For a Bill Nye video on reactions, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "Bill Nye reactions". (you can get the username and password from your teacher)
- For videos and clips on reactions, go to http://www.uen.org/dms/. Go to the k-12 library. Search for "reactions" or "chemical equations". (you can get the username and password from your teacher)
- Vision Learning: Chemical Equations http://visionlearning.com/library/module_viewer.php?mid=56&l=&c3=
- Balancing Equations Tutorial: http://www.mpcfaculty.net/mark_bishop/balancing_equations_tutorial.htm
- Balancing Equations Tutorial: http://www.wfu.edu/~ylwong/balanceeq/balanceq.html
- Law of Conservation of Mass (YouTube): http://www.youtube.com/watch%3Fv%3DdExpJAECSL8

7.4 Review Questions

Balance the following equations.

1) $Cu + O_2 \rightarrow CuO$
2) $H_2O \rightarrow H_2 + O_2$
3) $Fe + H_2O \rightarrow H_2 + Fe_2O_3$
4) $NaCl \rightarrow Na + Cl_2$
5) $AsCl_3 + H_2S \rightarrow As_2S_3 + HCl$
6) $CaCO_3 \rightarrow CaO + CO_2$
7) $H_2S + KOH \rightarrow HOH + K_2S$
8) $XeF_6 + H_2O \rightarrow XeO_3 + HF$
9) $Cu + AgNO_3 \rightarrow Ag + Cu(NO_3)_2$
10) $Fe + O_2 \rightarrow Fe_2O_3$
11) $Al(OH)_3 + Mg_3(PO_4)_2 \rightarrow AlPO_4 + Mg(OH)_2$
12) $Al + H_2SO_4 \rightarrow H_2 + Al_2(SO_4)_3$

13) $H_3PO_4 + NH_4OH \rightarrow HOH + (NH_4)_3PO_4$
14) $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$
15) $Al + O_2 \rightarrow Al_2O_3$
16) $CH_4 + O_2 \rightarrow CO_2 + H_2O$

17) When the following equation is balanced, what is the coefficient found in front of the O_2? $P_4 + O_2 + H_2O \rightarrow H_3PO_4$
18) When properly balanced, what is the **sum** of all the coefficients in the following chemical equation? $SF_4 + H_2O \rightarrow H_2SO_3 + HF$
19) Explain in your own words why it is essential that subscripts remain constant but coefficients can change when balancing a reaction.

7.5: Types of Reactions

Objectives

- Classify a chemical reaction as a synthesis, decomposition, single replacement, double replacement, or a combustion reaction.
- Predict the products of simple reactions.

Introduction

Chemical reactions are classified into types to help us analyze them and also to help us predict what the products of the reaction will be. The five major types of chemical reactions are synthesis, decomposition, single replacement, double replacement, and combustion.

Synthesis Reactions

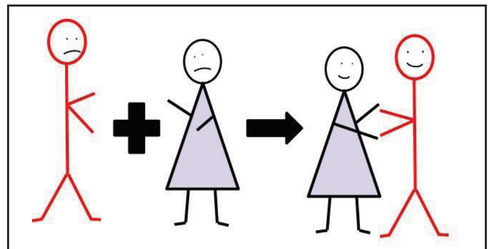
A synthesis reaction is one in which two or more reactants combine to make one type of product.

General equation: $A + B \rightarrow AB$

Synthesis reactions occur as a result of two or more simpler elements or molecules combining to form a more complex molecule. Look at the example below. Here two elements (hydrogen and oxygen) are combining to form one product (water).

Example: $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(l)$

We can always identify a synthesis reaction because there is only one product of the reaction.



A synthesis reaction is similar to forming a couple, which behaves and acts differently than the two single individuals.
CC – Tracy Poulsen

You should be able to write the chemical equation for a synthesis reaction if you are given a product by picking out its elements and writing the equation. Also, if you are given elemental reactants and told that the reaction is a synthesis reaction, you should be able to predict the products.

Example:
(a) Write the chemical equation for the synthesis reaction of silver bromide, AgBr.
(b) Predict the products for the following reaction: $CO_2(g) + H_2O(l)$

Solution:
(a) $2\,Ag + Br_2 \rightarrow 2\,AgBr$
(b) $CO_2(g) + H_2O(l) \rightarrow H_2CO_3$

Decomposition Reactions

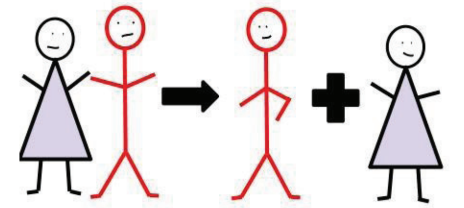
When one type of reactant breaks down to form two or more products, we have a decomposition reaction. The best way to remember a decomposition reaction is that for all reactions of this type, there is only one reactant.

General Equation: $AB \rightarrow A + B$

Look at the example below for the decomposition of ammonium nitrate to dinitrogen oxide and water.

Example: $NH_4NO_3 \rightarrow N_2O + 2\,H_2O$

Notice the one reactant, NH_4NO_3, is on the left of the arrow and there is more than one on the right side of the arrow. This is the exact opposite of the synthesis reaction type.



A decomposition reaction is similar to breaking up a couple, in which the individuals have different properties from the couple they started out in.
CC – Tracy Poulsen

When studying decomposition reactions, we can predict reactants in a similar manner as we did for synthesis reactions. Look at the formula for magnesium nitride, Mg_3N_2. What elements do you see in this formula? You see magnesium and nitrogen. Now we can write a decomposition reaction for magnesium nitride. Notice there is only one reactant.

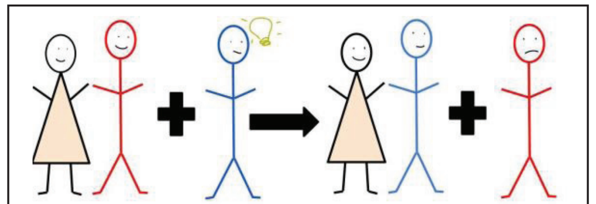
$$Mg_3N_2 \rightarrow 3\,Mg + N_2$$

Example: Write the chemical equation for the decomposition of:
(a) Al_2O_3
(b) Ag_2S
(c) MgO

Solution:
(a) $2\,Al_2O_3 \rightarrow 4\,Al + 3\,O_2$
(b) $Ag_2S \rightarrow 2\,Ag + S$
(c) $2\,MgO \rightarrow 2\,Mg + O_2$

Single Replacement Reactions

A third type of reaction is the single replacement reaction. In single replacement reactions one element reacts with one compound to form products. The single element is said to replace an element in the compound when products form, hence the name single replacement. Metal elements will always replace other metals in ionic compounds or hydrogen in an acid. Nonmetal elements will always replace another nonmetal in an ionic compound.



In a single replacement reaction, a single element (individual) takes the place of an element within a compound (couple), leaving a different element (individual) separate.
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General equation: $A + BC \rightarrow B + AC$

Consider the following examples.

$$Zn(s) + Cu(NO_3)_2(aq) \rightarrow Zn(NO_3)_2(aq) + Cu(s)$$

Notice that the metal element, Zn, replaced the metal in the compound $Cu(NO_3)_2$. A metal element will always replace a metal in an ionic compound. Also, note that the charges of the ionic compounds must equal zero. To correctly predict the formula of the ionic product, you must know the charges of the ions you are combining.

$$Zn(s) + 2\,HBr(aq) \rightarrow ZnCl_2(aq) + H_2(g)$$

When a metal element is mixed with acid, the metal will replace the hydrogen in the acid and release hydrogen gas a product. Once again, note that the charges of the ionic compounds must equal zero. To correctly predict the formula of the ionic product, you must know the charges of the ions you are combining, in this case Zn^{2+} and Cl^-.

$$Cl_2(g) + 2\,KI(aq) \rightarrow 2\,KCl(aq) + I_2(s)$$

When a nonmetal element is added to an ionic compound, the element will replace the nonmetal in the compound. Also, to correctly write the formulas of the products, you must first identify the charges of the ions that will be in the ionic compound.

Example: What would be the products of the reaction between solid aluminum and iron(III) oxide? The reactants are: $Al + Fe_2O_3 \rightarrow$

Solution: In order to predict the products we need to know that aluminum will replace iron and form aluminum oxide (the metal will replace the metal ion in the compound). Aluminum has a charge of +3 and oxygen has a charge of -2. The compound formed between aluminum and oxygen, therefore, will be Al_2O_3. Since iron is replaced in the compound by aluminum, the iron will now be the single element in the products. The unbalanced equation will be:

$$Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$$

and the balanced equation will be:

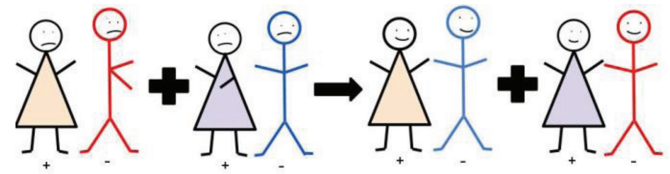
$$2\,Al + Fe_2O_3 \rightarrow Al_2O_3 + 2\,Fe$$

Example:
(a) Write the chemical equation for the single replacement reaction between zinc solid and lead(II) nitrate solution to produce zinc nitrate solution and solid lead. (*Note: zinc forms ions with a +2 charge)
(b) Predict the products for the following reaction: $Fe + CuSO_4$ (in this reaction, assume iron forms ions with a +2 charge)
(c) Predict the products for the following reaction: $Al + CuCl_2$
(d) Complete the following reaction. Then balance the equation: $Al + HNO_3 \rightarrow$

Solution:
(a) $Zn + Pb(NO_3)_2 \rightarrow Pb + Zn(NO_3)_2$
(b) $Fe + CuSO_4 \rightarrow Cu + FeSO4$
(c) $2\,Al + 3\,CuCl_2 \rightarrow 3\,Cu + 2\,AlCl_3$
(d) $2\,Al + 6\,HNO_3 \rightarrow 2\,Al(NO_3)_3 + 3\,H_2$

Double Replacement

For double replacement reactions two ionic compound reactants will react by having the cations exchange places, forming two new ionic compounds. The key to this type of reaction, as far as identifying it over the other types, is that it has two compounds as reactants. This type of reaction is more common than any of the others and there are many different types of double replacement reactions. Precipitation and neutralization reactions are two of the most common double replacement reactions. **Precipitation reactions** are ones where two aqueous compound reactants combine to form products where one of the products is an insoluble solid. A **neutralization reaction** is one where the two reactant compounds are an acid and a base and the two products are a salt and water (i.e. acid + base salt + water).



In a double replacement reaction, the cations (girls) in two compounds (couples) trade places to form two new ionic compounds (couples)
CC – Tracy Poulsen

General equation: $AB + CD \rightarrow AD + CB$

For example, when solutions of silver nitrate and sodium chloride are mixed, the following reaction occurs:

$$AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$$

This is an example of a precipitate reaction. Notice that two aqueous reactants form one solid, the precipitate, and another aqueous product.

An example of a neutralization reaction occurs when sodium hydroxide, a base, is mixed with sulfuric acid:

$$2\ NaOH(aq) + H_2SO_4(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

In order to write the products for a double displacement reaction, you must be able to determine the correct formulas for the new compounds. Remember, the total charge of all ionic compounds is zero. To correctly write the formulas of the products, you must know the charges of the ions in the reactants. Let's practice with an example or two.

Example: A common laboratory experiment involves the reaction between lead(II) nitrate and sodium iodide, both colorless solutions. The reactants are given below. Predict the products.

$$Pb(NO_3)_2(aq) + NaI(aq) \rightarrow$$

Solution:
We know that the cations exchange anions. We now have to look at the charges of each of the cations and anions to see what the products will be. In $Pb(NO_3)_2$, the nitrate, NO_3^- has a charge of -1. This means the lead must be +2, Pb^{2+}. In the sodium iodide, we are combining Na^+ and I^-.

Now we switch ions and write the correct subscripts so the total charge of each compound is zero. The Pb^{2+} will combine with the I^- to form PbI_2. The Na^+ will combine with the NO_3^- to form $NaNO_3$.

Only after we have the correct formulas can we worry about balancing the two sides of the reaction. The final balanced reaction will be:

$$Pb(NO_3)_2(aq) + 2\ NaI(aq) \rightarrow PbI_2(s) + 2\ NaNO_3(aq)$$

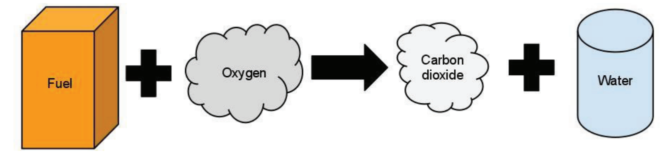
Example:
(a) Write a chemical equation for the double replacement reaction between calcium chloride solution and potassium hydroxide solution to produce potassium chloride solution and a

precipitate of calcium hydroxide.
(b) Predict the products for the following reaction: $AgNO_3(aq) + NaCl(aq) \rightarrow$
(c) Predict the products for the following reaction: $FeCl_3(aq) + KOH(aq) \rightarrow$

Solution:
(a) $CaCl_2(aq) + 2\ KOH(aq) \rightarrow Ca(OH)_2(s) + 2\ KCl(aq)$
(b) $AgNO_3(aq) + NaCl(aq) \rightarrow AgCl(s) + NaNO_3(aq)$
(c) $FeCl_3(aq) + KOH(aq) \rightarrow Fe(OH)_3(s) + KCl(aq)$

Combustion

In a combustion reaction oxygen reacts with another substance to produce carbon dioxide and water. This is what happens when fuel burns. In a particular branch of chemistry, known as organic chemistry, we study compounds known as hydrocarbons. A **hydrocarbon** is compound consisting of only hydrogen and carbon. Hydrocarbons represent the major components of all organic material including fuels. Combustion reactions usually have the same products, CO_2 and H_2O, and one of its reactants is always oxygen. In other words, the only part that changes from one combustion reaction to the next is the actual hydrocarbon that burns. The general equation is given below. Notice the oxygen, carbon dioxide, and water parts of the reaction are listed for you to show you how these reactants and products remain the same from combustion reaction to combustion reaction.



In a combustion reaction, a fuel (C_xH_y compound) reacts with oxygen (O_2) to form CO_2 and H_2O.
CC – Tracy Poulsen

General equation: $C_xH_y\ (hydrocarbon) + O_2 \rightarrow CO_2 + H_2O$

Look at the reaction for the combustion of octane, C_8H_{18}, below. Octane has 8 carbon atoms hence the prefix "oct".

Example: $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$

This reaction is referred to as complete combustion. Complete combustion reactions occur when there is enough oxygen to burn the entire hydrocarbon. This is why there are only carbon dioxide and water as products.

Example: Write the balanced reaction for the complete combustion of propane, C_3H_8.

Solution: The reactants of all combustion reactions include the fuel (a compound with carbon and hydrogen) reacting with oxygen. The products are always carbon dioxide and water.

$$C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$$

Lesson Summary

The Five Types of Chemical Reactions

Reaction Name	Reaction Description
Synthesis:	two or more reactants form one product.
Decomposition:	one type of reactant forms two or more products.
Single replacement:	one element reacts with one compound to form products.

Double replacement:	two compounds act as reactants.
Combustion:	a fuel reactant reacts with oxygen gas.

Vocabulary

- **Synthesis reaction:** a reaction in which two or more reactants combine to make one product
- **Decomposition reaction:** a reaction in which one reactant breaks down to form two or more products
- **Single replacement reaction:** a reaction in which an element reacts with a compound to form products
- **Double replacement reaction:** a reaction in which two reactants form products by having the cations exchange places with the anions
- **Combustion reaction:** a reaction in which oxygen reacts with another substance to produce carbon dioxide and water
- **Hydrocarbon:** an organic substance consisting of only hydrogen and carbon

7.4: Review Questions

Classify each type of reaction as synthesis, decomposition, single replacement, double replacement or combustion.

1) $Cu + O_2 \rightarrow CuO$
2) $H_2O \rightarrow H_2 + O_2$
3) $Fe + H_2O \rightarrow H_2 + Fe_2O_3$
4) $AsCl_3 + H_2S \rightarrow As_2S_3 + HCl$
5) $Fe_2O_3 + H_2 \rightarrow Fe + H_2O$
6) $CaCO_3 \rightarrow CaO + CO_2$
7) $H_2S + KOH \rightarrow HOH + K_2S$
8) $NaCl \rightarrow Na + Cl_2$
9) $Al + H_2SO_4 \rightarrow H_2 + Al_2(SO_4)_3$
10) $CH_4 + O_2 \rightarrow CO_2 + H_2O$

11) Distinguish between synthesis and decomposition reactions.

12) When dodecane, $C_{10}H_{22}$, burns in excess oxygen, what will be the products?

13) When iron rods are placed in liquid water, a reaction occurs. Hydrogen gas evolves from the container and iron(III) oxide forms onto the iron rod. Classify the type of reaction and write a balanced chemical equation for the reaction.

Classify each of the following reactions and predict products for each reaction.

14) $H_3PO_4 + NH_4OH \rightarrow$
15) $C_3H_8 + O_2 \rightarrow$
16) $Al + O_2 \rightarrow$
17) $BaCl_2 + Na_2SO_4 \rightarrow$
18) $Ca + HCl \rightarrow$
19) $FeS + HCl \rightarrow$
20) $NaI + Br_2 \rightarrow$

7.6: Stoichiometry

Objectives

- Explain the meaning of the term "stoichiometry".
- Determine mole ratios in chemical equations.
- Calculate the number of moles of any reactant or product from a balanced equation given the number of moles of one reactant or product.
- Calculate the mass of any reactant or product from a balanced equation given the mass of one reactant or product.

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 quantities of reactants and products in a balanced equation. This is known as stoichiometry.

Stoichiometry, by definition, is the calculation of the quantities of reactants or products in a chemical reaction using the relationships found in the balanced chemical equation. The word stoichiometry is actually Greek coming from two words stoikheion, which means element and metron, which means measure.

Interpreting Chemical Equations

The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Look at the chemical equation below.

$2\ CuSO_4$	+	$4\ KI$	$\rightarrow$	$2\ CuI$	+	$2\ K_2SO_4$	+	I_2
2 formula units $CuSO_4$	+	*4 formula units KI*	$\rightarrow$	*2 formula units CuI*	+	*2 formula units K_2SO_4*	+	*1 molecule I_2*
2 moles $CuSO_4$	+	*4 moles KI*	$\rightarrow$	*2 moles CuI*	+	*2 moles K_2SO_4*	+	*1 moles I_2*

The coefficients used, as we have learned, tell us the relative amounts of each substance in the equation. So for every 2 units of copper (II) sulfate ($CuSO_4$) we have, we need to have 4 units of potassium iodide (KI). For every two dozen copper(II) sulfates, we need 4 dozen potassium iodides. Because the unit "mole" is also a counting unit, we can interpret this equation in terms of moles, as well: For every two moles of copper(II) sulfate, we need 4 moles potassium iodide.

Look at the chemical equation below. This reaction can be interpreted many ways.

$$N_2O_3 + H_2O \rightarrow 2\ HNO_2$$

- One molecule of dinitrogen trioxide plus one molecule of water yields two molecules of hydrogen nitrite.
- One mole of dinitrogen trioxide plus one mole of water yields two moles of hydrogen nitrite.

Example: For each of the following equations, indicate the number of formula units or molecules, and the number of moles present in the balanced chemical equation.

(a) $2\ C_2H_6 + 7\ O_2 \rightarrow 4\ CO_2 + 6\ H_2O$

(b) $KBrO_3 + 6\ KI + 5\ HBr \rightarrow 7\ KBr + 3\ I_2 + 3\ H_2O$

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.
Two moles of C_2H_6 plus seven moles of O_2 yields four moles of CO_2 plus six moles of H_2O.
(b) Two formula units of $KBrO_3$ plus six formula units of KI plus six formula units of HBr yields seven formula units of KBr plus three molecules of I_2 and three molecules of H_2O.
Two moles 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.

Stoichiometry

In chemistry, we "talk" to each other using chemical equations, the same way mathematicians talk to each other using mathematical equations. In chemistry, we also want to talk about quantities. Using stoichiometry, you can predict the quantities of reactants as products that can be used and produced in a chemical reaction. This requires working with balanced chemical equations.

In the previous section we explored mole relationships in balanced chemical equations. In this section, we will use the mole as a conversion factor to calculate moles of product from a given number of moles of reactant or 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 you require to produce a given amount of product or to calculate the mass of product you can obtain from a given mass of reactant.

Mole Ratios

A mole ratio is the relationship of the number of moles of the substances in a reaction. For instance, in the following reaction we read the coefficients as molecules (or formula units) and moles:

$$2\ H_2(g) + O_2(g) \rightarrow 2\ H_2O(l)$$

2 moles of H_2 react with 1 mole of O_2 to produce 2 moles of H_2O. Or, an alternate method to represent this information is with mole ratios. The following mole ratios can be obtained from this reaction:

$$\frac{2\ mol\ H_2}{1\ mol\ O_2}\ or\ \frac{1\ mol\ O_2}{2\ mol\ H_2}\ or\ \frac{2\ mol\ H_2O}{2\ mol\ H_2}\ or\ \frac{2\ mol\ H_2O}{1\ mol\ O_2}\ or\ \frac{2\ mol\ H_2}{2\ mol\ H_2O}\ or\ \frac{1\ mol\ O_2}{2\ mol\ H_2O}$$

Using the coefficients of a balanced reaction, you can compare any two substances in the reaction you are interested in, whether they are reactants or products. 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.

Mole-Mole Calculations

We have already learned the process through which chemists solve many math problems, the factor-label method. The mole-mole ratio we obtain from a balanced reaction can be used as a ratio in part of that process.

Example: If only 0.050 mol of magnesium hydroxide, $Mg(OH)_2$, is present, how many moles of phosphoric acid, H_3PO_4, would be required for the reaction?

$$2\ H_3PO_4 + 3\ Mg(OH)_2 \rightarrow Mg_3(PO_4)_2 + 6\ H_2O$$

Solution: We need to set up this problem using the same steps of dimensional analysis.
Given: 0.050 mol $Mg(OH)_2$
Find: mol H_3PO_4
The ratio we need is one that compares mol $Mg(OH)_2$ to mol H_3PO_4. This is the ratio obtained in the balanced reaction. Note that there are other reactants and products in this reaction, but we don't need to use them to solve this problem.

$$0.050\ \cancel{mol\ Mg(OH)_2} \cdot \frac{2\ mol\ H_3PO_4}{3\ \cancel{mol\ Mg(OH)_2}} = \mathbf{0.033\ mol\ H_3PO_4}$$

Notice if the equation was not balanced, the amount of H_3PO_4 would have been different. The reaction MUST be balanced to use the reaction in any calculations. As you can see, the mole ratios are useful for converting between the number of moles of one substance and another.

Calculations Using a Mole Map

Being able to perform mass-mass calculations allows you to determine the mass of reactant (how many grams) you require to produce a given amount of product; or to calculate the mass of product you can obtain from a given mass of reactant or the mass of reactant needed to react with a specific amount of another reactant. Just as when working with mole ratios, it is important to make sure you have a balanced chemical equation before you begin.

These types of problems can be done using dimensional analysis, also called the factor-label method. 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 the correct answer.

For example, 15.0 g of chlorine gas is bubbled over liquid sulfur to produce disulfur dichloride. How much sulfur, in grams, is needed according to the balanced equation:

$$Cl_2(g) + 2\ S(l) \rightarrow S_2Cl_2(l)$$

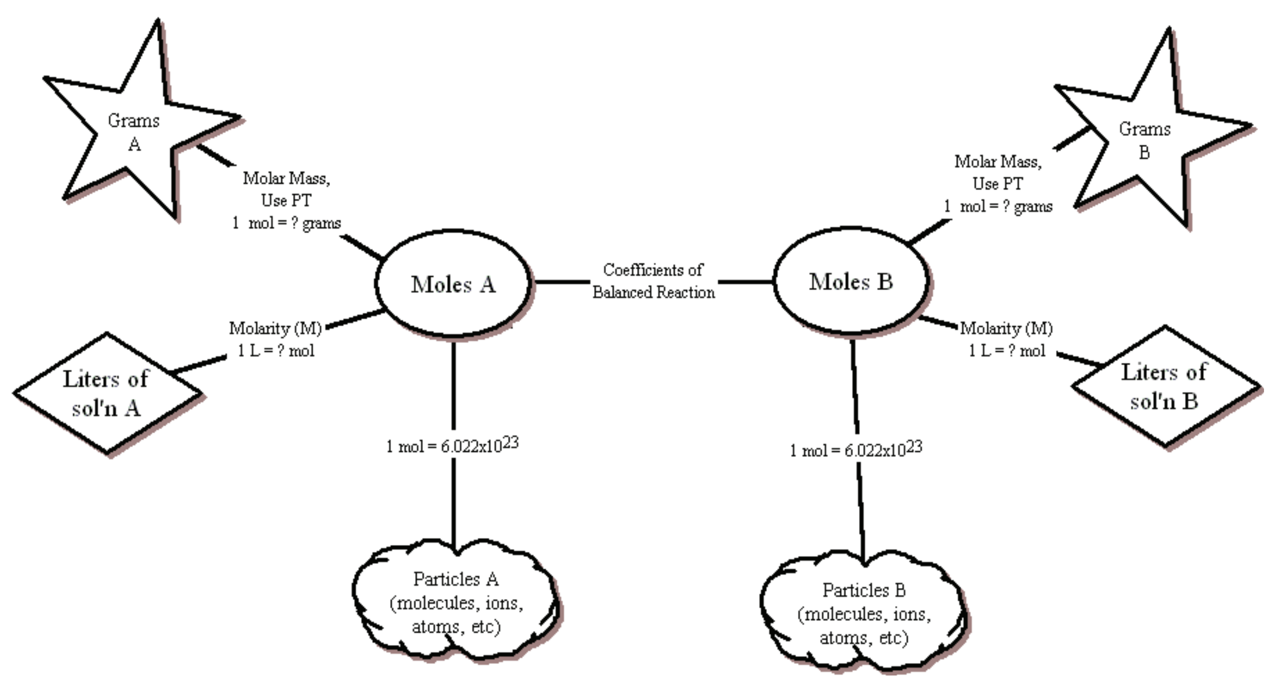
1. Identify the **given:** 15.0 g Cl_2
2. Identify the **find:** g S
3. Next, use the correct **ratios** that allow you to cancel the units you don't want and get to the unit you are calculating for.

$$15.0\ \cancel{g\ Cl_2} \times \frac{1\ \cancel{mol\ Cl_2}}{71.0\ \cancel{g\ Cl_2}} \times \frac{2\ \cancel{mol\ S}}{1\ \cancel{mol\ Cl_2}} \times \frac{32.1\ g\ S}{1\ \cancel{mol\ S}} = 13.6\ g\ S$$

If we combine the mole-mole ratio with ratios we learned previously, when we first learned about the mole, we have several ratios we can use to solve a wide variety of problems. The mole map is a tool we can use to help us to know which ratios to use when solving problems.

You use this map much like you would use a road map. You must first find out where you are on the map (your given units) and where you would like to go (your "find" units). The map will then let you know which roads (ratios) to take to get there. Let's see how this works with a couple of example problems.

The Mole Map



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Example: The thermite reaction is a very exothermic reaction which produces liquid iron, given by the following balanced equation:

$$Fe_2O_3(s) + 2\ Al(s) \rightarrow 2\ Fe(l) + Al_2O_3(s)$$

If 5.00 g of iron is produced, how much iron(III) oxide was placed in the original container?

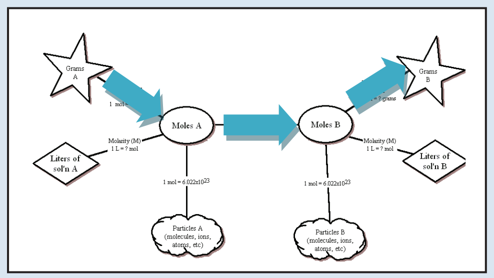
Solution:

1) Identify the "given": 5.00 g iron. (Even though this is a product, it is still the measurement given to us in the problem.)

2) Identify the units of the "find": g Fe_2O_3 (remember, mass is measured in grams)

3) Ratios: This is where the map comes in handy. To start with, we are at 5.00 g Fe. For this problem, then, "A" on the map stands for Fe. We start at grams A.

We want to know g Fe_2O_3. For this problem, "B" stands for Fe_2O_3. We are heading to grams B.



Our map tells us this problem will take 3 ratios (3 roads from g A to g B): molar mass of A, mol:mol ratio from a balanced reaction, and molar mass of B. To solve our problem, the work will look like:

$$5.00\ \cancel{g\ Fe} \cdot \frac{1\ \cancel{mol\ Fe}}{55.85\ \cancel{g\ Fe}} \cdot \frac{1\ \cancel{mol\ Fe_2O_3}}{2\ \cancel{mol\ Fe}} \cdot \frac{159.7\ g\ Fe_2O_3}{1\ \cancel{mol\ Fe_2O_3}} = \mathbf{7.17\ g\ Fe_2O_3}$$

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? The balanced reaction is:

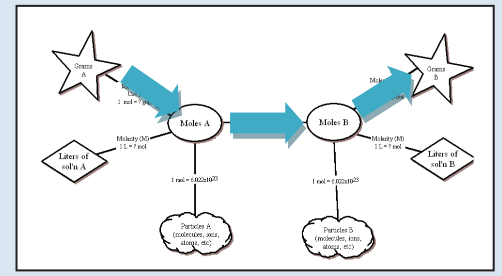
$$2\ C_{13}H_{18}O_2 + 33\ O_2 \rightarrow 26\ CO_2 + 18\ H_2O$$

Solution:

Given: 200. g $C_{13}H_{18}O_2$ (g A on the map)

Find: g CO_2 (g B on the map)

Ratios: The map says we need to use the molar mass of $C_{13}H_{18}O_2$, then the coefficients of the balanced reaction, then the molar mass of CO_2.



$$200.\ \cancel{g\ C_{13}H_{18}O_2} \cdot \frac{1\ \cancel{mol\ C_{13}H_{18}O_2}}{206.3\ \cancel{g\ C_{13}H_{18}O_2}} \cdot \frac{26\ \cancel{mol\ CO_2}}{2\ \cancel{mol\ C_{13}H_{18}O_2}} \cdot \frac{44.01\ g\ CO_2}{1\ \cancel{mol\ CO_2}} = \mathbf{555\ g\ CO_2}$$

Example: If sulfuric acid is mixed with sodium cyanide, the deadly gas hydrogen cyanide is produced. How many moles of sulfuric acid would have been placed in the container to produce 12.5 g of hydrogen cyanide? The balanced reaction is:

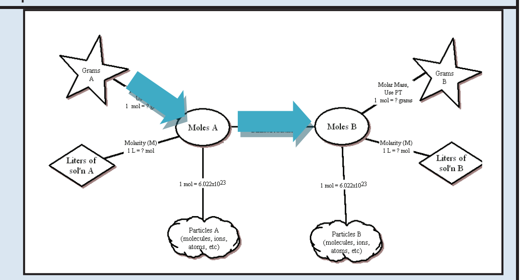
$$2\ NaCN + H_2SO_4 \rightarrow Na_2SO_4 + 2\ HCN$$

Solution:

Given: 12.5 g HCN (g A on map)

Find: mol H_2SO_4 (mol A on map)

Ratios: The mole map says we need the molar mass of HCN and the coefficients of the balanced reaction.



$$12.5\ \cancel{g\ HCN} \cdot \frac{1\ \cancel{mol\ HCN}}{27.0\ \cancel{g\ HCN}} \cdot \frac{1\ mol\ H_2SO_4}{2\ \cancel{mol\ HCN}} = \mathbf{0.231\ mol\ H_2SO_4}$$

Example: How many atoms of carbon would be released from the complete dehydration of 18.0 g of sugar ($C_6H_{12}O_6$) with sulfuric acid? The balanced reaction is:

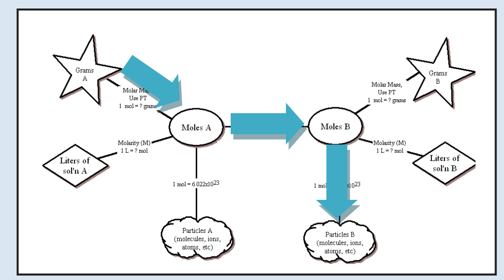
$$C_6H_{12}O_6 + H_2SO_4 \rightarrow 6\ C + 7\ H_2O + SO_3$$

Solution:

Given: 18 g C6H12O6

Find: atoms C

Ratios: the mole map says we need the molar mass of the sugar, the balanced reaction, and finally Avogadro's number.



$$18.0\ g\ C_6H_{12}O_6 \times \frac{1\ mol\ C_6H_{12}O_6}{180.g\ C_6H_{12}O_6} \times \frac{6\ mol\ C}{1\ mol\ C_6H_{12}O_6} \times \frac{6.02\times10^{23} atoms\ C}{1\ \cancel{mol\ C}} = \mathbf{3.6\times10^{23} atoms\ C}$$

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.
- The coefficients in a balanced chemical equation represent the reacting ratios of the substances in the reaction.
- The coefficients of the balanced equation can be used to determine the ratio of moles of all the substances in a reaction.

Vocabulary

- **Stoichiometry:** the calculation of quantitative relationships of the reactants and products in a balanced chemical equation
- **formula unit:** the empirical formula of an ionic compound
- **Mole ratio:** the ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation

Further Reading / Supplemental Links

- Stoichiometry: http://www.lsua.us/chem1001/stoichiometry/stoichiometry.html

7.6: Review Questions

1) Given the reaction between ammonia and oxygen to produce nitrogen monoxide, how many moles of water vapor can be produced from 2 mol of ammonia? The balanced reaction is: $4\ NH_3(g) + 5\ O_2(g) \rightarrow 4\ NO(g) + 6\ H_2O(g)$

2) When properly balanced, how many moles of bismuth(III) oxide can be produced from 0.625 mol of bismuth? The unbalanced reaction is: $Bi(s) + O_2(g) \rightarrow Bi_2O_3(s)$

3) Solid lithium reacts with an aqueous solution of aluminum chloride to produce aqueous lithium chloride and solid aluminum. The reaction is: $3\ Li + AlCl_3 \rightarrow 3\ LiCl + Al$. How many moles of lithium chloride are formed if 5.0 mol aluminum were produced?

For the given balanced reaction: $Ca_3(PO_4)_2 + 3\ SiO_2 + 5\ C \rightarrow 3\ CaSiO_3 + 5\ CO + 2\ P$

4) How many moles of silicon dioxide are required to react with 0.35 mol of carbon?
5) How many moles of calcium phosphate are required to produce 0.45 mol of calcium silicate?

For the given balanced reaction, $4\ FeS + 7\ O_2 \rightarrow 2\ Fe_2O_3 + 4\ SO_2$

6) How many moles of iron(III) oxide are produced from 1.27 mol of oxygen?
7) How many moles of iron(II) sulfide are required to produce 3.28 mol of sulfur dioxide?

8) Given the reaction between copper (II) sulfide and nitric acid, how many grams of nitric acid will react with 2.00 g of copper(II) sulfide?
$$3\ CuS(s) + 8\ HNO_3(aq) \rightarrow 3\ Cu(NO_3)_2(aq) + 2\ NO(g) + 4\ H_2O(l) + 3\ S(s)$$

9) When properly balanced, what mass of iodine was needed to produce 2.5 g of sodium iodide in the equation below? $I_2(aq) + Na_2S_2O_3(aq) \rightarrow Na_2S_4O_6(aq) + NaI(aq)$

10) Determine the mass of lithium hydroxide produced when 0.38 grams of lithium nitride reacts with water according to the following equation: $Li_3N + 3H_2O \rightarrow NH_3 + 3LiOH$

11) If 3.01×10^{23} formulas of cesium hydroxide are produced according to this reaction: $2Cs + 2H_2O \rightarrow 2CsOH + H_2$, how many grams of cesium reacted?

12) How many liters of oxygen are necessary for the combustion of 425 g of sulfur, assuming that the reaction occurs at STP? The balanced reaction is:
$S + O_2 \rightarrow SO_2$ (hint: one mole of oxygen is 22.4 Liters at STP)

13) If I have 2.0 grams of carbon monoxide, how many molecules of carbon monoxide are there?

14) What mass of oxygen is needed to burn 3.5 g of propane (C_3H_8) is burned according to the following equation: $C_3H_8 + 5O_2 \rightarrow 4H_2O + 3CO_2$

15) How many grams of water are produced if 5 moles of oxygen react according to the following reaction? $2H_2 + O_2 \rightarrow 2H_2O$

7.7: Reversible reaction & Equilibrium

Objectives

- Describe the three possibilities that exist when reactants come together.
- Describe what is occurring in a system at equilibrium.

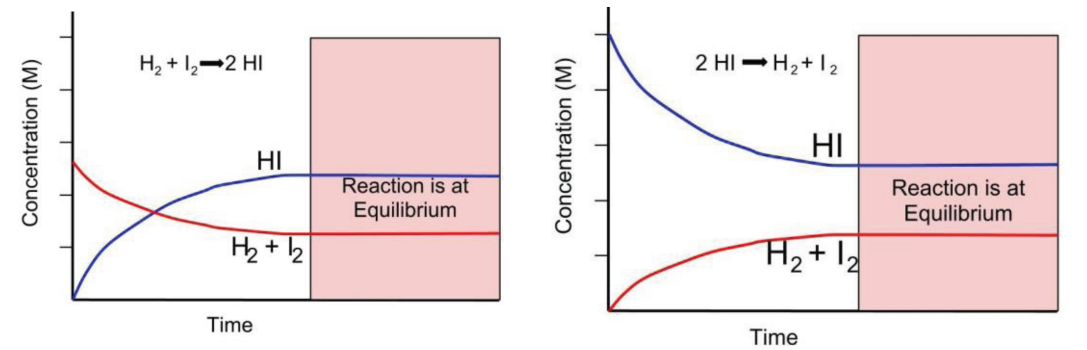
Introduction

Think for a minute about sitting down to a table to eat dinner. There are three possibilities that could happen when you eat dinner. You could (1) finish your entire dinner, (2) you could not want any of it and leave it all on your plate, or (3) you could eat some of it and leave some of it. Reactions have the same possibilities. Reactions also do not always proceed all the way from start to finish. You may have reactions that (1) go to completion so that at the end the reaction vessel contains all products and only products. Some reactions (2) may not start at all so at the end the reaction vessel contains all reactants and only reactants. And some reactions (3) may start but not go to completion, that is, the reaction might start but not go completely to products. In this last case, at the end, the reaction vessel would contain come reactants and some products. In this chapter, we are going to take a closer look at the third type of reaction.

Reversible Reactions and Equilibrium

Consider the hypothetical reaction: $A + B \rightarrow C + D$. If we looked at this reaction using what we have learned, this reaction will keep going, forming C and D until A and B run out. This is what we call an “irreversible reaction” or a “reaction that goes to completion”.

Some reactions, however, are **reversible**, meaning the reaction can go backwards in which products react to form reactants, so that: $A + B \leftarrow C + D$. The direction of the arrow shows that C and D are reacting to form A and B. What if the two reactions, the forward reaction and the reverse reaction, were occurring at the same time? What would this look like? If you could peer into the reaction, you would be able to find A, B, C, and D particles. A and B would react to form C and D at the same time that C and D are reacting to form A and B. If the forward and reverse reactions are happening at the same rate, the reaction is said to be at **equilibrium** or **dynamic equilibrium**. At this point, the concentrations of A, B, C, and D are not changing (or are constant) and we would see no difference in our reaction container, but reactions are still occurring in both directions. It is important to point out that having constant amounts of reactants and products does NOT mean that the concentration of the reactants is equal to the concentration of the products. It means they are not changing. These reactions appear to have stopped before one of the reactants has run out.

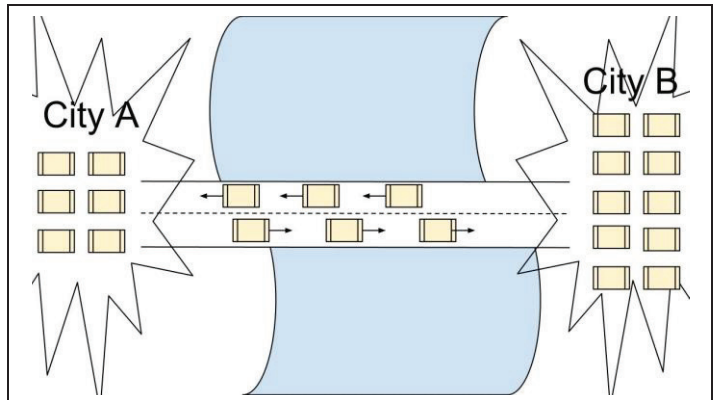


In reversible reactions, not all of the reactants are used to make products. Instead, both reactants and products are left over at the end.

Chemists use a double-headed arrow, ⇌, to show that a reaction is at equilibrium. We would write the example reaction as: $A + B \rightleftharpoons C + D$. The arrow indicates that both directions of the reaction are happening.

Another way to think about reversible and irreversible reactions is to compare them to two types of games of tag. Reversible reactions are in many ways like a traditional game of tag: The "it" person can become "not it" and somebody who is "not it" is tagged and becomes "it". In this way it is a reversible change. It is also like a reaction at equilibrium, because overall no change is occurring. There is always a constant number of "it" people and "not it" people in the game. Also, having constant numbers of "it" and "not it" people in our game does not mean that the number of "it" people (reactants) is equal to the number of "not it" people. Furthermore, this is similar to equilibrium in that this game never truly ends (unless everybody gets tired of playing). The game could go on forever. We could write this as the following reversible reaction:

"It" ⇋ "Not it"



Pretend a bridge connects two cities separated by a river. This situation models equilibrium if the rate that the cars move between the cities is the same. This does not mean that the same number of cars are in City A as are in City B.
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Irreversible reactions (those that only go in one direction from reactants to products and cannot reach a state of equilibrium) is more like a game of sharks and minnows. In sharks and minnows almost everybody starts out as a minnow. Once tagged, they become a shark. However, the difference here is that once you are a shark you are always a shark; there is no way to go back to becoming a minnow. The game continues until everybody has been tagged and becomes a shark. This is similar to irreversible reactions in that the reactants turn into products, but can't change back. Furthermore, the reaction will proceed until the reactants have been used up and there isn't any more left. We could write the reaction as:

Minnow → Shark

Lesson Summary

- There are a few possible ways a reaction can go: It can go to completion (reactants → products); it can occur but not go to completion. Instead it would reach chemical equilibrium (reactants ⇌ products).
- Chemical equilibrium occurs when the number of particles becoming products is equal to the number of particles becoming reactants.
- A dynamic equilibrium is a state where the rate of the forward reaction is equal to the rate of the reverse reaction.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki/Chemical_equilibrium

Vocabulary

- Equilibrium: A state that occurs when the rate of forward reaction is equal to the rate of the reverse reaction.

7.7: Review Questions

1) For the reaction $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$, describe what is happening to make this an equilibrium reaction.

2) If the following table of concentration vs. time was provided to you for the ionization of acetic acid. When does the reaction reach equilibrium? How do you know?

#2	
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

3) The word "equilibrium" comes from the word "equal". What does the term equal mean in this definition?

Indicate whether each of the following statements is true or false for a system in equilibrium.

4) The amount of products is equal to the amount of reactants.
5) The amount of product is not changing.
6) The amount of reactant is not changing.
7) Particles (atoms/molecules) are not reacting.
8) The rate of the forward reaction is equal to the rate of the reverse reaction.

7.8: Equilibrium Constant

Objectives

- Write equilibrium constant expressions.
- Use equilibrium constant expressions to solve for unknown concentrations.
- Use known concentrations to solve for the equilibrium constants.
- Explain what the value of K means in terms of relative concentrations of reactants and products.

Introduction

In the previous section, you learned about reactions that can reach a state of equilibrium, in which the concentration of reactants and products aren't changing. If these amounts are changing, we should be able to make a relationship between the amount of product and reactant when a reaction reaches equilibrium.

The Equilibrium Constant

Equilibrium reactions are those that do not go to completion but are in a state where the reactants are reacting to yield products and the products are reacting to produce reactants. In a reaction at equilibrium, the equilibrium concentrations of all reactants and products can be measured. The **equilibrium constant** (K) is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants. Sometimes, subscripts are added to the equilibrium constant symbol K, such as K_{eq}, K_c, K_p, K_a, K_b, and K_{sp}. These are all equilibrium constants and are subscripted to indicate special types of equilibrium reactions.

There are some rules about writing equilibrium constant expressions that you must learn:

1. Concentrations of products are multiplied on the top of the expression. Concentrations of reactants are multiplied together on the bottom.
2. Coefficients in the equation become exponents in the equilibrium expression.
3. Leave out solids and liquids, as their concentrations do not change in a reaction

Example: Write the equilibrium expression for:

$$CO(g) + 3\,H_2(g) \rightleftharpoons CH_4(g) + H_2O(g)$$

Solution:

$$K = \frac{[CH_4][H_2O]}{[CO][H_2]^3}$$

*Note that the coefficients become exponents. Also, note that the concentrations of products in the numerator are *multiplied*. The same is true of the reactants in the denominator

Example: Write the equilibrium expression for:

$$2\,TiCl_3(s) + 2\,HCl(g) \rightleftharpoons 2\,TiCl_4(s) + H_2(g)$$

Solution:

$$K = \frac{[H_2]}{[HCl]^2}$$

*Note that the solids are left out of the expression completely

Example: Write the equilibrium expression for:

$$P_4(s) + 6\,Cl_2(g) \rightleftharpoons 4\,PCl_3(s)$$

Solution:

$$K = \frac{1}{[Cl_2]^6}$$

*Note that the only product is a solid, which is left out. That leaves just 1 on top in the numerator

Example: Write the equilibrium expression for:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

Solution:

$$K = [H^+][OH^-]$$

Mathematics with Equilibrium Expressions

The equilibrium constant value is the ratio of the concentrations of the products over the reactants. This means we can use the value of K to predict whether there are more products or reactants at equilibrium for a given reaction.

If the equilibrium constant is "1" or nearly "1", it indicates that the molarities of the reactants and products are about the same. If the equilibrium constant value was a large number, like 100, or a very large number, like $1\text{x}10^{15}$, it indicates that the products (numerator) is a great deal larger than the reactants. That means that at equilibrium, the great majority of the material is in the form of products and we say the "products are strongly favored". If the equilibrium constant is small, like 0.10, or very small, like $1\text{x}10^{-12}$, it indicates that the reactants are much larger than the products and the reactants are strongly favored. With large K values, most of the material at equilibrium is in the form of products and with small K values, most of the material at equilibrium is in the form of the reactants.

The equilibrium expression is an equation that we can use to solve for K or for the concentration of a reactant or product.

Example : For the reaction, $SO_2(g) + NO_2(g) \rightleftharpoons SO_3(g) + NO(g)$ determine the value of K when the equilibrium concentrations are: $[SO_2]=1.20$ M, $[NO_2]=0.60$ M, $[NO]=1.6$ M, and $[SO_3]=2.2$ M.

Solution:

Step 1: Write the equilibrium constant expression:

$$K = \frac{[SO_3][NO]}{[SO_2][NO_2]}$$

Step 2: Substitute in given values and solve:

$$K = \frac{(2.2)(1.6)}{(1.20)(0.60)} = \mathbf{4.9}$$

Example: Consider the following reaction: $CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$; K=1.34 If the $[H_2O]=0.100$ M, $[H_2]=0.100$ M, and $[CO_2]=0.100$ M at equilibrium, what is the equilibrium concentration of CO?

Solution:

Step 1: Write the equilibrium constant expression:

$$K = \frac{[H_2][CO_2]}{[CO][H_2O]}$$

Step 2: Substitute in given values and solve:

$$1.34 = \frac{(0.100)(0.100)}{[CO](0.100)}$$

Solving for [CO], we get: [CO]=**0.0746 M**

Lesson Summary

- The equilibrium expression is a mathematical relationship that shows how the concentrations of the products vary with the concentration of the reactants.
- If the value of K is greater than 1, the products in the reaction are favored; if the value of K is less than 1, the reactants in the reaction are favored; if K is equal to 1, neither reactants nor products are favored.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki/Chemical_equilibrium
- Crabapples & Equlibrium: http://www.chem.ox.ac.uk/vrchemistry/ChemicalEquilibrium/HTML/page05.htm
- Equilibrium Animation / Applet: Dots: http://chemconnections.org/Java/equilibrium/

Vocabulary

- Equilibrium constant (K): A mathematical ratio that shows the concentrations of the products divided by concentration of the reactants.

7.8: Review Questions

1) Which phases of substances are not included in the equilibrium expression?

Write an equilibrium expression for each reaction:

2) $2\ H_2(g) + O_2(g) \rightleftharpoons 2\ H_2O(g)$
3) $2\ NO(g) + Br_2(g) \rightleftharpoons 2\ NOBr(g)$
4) $NO(g) + O_3(g) \rightleftharpoons O_2(g) + NO_2(g)$
5) $CH_4(g) + H_2O(g) \rightleftharpoons CO(g) + 3\ H_2(g)$
6) $CO(g) + 2\ H_2(g) \rightleftharpoons CH_3OH(g)$
7) $2\ C_2H_6(g) + 7\ O_2(g) \rightleftharpoons 4\ CO_2(g) + 6\ H_2O(g)$
8) $C_2H_6(g) \rightleftharpoons C_2H_4(g) + H_2(g)$
9) $Hg(g) + I_2(g) \rightleftharpoons HgI_2(g)$
10) $SnO_2(s) + 2\ CO(g) \rightleftharpoons Sn(s) + 2\ CO_2(g)$
11) $Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2\ OH^-(aq)$

12) What does a large value for K imply?
13) What does a small value of K imply?

14) Consider the following equilibrium system: $2\ NO(g) + Cl_2(g) \rightleftharpoons 2\ NOCl(g)$. At a certain temperature, the equilibrium concentrations are as follows: [NO]=0.184 *M*, $[Cl_2]$=0.165 *M*, [NOCl]=0.060 *M*. What is the equilibrium constant for this reaction?

15) For the reaction: $MgCl_2(s) + \frac{1}{2}\ O_2(g) \rightleftharpoons 2\ MgO(s) + Cl_2(g)$. The equilibrium constant was found to be 3.86 at a certain temperature. If $[O_2]$=0.560 M at equilibrium, what is the concentration of $Cl_2(g)$?

16) Consider the equilibrium: $CO(g) + H_2O(g) \rightleftharpoons H_2(g) + CO_2(g)$.
   a) Write an equilibrium expression for this reaction.
   b) If [CO]=0.200M, $[H_2O]$=0.500M, $[H_2]$=0.32M and $[CO_2]$=0.42M, find K.

17) Hydrogen sulfide decomposes according to the equation: $2H_2S(g) \rightleftharpoons 2H_2(g) + S_2(g)$.
   a) Write an equilibrium expression for this reaction.
   b) At equilibrium, the concentrations of each gas are as follows: $[H_2S]=7.06\times10^{-3}$M, $[H_2]=2.22\times10^{-3}$M and $[S_2]=1.11\times10^{-3}$M. What is K_{eq}?

18) Given the following system in equilibrium: $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$
   a) Write an equilibrium expression for the reaction.
   b) If K=85.0, would you expect to find more reactants or products at equilibrium? Why?
   c) If $[SO_2]$=0.0500 M and $[O_2]$=0.0500M, what is the concentration of SO_3 at equilibrium?

7.9: The Effects of Applying Stress to Reactions at Equilibrium

Objectives

- State Le Châtelier's Principle.
- Describe the effect of concentration on an equilibrium system.
- Describe the effect of temperature as a stress on an equilibrium system.

Introduction

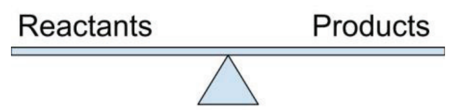
When a reaction has reached equilibrium with a given set of conditions, if the conditions are not changed, the reaction will remain at equilibrium forever. The forward and reverse reactions continue at the same equal and opposite rates and the macroscopic properties remain constant.

It is possible, however, to disturb that equilibrium by changing 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 and 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. That's the topic of this section.

Le Châtelier's Principle

In the late 1800's, a chemist by the name of Henry-Louis Le Châtelier was studying stresses that were applied to chemical equilibria. He formulated a principle from this research and, of course, the principle is called Le Chatelier's Principle. **Le Châtelier's Principle** 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.

Le Chatelier's principle is not an explanation of what happens on the molecular level to cause the equilibrium shift, 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.

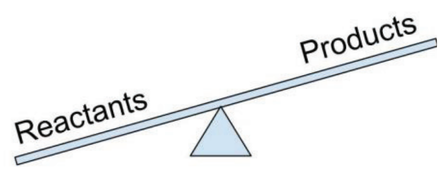


**Thinking of a teeter-totter is a good way to remember Le Chatlier's Principle. If a change is made, what does the reaction need to do to get back to equilibrium?*
CC Tracy Poulsen

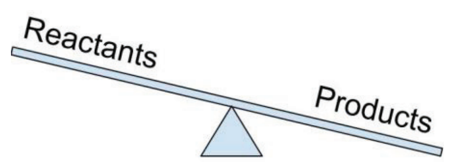
Effect of Concentration Changes on a System at Equilibrium

For instance, if a stress is applied by increasing the concentration of a reactant, the reaction will adjust in such a way that the reactants and products can get back to equilibrium. In this case, you made it so there is too much reactant. The reaction will use up some of the reactant to make more product. We would say the reaction "shifts to the products" or "shifts to the right". If you increase the concentration of a product, you have the opposite effect. The reaction will use up some of the product to make more reactant. The reaction "shifts to the reactants" or "shifts to the left".

What is we remove some reactant or product? If a stress is applied by lowering a reactant concentration, the reaction will try to replace some of the missing reactant. It uses up some of the product to make more reactant, and the reaction "shifts to the reactants". If a stress is applied by reducing the concentration of a product, the equilibrium position will shift toward the products.



**If more reactant is added or a product is removed, the reaction must shift to make more products to get back to equilibrium.*



**If more product is added or a reactant is removed, the reaction must shift to make more reactants to get back to equilibrium.*
CC Tracy Poulsen

Example: For the reaction: $SiCl_4(g) + O_2(g) \rightleftharpoons SiO_2(s) + 2\ Cl_2(g)$, what would be the effect on the equilibrium system if:
(a) $[SiCl_4]$ increases
(b) $[O_2]$ increases
(c) $[Cl_2]$ increases

Solution:
(a) $[SiCl_4]$ increases: The equilibrium would shift to the right
(b) $[O_2]$ increases: The equilibrium would shift to the right
(c) $[Cl_2]$ increases: The equilibrium would shift left

Example: Here's a reaction at equilibrium. $A(aq) + B(aq) \rightleftharpoons C(aq) + D(aq)$
(a) Which way will the equilibrium shift if you add some A to the system without changing anything else?
(b) Which way will the equilibrium shift if you add some C to the system without changing

anything else?

Solution:
(a) The equilibrium will shift toward the products (forward).
(b) The equilibrium will shift toward the reactants (backward).

The Effect of Changing Temperature on a System at Equilibrium

Le Chatelier's principle also correctly predicts the equilibrium shift when systems at equilibrium are heated and cooled. An increase in temperature is the same as adding heat to the system. Consider the following equilibrium:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) \quad \Delta H = -\ 191\text{ kJ}$$

We will learn more about this later, but ΔH has to do with the change in energy, usually heat, for this reaction. The negative sign (-) in the ΔH indicates that energy is being given off. This equation can also be written as:

$$2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g) + 191\text{ kJ of heat}$$

What's important to remember about increasing the temperature of an equilibrium system, is the energy can be thought as just another product or reactant. In this example, you can clearly see that the 191 kJ are a product. Therefore when the temperature of this system is raised, heat is being added and the effect will be the same as increasing any other product. Increasing a product causes the reaction will use up some of the products to make more reactants. And, if the temperature for this equilibrium system is lowered, the equilibrium will shift to make up for this stress. When the temperature is decreased for this reaction, the reaction will shift toward the products in an attempt to counteract the decreased temperature. Therefore, the $[SO_3]$ will increase and the $[SO_2]$ and $[O_2]$ will decrease.

In some reactions, though, heat is a reactant. These reactions are called **endothermic reactions**. These reactions would have the opposite effect. If heat is a reactant, adding head adds a reactant and the reaction will shift towards the products. If heat is removed (by lowering the temperature) from an endothermic reaction, a reactant is removed and the reaction will shift to make more reactants.

The Effect of Temperature on an Endothermic and an Exothermic Equilibrium System

Temperature Change	Exothermic (-ΔH)	Endothermic (+ΔH)
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.
(a) $H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g) \quad \Delta H = +\ 40\text{kJ/mol}$
(b) $2\ SO_2(g) + O_2(g) \rightleftharpoons 2\ SO_3(g)$ + energy

Solution:
(a) The reaction is endothermic, because ΔH is positive, meaning heat is a reactant. We would write: $40\text{ kJ} + H_2(g) + CO_2(g) \rightleftharpoons CO(g) + H_2O(g)$
With an increase in temperature for an endothermic reaction, the reactions will shift right

producing more products.
(b) The reaction is exothermic, meaning heat is a product. With an increase in temperature for an exothermic reaction, the reactions will shift left producing more reactants.

The Haber Process

Let's look at a particularly useful reaction and how chemists applied Le Chatlier's Principle to make more of a desires product. The reaction between nitrogen gas and hydrogen gas can produce ammonia, NH_3. 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 small yield of ammonia. The reaction is as follows:

$$N_2(g) + 3\ H_2(g) \rightleftharpoons 2\ NH_3(g) + \text{energy}$$

Fritz Haber, a German chemist working in the early years of the 20th century, applied Le Châtelier's principle to help solve this problem. Decreasing the concentration of ammonia, for instance, by immediately removing it from the reaction container will cause the equilibrium to shift to the right and continue to produce more product. There were a number of other ways that

One more factor that will affect this equilibrium system is the temperature. Since the forward reaction is exothermic (heat is released as a product), lowering the temperature will once again shift the equilibrium system to the right and increase the ammonia that is produced. Specifically the conditions that were found to produce the greatest yield of ammonia are 550°C (in commercial situations this is a "low" temperature) and 250 atm of pressure. Once the equilibrium system is producing the ammonia, the product is removed, cooled and dissolved in water.

Lesson 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 producing less products.
- Decreasing the concentration of a product causes the equilibrium to shift to the right producing more products.
- For a forward exothermic reaction, an increase in temperature shifts the equilibrium toward the reactant side whereas a decrease in temperature shifts the equilibrium toward the product side.

Further Reading / Supplemental Links

- http://en.wikipedia.org/wiki
- Tutorial: Le Chatlier's Principle: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/lechv17.swf

Vocabulary

- Le Châtelier's Principle: Applying a stress to an equilibrium system causes the equilibrium position to shift to offset that stress and regain equilibrium.

- Exothermic reaction: A reaction in which heat is released, or is a product of a reaction.
- Endothermic reaction: A reaction in which heat is absorbed, or is a reactant of a reaction.
- Catalyst: A substance that increases the rate of a chemical reaction but is, itself, left unchanged, at the end of the reaction.

7.9: Review Questions

1) What is the effect on the equilibrium if the concentration of a reactant is increased?
2) What is the effect on the equilibrium if the concentration of a reactant is decreased?

For the reaction: $N_2O_5(s) \rightleftharpoons NO_2(g) + O_2(g)$, what would be the effect on the equilibrium if:

3) $[NO_2]$ decreases
4) $[NO_2]$ increases
5) $[O_2]$ increases

For the reaction: $C(s) + H_2O(g) \rightleftharpoons CO(g) + H_2(g)$, what would be the effect on the equilibrium system if:

6) $[H_2O]$ increases
7) $[CO]$ increases
8) $[H_2]$ decreases

Predict the effect on the equilibrium position if the temperature is increased in each of the following.

9) $H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$ $\Delta H = +51.9$ kJ
10) $P_4O_{10}(s) + H_2O(l) \rightleftharpoons H_3PO_4(aq)$ + heat
11) $Ag^+(aq) + Cl^-(aq) \rightleftharpoons AgCl(s)$ $\Delta H = -112$ kJ/mol.
12) $2\ NOBr_{(g)} \rightleftharpoons 2\ NO(g) + Br_2(g)$ $\Delta H = +16.1$ kJ.

In the following reaction, what would be the effect of each of the following changes to the system at equilibrium? $C(s) + O_2(g) \rightleftharpoons CO_2(g)$ $\Delta H = -393.5$ kJ/mol

13) increase O_2
14) increase the temperature

Predict the effect on the equilibrium: $H_2O(g) + CO(g) \rightleftharpoons H_2(g) + CO_2(g)$ $\Delta H = -42$ kJ when each of the following changes are made to the equilibrium system.

15) Temperature is increased
16) $[CO_2]$ decreases
17) $[H_2O]$ increases
18) $[H_2]$ decreases

Predict the effect on the chemical equilibrium $2\ SO_3(g) + \text{heat} \rightleftharpoons 2\ SO_2(g) + O_2(g)$, when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?

19) Temperature is increased
20) $[O_2]$ decreases

Predict the effect on the chemical equilibrium: $N_2O_4(g)$ + heat ⇌ 2 $NO_2(g)$, when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?

21) Temperature is decreased

22) $[N_2O_4]$ decreases

Chapter 8: Describing Acids & Bases

8.1: Classifying Acids and Bases

Objectives

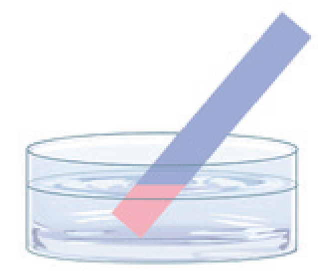
- List the properties of acids.
- List the properties of bases.
- Define an Arrhenius acid and list some substances that qualify as acids under this definition.
- Define an Arrhenius base and list some substances that qualify as bases under this definition.

Introduction

We may not realize how much acids and bases affect our lives. Have you ever thought of drinking a can of soda pop and actually drinking acid? Have you looked at bottles of household cleaners and noticed what the main ingredients were? Have you ever heard a shampoo commercial and heard them say that the shampoo was "pH balanced" and wondered what this means and why it is so important for hair? Thanks to the beginning work of scientists in the latter part of the 19^{th} century, we started to learn about acids and bases; our study continued and is constantly growing. Let's begin our study of this wonderful branch of chemistry.

Properties of Acids

Acids are a special group of compounds with a set of common properties. This helps to distinguish them from other compounds. Thus, if you had a number of compounds and you were wondering whether these were acids or otherwise, you could identify them by their properties. But what exactly are the properties? Think about the last time you tasted lemons. Did they taste sour, sweet, or bitter? Lemons taste sour. This is a property of acids. Another property of acids is that they turn blue litmus paper red. Litmus paper is an **indicator**, which is a substance that changes color depending on how acidic or basic something is. If blue litmus paper turns red when it is dipped into a solution, then the solution is an acid. Another property of acids that many people are familiar with is their ability to cause burns to skin. This is why it is a bad idea to play with battery acid or other acids.



Acids cause blue litmus paper to turn red

Acids react with many metals to produce hydrogen gas. For some examples, look at the reactions below:

$Zn(s) + 2\ HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

$Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

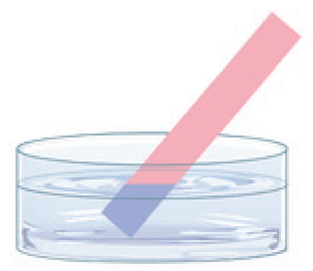
What do you notice that is the same for all three equations? In each case, the reactants are a metal (Zn or Mg) and an acid (HCl). They all produce hydrogen gas, H_2. This is another property of acids. Acids react with most metals to produce hydrogen gas.

Think about the last time you took an aspirin or a vitamin C tablet. Aspirin is acetylsalicylic acid while vitamin C is ascorbic acid; both are acids that can produce H^+ ions when dissolved in water. Acetic acid ($HC_2H_3O_2$) is a component of vinegar, hydrochloric acid (HCl) is stomach acid, phosphoric acid (H_3PO_4) is commonly found in dark soda pop, sulfuric acid H_2SO_4 is used in car batteries and formic acid HCO_2H is what causes the sting in ant bites. For all of these acids, the chemical formula of an acid begins with one or more hydrogen atoms. Acids dissolve in water to make H^+ ions. Because they make ions (charged particles) when they are dissolved, acids will also conduct electricity when they are dissolved in water.

We interact with acids on a daily basis so some knowledge of their properties and interactions is essential. Acids are present in our everyday lives.

Properties of Bases

There is one common base that some may have had the opportunity to taste: milk of magnesia, which is a slightly soluble solution of magnesium hydroxide. This substance is used for acid indigestion. Flavorings have been added to improve the taste, otherwise it would have a bitter taste when you drink it. Other common bases include substances like Windex, Drano, oven cleaner, soaps and many cleaning other products. Please note: do not taste any of these substances. A bitter taste is one property you will have to take for granted. Bases also tend to have a slippery feel. This matches what you have experienced with soaps and detergents.



Bases cause red litmus paper to turn blue

As with acids, bases have properties that allow us to distinguish them from other substances. We have learned that acids turn blue litmus paper red. Bases turn red litmus paper blue. Notice that the effect of the indicator is the opposite of that of acids.

Most acids have formulas that start with H. On the other hand, most of the bases we will be using in this course have formulas that end with –OH. These bases contain the polyatomic ion called hydroxide. When bases dissolve in water, they produce hydroxide (OH^-) ions. Because they dissolve into charged particles, bases will also conduct electricity when they are dissolved.

Although many people have already heard of the danger of acids at causing burns, many bases are equally dangerous and can also cause burns. It is important to be very careful and to follow correct safety procedures when dealing with both acids and bases.

Acids & Bases Defined

Although scientists have been able to classify acids and bases based on their properties for some time, it took a while to come up with a theory explaining why some substances were acidic and others were basic. Svante Arrhenius set the groundwork for our current understanding of acid-base theory. We will focus on his famous acid-base definitions. This was quite an accomplishment for a scientist in the late 19th century with very little technology, but with the combination of knowledge and intellect available at the time Arrhenius led the way to our understanding of how acids and bases differed, their properties, and their reactions. Keep in mind that Arrhenius came up with these theories in the late 1800's so his definitions came with some limitations. For now we will focus on his definitions.

Arrhenius Acids

Take a look at all of the following chemical equations. What do you notice about them? What is common for each of the equations below?

Hydrochloric acid: $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Nitric acid: $HNO_3(aq) \rightarrow H^+(aq) + NO_3^-(aq)$

Perchloric acid: $HClO_4(aq) \rightarrow H^+(aq) + ClO_4^-(aq)$

One of the distinguishable features about acids is the fact that acids produce H^+ ions in solution. If you notice in all of the above chemical equations, all of the compounds dissociated to produce H^+ ions. This is the one main, distinguishable characteristic of acids and the basis for the Arrhenius definition of acids. An Arrhenius **acid** is a substance that produces H^+ ions in solution.

Arrhenius Bases

In contrast, an Arrhenius **base** is a substance that releases OH^- ions in solution. Many bases are ionic substances made up of a cation and the anion hydroxide, OH^-. The dissolving equation for the base sodium hydroxide, NaOH, is shown below:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

Barium hydroxide produces a similar reaction when dissociating in water:

$$Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2\ OH^-(aq)$$

The production of OH^- ions is the definition of bases according to the Arrhenius.

Lesson Summary

- Acids turn blue litmus paper red, taste sour, and react with metals to produce hydrogen gases.
- Common acids include vinegar ($HC_2H_3O_2$), phosphoric acid in soda pop (H_3PO_4) and stomach acid HCl.
- Bases turn red litmus paper blue, have a bitter taste, and are slippery to the touch.
- Common bases include Drano (NaOH), soaps and detergents, milk of magnesia ($Mg(OH)_2$) and Windex (NH_4OH).
- Arrhenius defined an acid as a substance that donates H^+ ions when dissociating in solution.
- An Arrhenius base is a substance that releases OH^- ions in solution.

Vocabulary

- Arrhenius acid: a substance that produces H^+ ions in solution
- Arrhenius base: a substance that produces OH^- ions in a solution

Further Reading / Supplemental Links

- Strong & Weak Acids animation: http://www.mhhe.com/physsci/chemistry/chang7/esp/folder_structure/ac/m2/s1/acm2s1_1.htm
- Tutorial: Acids & Bases http://visionlearning.com/library/module_viewer.php?mid=58&l=&c3=

8.1: Review Questions

Indicate whether each of the following is a property of acids, bases, or both acids and bases.

1) Have a sour taste
2) Taste bitter
3) Turns litmus paper red
4) Feels slippery
5) React with metals
6) Turns litmus paper blue
7) What is the Arrhenius definition of an acid?

8.2: pH

Objectives

- State the $[H^+]$, $[OH^-]$, and K_w values for the self-ionization of water.
- Define and describe the pH scale and describe how logarithmic scales work
- Calculate $[H^+]$, $[OH^-]$, and pH given the value of any one of the other values in a water solution at 25°C.
- Explain the relationship between the acidity or basicity of a solution and the hydrogen ion concentration, $[H^+]$, and the hydroxide ion concentration, $[OH^-]$, of the solution.
- Predict whether an aqueous solution is acidic, basic, or neutral from the $[H^+]$, $[OH^-]$, or the pH.

Introduction

We have been discussing what makes an acid or a base and what properties acids and bases have. It is frequently useful to compare how acidic or basic a solution is in comparison to other solutions. A couple of ways to do this is to compare $[H^+]$ to $[OH^-]$ or to find the pH of a solution.

Relationship Between $[H^+]$ and $[OH^-]$

We have learned that acids and bases are related to hydrogen ions $[H^+]$ and hydroxide ions $[OH^-]$. Both of these ions are present in both acids and bases. However, they are also present in pure water. Water self-ionizes according to the following reaction:

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

The equilibrium expression for this reaction would be:

$$K_w=[H^+][OH^-]$$

The equilibrium constant for this particular equilibrium is K_w, meaning the equilibrium constant for water. From experimentation, chemists have determined that in pure water, $[H^+]=1\text{x}10^{-7}$ M and $[OH^-]=1\text{x}10^{-7}$ M. If you substitute these values into the equilibrium expression, you find that $K_w=1\text{x}10^{-14}$. Any solution which contains water, even if other things are added, will shift to establish this equilibrium. Therefore, for any solution, the following relationship will always be true:

$$K_w=1\text{x}10^{-14}=[H^+]\,[OH^-] \text{ @ } 25°C$$

We can describe whether a solution is acidic, basic, or neutral according to the concentrations in this equilibrium.

- If $[H^+] = [OH^-]$, the solution is neutral (such as in pure water)
- If $[H^+] > [OH^-]$, the solution is acidic. This means that $[H^+] > 1\text{x}10^{-7}$ M.

- If $[H^+] < [OH^-]$, the solution is basic. This means that $[OH^-] > 1\text{x}10^{-7}$ M.

We can use this equation to calculate the concentrations of H^+ and OH^-. Consider the following example.

Example: Suppose acid is added to some water, and $[H^+]$ is measured to be $1\text{x}10^{-4}$ M. What would $[OH^-]$ be?

Solution: substitute what we know into the equilibrium expression:

$$K_w=1\text{x}10^{-14}=[H^+]\,[OH^-]$$
$$1\text{x}10^{-14}=[1\text{x}10^{-4}][OH^-]$$

To isolate $[OH^-]$, divide by sides by $1\text{x}10^{-4}$.

This leaves, $\mathbf{[OH^-]=1x10^{-10}\ M}$

Note that because $[H^+] > [OH^-]$, the solution must be acidic.

Suppose, on the other hand, something is added to the solution that reduces the hydrogen ion concentration, a base.

Example: If the final hydrogen ion concentration is $1\text{x}10^{-12}$ M, we can calculate the final hydroxide ion concentration.

Solution:

$$K_w=1\text{x}10^{-14}=[H^+]\,[OH^-]$$
$$1\text{x}10^{-14}=[1\text{x}10^{-12}]\,[OH^-]$$

To isolate $[OH^-]$, divide by sides by $1\text{x}10^{-12}$.

This leaves, $\mathbf{[OH^-]=1x10^{-2}\ M}$

Note that because $[H^+] < [OH^-]$, the solution must be basic.

Using the K_w expression, anytime we know either the $[H^+]$ or the $[OH^-]$in a water solution, we can always calculate the other one.

Example: What would be the $[H^+]$ for a grapefruit found to have a $[OH^-]$ of $1.26\text{x}10^{-11}$? What is $[H^+]$ and is the solution acidic, basic, or neutral?

Solution:

$$K_w=1\text{x}10^{-14}=[H^+]\,[OH^-]$$
$$1\text{x}10^{-14}=[H^+]\,[1.26\text{x}10^{-11}]$$

To isolate $[H^+]$, divide by sides by $1.26\text{x}10^{-11}$.

This leaves, $\mathbf{[H^+]=7.94x10^{-4}\ M}$

Also, the solution must be acidic because $[H^+] > [OH^-]$

pH Scale

A few very concentrated acid and base solutions are used in industrial chemistry and inorganic laboratory situations. For the most part, however, acid and base solutions that occur in nature, those used in cleaning, and those used in organic or 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 $1.0\text{x}10^{-14}$ M. Expressing hydrogen ion concentrations in exponential numbers becomes tedious and is difficult for those not trained in chemistry. A Danish chemist named Søren Sørensen developed a shorter method for expressing acid strength or hydrogen ion concentration with a non-exponential number. He named his method **pH**. The p from pH comes from the German word *potenz* meaning "power or the exponent of". Sørensen's idea that the pH would be a simpler number to deal with in terms of discussing acidity level led him to a formula that relates pH and $[H^+]$:

$$pH = -\log [H^+]$$

If the hydrogen ion concentration is between 1.0 M and 1.0×10^{-14}, the value of the pH will be between 0 and 14.

Example: Calculate the pH of a solution given that $[H^+]=0.01$ M.
Solution: $pH = -\log(0.01)$ **pH = 2**

Sometimes you will need to use a calculator.

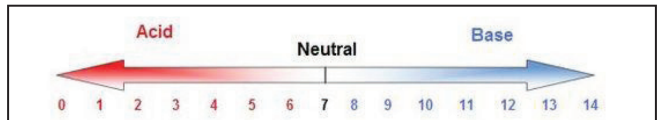
Example: Calculate the pH of saliva with $[H^+]=1.58 \times 10^{-6}$ M.
Solution: $pH = -\log(1.58 \times 10^{-6})$ **pH = 5.8**

If you are given [OH-] it is still possible to find the pH, but it requires one more step. You must first find [H+] and then use the pH equation.

Example: Calculate the pH of a solution with $[OH^-]=7.2 \times 10^{-4}$ M.
Solution: In order to find pH, we need [H+]. $K_w=1 \times 10^{-14}=[H^+][OH^-]$ $1 \times 10^{-14}=[H^+][7.2 \times 10^{-4}]$ To isolate $[H^+]$, divide by sides by 7.2×10^{-4}.   This leaves, $[H^+]=1.39 \times 10^{-11}$ M   We can now find the pH $pH = -\log(1.39 \times 10^{-11})$ **pH = 10.9**

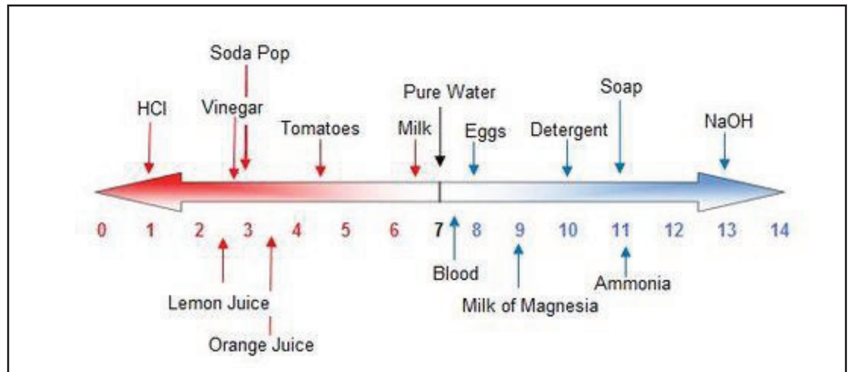
The pH scale developed by Sørensen is a logarithmic scale, which means that a difference of 1in pH units indicates a difference of a factor of 10 in the hydrogen ion concentrations. A difference of 2 in pH units indicates a difference of a factor of 100 in the hydrogen ion concentrations. Not only is the pH scale a logarithmic scale but by defining the pH as the *negative* log of the hydrogen ion concentration, the numbers on the scale get smaller as the hydrogen ion concentration gets larger. For example, pH=1is a stronger acid than pH=2 and, it is stronger by a factor of 10 (the difference between the pH's is 1).

The closer the pH is to 0 the greater the concentration of $[H^+]$ ions and therefore the more acidic the solution. The closer the pH is to 14, the higher the concentration of OH^- ions and the stronger the base.



The pH Scale.

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 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 to see how these common examples relate in terms of their pH.



pH Scale for Common Substances.

Example: Compare lemon juice (pH=2.5) to milk (pH=6.5). Answer each of the following: a) Label each as acidic, basic, or neutral   b) Which has a higher concentration of H^+ ions?   c) How many times more H^+ does that solution have?
Solution: a) Both lemon juice and milk are acidic, because their pH's are less than 7. (*Note: milk is only very slightly acidic as its pH is very close to 7)   b) The lower the pH, the higher the concentration of H^+ ions. Therefore, lemon juice has more H^+.   c) Each step down on the pH scale increases the H^+ concentration by 10 times. It is 4 steps down on the pH scale to go from 6.5 to 2.5. Therefore, lemon juice has 10x10x10x10 or 10,000 times more H^+ ions than milk.

Lesson Summary

- Water ionizes slightly according to the equation $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$
- The equilibrium constant for the dissociation of water is: $K_w=1 \times 10^{-14}=[H^+][OH^-]$
- $pH = -\log [H^+]$.

8.2: Review Questions

1) In saturated limewater, $[H^+]=3.98 \times 10^{-13}$ M.
   a) Find $[OH]^-$
   b) What is the pH?
   c) Is the solution acidic, basic, or neutral?
2) In butter, $[H^+]=6.0 \times 10^{-7}$ M.
   a) Find $[OH]^-$
   b) What is the pH?
   c) Is the solution acidic, basic, or neutral?
3) In peaches, $[OH^-]=3.16 \times 10^{-11}$ M
   a) Find $[H^+]$
   b) What is the pH?

c) Is the solution acidic, basic, or neutral?

4) During the course of the day, human saliva varies between being acidic and basic. If $[OH^-]=3.16\times10^{-8}$ M,
   a) Find $[H^+]$
   b) What is the pH?
   c) Is the solution acidic, basic, or neutral?
5) A solution contains 4.33×10^{-8}M hydroxide ions. What is the pH of the solution?

6) A solution contains a hydrogen ion concentration of 6.43×10^{-9} M. What is the pH of the solution?

7) If the pH of one solution is 5 less than another solution, how does the amount of H^+ in each solution compare? Which has more H^+? How many times more?

8.3: Neutralization

Objectives

- Explain what is meant by a neutralization reaction
- Write the balanced equation for the reaction that occurs when an acid reacts with a base.

Introduction

Neutralization is a reaction between an acid and a base that produces water and a salt. The general reaction for the neutralization reaction is shown below.

$$\text{acid} + \text{base} \rightarrow \text{salt} + \text{water}$$

In this section, we will be writing the products of neutralization reactions.

Neutralization Reactions

Acids are a combination of hydrogen ions (H^+) and an anion. Examples include HCl, HNO_3, and $HC_2H_3O_2$. Bases can be a combination of metal cations and hydroxide ions, OH^-. Examples include NaOH, KOH, and $Mg(OH)_2$. According to the Arrhenius definitions of acids and bases, the acid will contribute the H^+ ion that will react to neutralize the OH^- ion, contributed by the base, to produce 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. Look at the following equations. What do they have in common?

$$HClO_4 + NaOH \rightarrow NaClO_4 + HOH$$
$$H_2SO_4 + 2\ KOH \rightarrow K_2SO_4 + 2\ HOH$$

(Note: HOH is the same as 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. The H^+ ion from the acid will neutralize the OH^- ion from the base to form water. The other product is a salt formed when the cation of the base combines with the anion of the acid. Remember, the total charge on the salt MUST be zero. You must have the correct number of cations and anions to cancel out the charges of each.

Example: Complete the following neutralization reactions.

(a) $H_2SO_4 + Ba(OH)_2 \rightarrow$

(b) $HCOOH + Ca(OH)_2 \rightarrow$

(c) $HCl + NaOH \rightarrow$

Solution:

(a) The H^+ in H_2SO_4 will combine with the OH^- part of $Ba(OH)_2$ to make water (H_2O or HOH). The salt produced is what is formed when Ba^{2+} (the cation from the base) combines with SO_4^{2-}(the anion from the acid). These have charges of +2 and -2, so the formula for this compound is $BaSO_4$.

Before it is balanced, the reaction is:

$$H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + H_2O$$

After balancing, we get:

$$\mathbf{H_2SO_4 + Ba(OH)_2 \rightarrow BaSO_4 + 2\ H_2O}$$

(b) The H^+ in HCOOH will combine with the OH^- part of $Ca(OH)_2$ to make water (H_2O or HOH). The salt produced is what is formed when Ca^{2+} (the cation from the base) combines with $COOH^-$ the anion from the acid). These have charges of +2 and -1, so the formula for this compound is $Ca(COOH)_2$.

Before it is balanced, the reaction is:

$$HCOOH + Ca(OH)_2 \rightarrow Ca(COOH)_2 + H_2O$$

After balancing, we get:

$$\mathbf{2\ HCOOH + Ca(OH)_2 \rightarrow Ca(COOH)_2 + 2\ H_2O}$$

(c) The H^+ in HCl will combine with the OH^- part of NaOH to make water (H_2O or HOH). The salt produced is what is formed when Na^+ (the cation from the base) combines with Cl^- the anion from the acid). These have charges of +1 and -1, so the formula for this compound is NaCl.

Before it is balanced, the reaction is:

$$\mathbf{HCl + NaOH \rightarrow NaCl + H_2O}$$

The reaction is already balanced, so we are done.

Lesson Summary

- A neutralization reaction between an acid and a base will produce a salt and water.

Vocabulary

- Neutralization: a reaction between an acid and a base that produces water and a salt

8.3: Review Questions

Write a balanced reaction for each of the following neutralization reactions:

1) $HNO_3 + KOH \rightarrow$
2) $HClO_4 + NH_4OH \rightarrow$
3) $H_2SO_4 + NaOH \rightarrow$
4) $HNO_3 + NH_4OH \rightarrow$
5) $HF + NH_4OH \rightarrow$
6) $HC_2H_3O_2 + KOH \rightarrow$
7) $HCl + KOH \rightarrow$

8) Milk of magnesia, $Mg(OH)_2$ is a common over- the - counter antacid that has, as its main ingredient, magnesium hydroxide. It is used by the public to relieve acid indigestion. Acid indigestion is caused by excess stomach acid, HCl, being present.
9) Hydrochloric acid (HCl) reacts with barium hydroxide.
10) Sodium hydroxide reacts with perchloric acid ($HClO_4$).

8.4: Titration

Objectives

- Explain what an acid/base indicator is.
- Explain how a titration is performed
- Calculate the concentration of unknown acid or base when given the concentration of the other and the volume needed to reach the equilibrium point in a titration.

Introduction

For acid-base neutralization reactions, the typical laboratory procedure for determining the stoichiometric amounts of acid and/or base in the reaction is to complete a titration. As we go through this section, we will use some of the prior knowledge we have obtained about acids and bases, chemical reactions, and molarity calculations, to apply them to the concept of titrations.

Indicators

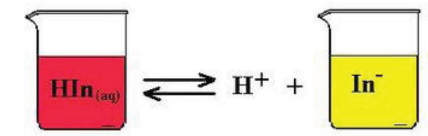
An **indicator** is a substance that changes color at a specific pH and is used to indicate the pH of the solution. Litmus paper is a paper that has been dipped in an indicator. The litmus paper is called an indicator because it is used to indicate whether the solution is an acid or a base. If the red litmus paper turns blue, the solution is basic (pH > 7), if the blue litmus turns red the solution is acidic (pH < 7).

Hydrangeas are also a natural indicator. The petals will change colors based on the pH of the soil.

The juice from red cabbage can be used to prepare an indicator paper. It contains the chemical anthrocyanin, which is the active ingredient in the indicator. Red beets, blueberries, and cranberries are other great examples of a naturally occurring indicators. Another example of a natural indicator is flowers. Hydrangea is a common garden plant with flowers that come in many colors depending on the pH of the soil. If you are travelling around and see a hydrangea plant with blue flowers, the soil is acidic, the creamy white flowers indicate the soil is neutral, and the pink flowers mean the soil is basic.

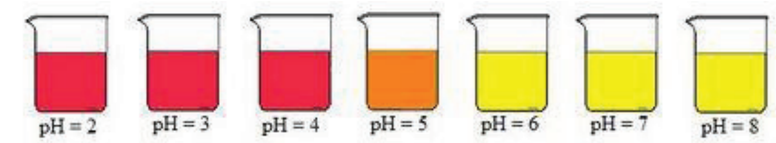
There are two requirements for a substance to function as an acid-base indicator; 1) the substance must have an equilibrium affected by hydrogen ion concentration, and 2) the two forms of the compound on opposite sides of the equilibrium must have different colors. Most indicators function in the same general manner and can be presented by a generic indicator equation. In the equation shown in the figure, 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^-.



The color of an indicator depends on the pH of the solution.

For the example above, HIn red and In^- yellow. If we add hydrogen ion to the solution, the equilibrium will be driven toward the reactants and the solution will turn red. If we add base to the solution (reduce hydrogen ion concentration), the equilibrium will shift toward the products and the solution will turn yellow. It is important to note that if this indicator changes color at pH=5, then at all pH values less than 5, the solution will be red and at all pH values greater than 5, the solution will be yellow. Therefore, putting this indicator into a solution and having the solution turn yellow does NOT tell you the pH of the solution . . . it only tells you that the pH is greater than 5 . . . it could be 6, 7, 8, 9, etc. There are many indicators that are available to be used to help determine the pH of solutions.

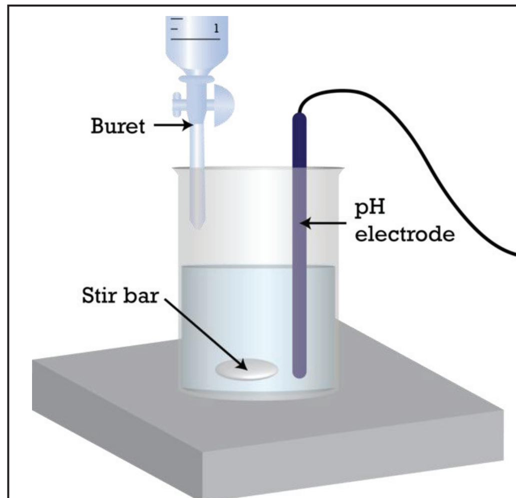


The color change of an indicator occurs over a very short range.

The Titration Process

One of the properties of acids and bases is that they neutralize each other to form water and a salt. In the laboratory setting, an experimental procedure where an acid is neutralized by a base (or vice versa) is known as titration. **Titration**, by definition, 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.

So what does one do in a titration? When doing a titration, you need to have a few pieces of equipment. A **buret** is used to accurately dispense the volume of the solution of known concentration (either the base or the acid). A flask is used to hold a known, measured volume of the unknown concentration of the other solution (either the acid or the base).



The set up of a titration

If the basic solution was in the buret, you would first read the volume of base in the buret at the beginning. You would add the base to the flask containing the acid until all of the acid has reacted and then read the volume of base in the buret again. To see how much was added, you would subtract the initial volume from the final volume.

In a titration, just enough base is added to completely react with all of the acid, without extra base being added. This is called the **equivalence point** because you have added equal moles of acid and base. For most acids and bases, this point is difficult to see, because the acid and base reactants as well as the salt and water products have no color. This is where indicators come in. An **indicator** is used to determine the equivalence of the titration. A few drops of the indicator are added to the flask before you begin the titration. If an appropriate indicator has been chosen, the indicator will only react and change color (and stay color changed) when all of the other acid has reacted. Therefore, the indicator will change color immediately after enough base was added to completely react with all of the acid (the equivalence point).

Some laboratories have pH meters that measures this point more accurately than the indicator, although an indicator is much more visual. The main purpose of a pH meter is to measure the changes in pH as the titration goes from start to finish. It is also possible to determine the equivalence point using the pH meter as the pH will change dramatically once all of the acid and base have been neutralized.

The Mathematics of Titration

For the calculations involved here, we will restrict our acid and base examples where the stoichiometric ratio of H^+ and OH^- is 1:1. The formula for these 1:1 reactions, in which 1 mole of acid is needed to react with 1 mole of base, has the structure:

$$(M_a)(V_a)=(M_b)(V_b)$$

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
- V_b is the volume of the base

This equation works because the left side calculates the number of moles of acid which react and the right side calculates the number of moles of base. To reach the equivalence point,

mol acid = mol base.

Example: When 10.0 mL of a 0.125 M solution of hydrochloric acid, HCl, is titrated with a 0.100 M solution of potassium hydroxide, KOH, what the volume of the hydroxide solution is required to neutralize the acid?

Solution:

Step 1: Write the balanced ionic chemical equation. Check that the acid:base ratio is 1:1.

$$HCl + KOH \rightarrow H_2O + KCl$$

Since 1 HCl is needed for each KOH, the reaction is 1:1.

Step 2: Use the formula and fill in all of the given information. The acid is HCl and the base is KOH.

$$M_a=0.125\text{ M}$$
$$V_a=10.0\text{ mL}$$
$$M_b=0.100\text{ M}$$
$$V_b=?$$

$$(M_a)(V_a)=(M_b)(V_b)$$
$$(0.125\text{ M})(10.0\text{ mL})=(0.100\text{ M})(V_b)$$
$$\mathbf{V_b=12.5\text{ mL}}$$

Therefore, for this weak acid-strong base titration, the volume of base required for the titration is 12.5 mL.

Lesson Summary

- An indicator is a substance that changes color at a specific pH and is used to indicate the pH of the solution.
- A titration is the addition of a known concentration of base (or acid) to a solution of acid (or base) of unknown concentration.
- The equivalence point is the point in the titration where the number of moles of acid equals the number of moles of base, and, if you chose an appropriate indicator, where the indicator changes color.
- For titrations where the stoichiometric ratio of mol H^+: mol OH^- is 1:1, the formula $(M_a)(V_a)=(M_b)(V_b)$ can be used to calculate concentrations or volumes for the unknown acid or base.

Vocabulary

- Titration: the lab process in which a known concentration of base (or acid) is added to a solution of acid (or base) of unknown concentration
- Indicator: a substance that changes color at a specific pH and is used to indicate the pH of the solution
- Equivalence point: the point in the titration where the number of moles of acid equals the number of moles of base

8.4: Review Questions

1) What is an indicator? What is it used for?
2) What is an equivalence point?
3) If 22.50 mL of a sodium hydroxide is necessary to neutralize 18.50 mL of a 0.1430 *M* HNO_3 solution, what is the concentration of NaOH?
4) Calculate the concentration of hypochlorous acid if 25.00 mL of HClO is used in a titration with 32.34 mL of a 0.1320 *M* solution of sodium hydroxide.
5) What volume of 0.45 *M* hydrochloric acid must be added to 15.0 mL of .997 *M* potassium hydroxide to neutralize the base? ($HCl + KOH \rightarrow H_2O + KCl$)
6) What volume of .20 *M* HI is needed to neutralize 25 mL of .50 *M* KOH?
7) What is the molarity of sodium hydroxide if .174L of the solution is neutralized by .20L of 1.2 M HCl? ($HCl + NaOH \rightarrow H_2O + NaCl$)
8) Suppose we used .150L of 0.500*M* NaOH and .250L of vinegar (acetic acid solution) of an unknown concentration. What is the molarity of the vinegar? (Balanced reaction is: $NaOH_{(aq)} + HC_2H_3O_{2\,(aq)} \rightarrow NaC_2H_3O_{2\,(aq)} + H_2O_{(l)}$)

Chapter 9: Energy of Chemical Changes

9.1: Energy

Objectives:

- Distinguish between kinetic and potential energy and give examples of each.

Introduction

Just like matter, energy is a term that we are all familiar with and use on a daily basis. Before you go on a long hike, you eat an *energy* bar; every month, the *energy* bill is paid; on TV, politicians argue about the *energy* crisis. But what is energy? If you stop to think about it, energy is very complicated. When you plug a lamp into an electric socket, you see energy in the form of light, but when you plug a heating pad into that same socket, you only feel warmth. Without energy, we couldn't turn on lights, we couldn't brush our teeth, we couldn't make our lunch, and we couldn't travel to school. In fact, without energy, we couldn't even wake up because our bodies require energy to function. We use energy for every single thing that we do, whether we're awake or asleep.

Types of Energy: Kinetic and Potential

Kinetic energy is energy associated with motion. When an object is moving, it has kinetic energy, and when the object stops moving, it has no kinetic energy. Although all moving objects have kinetic energy, not all moving objects have the same amount of kinetic energy. The amount of kinetic energy possessed by an object is determined by its mass and its speed. The heavier an object is and the faster it is moving, the more kinetic energy it has. Kinetic energy is very common and is easy to spot in the world around you. Sometimes we even capture kinetic energy and use it to power things like our home appliances. Forms of kinetic energy include heat, light, sound, and electricity.

Potential energy is stored energy that remains available until we choose to use it. Think of a battery in a flashlight. If you leave a flashlight on, the battery will run out of energy within a couple of hours. If, instead, you only use the flashlight when you need it and turn it off when you don't, the battery will last for days or even months. Because the battery stores potential energy, you can choose to use the energy all at once, or you can save it and use a small amount at a time.

Any stored energy is potential energy and has the "potential" to be used at a later time. Unfortunately, there are a lot of different ways in which energy can be stored, making potential energy very difficult to recognize. Generally speaking, an object has potential energy due to its position relative to another object.

For some examples of potential energy, though, it's harder to see how "position" is involved. In chemistry, we are often interested in what is called chemical potential energy. **Chemical potential energy** is energy stored in the atoms, molecules, and chemical bonds that make up matter. How does this depend on position? The world and all of the chemicals in it are made up of atoms. These atoms store potential energy that is dependent on their positions relative to one another. Although we cannot see atoms, scientists know a lot about the ways in which atoms interact. This allows them to figure out how much potential energy is stored in a specific quantity of a particular chemical. *Different chemicals have*

different amounts of potential energy because they are made up of different atoms, and those atoms have different positions relative to one another.

Since different chemicals have different amounts of potential energy, scientists will sometimes say potential energy depends on not only position but also composition. Composition affects potential energy because it determines which molecules and atoms end up next to each other. For example, the total potential energy in a cup of pure water is different than the total potential energy in a cup of apple juice because the cup of water and the cup of apple juice are composed of different amounts of different chemicals.

The Law of Conservation of Matter and Energy

While it's important to understand the difference between kinetic energy and potential energy, the truth is energy is constantly changing. Kinetic energy is constantly being turned into potential energy, and potential energy is constantly being turned into kinetic energy. Even though energy can change form, it must still follow the fundamental law: *energy cannot be created or destroyed*, it can only be changed from one form to another. This law is known as the **law of conservation of energy**.

9.2: Endothermic and Exothermic Changes

Objectives

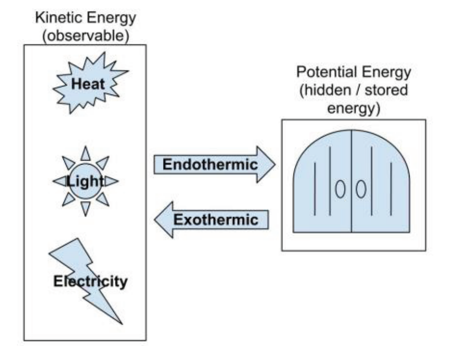
- Define potential energy and kinetic energy.
- Define endothermic and exothermic reactions.
- Describe how heat is transferred in endothermic and exothermic reactions.
- Determine whether a reaction is endothermic or exothermic through observations, temperature changes, or an energy diagram.

All Chemical Reactions Involve Energy

Remember that all chemical reactions involve a change in the bonds of the reactants. The bonds in the reactants are broken and the bonds of the products are formed. Chemical bonds have **potential energy** or "stored energy". Because we are changing the bonding, this means we are also changing how much of this "stored energy" there is in a reaction.

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 by the reaction or energy is taken in by the reaction. There are many types of energy that can be involved in these changes. Different types of energy include:

- Heat
- Electricity
- Light
- Chemical potential energy

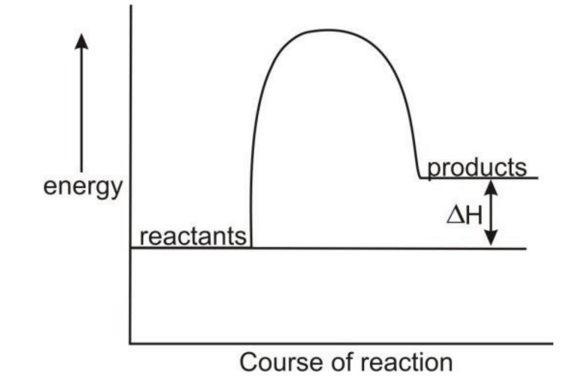


In endothermic changes, kinetic energy (such as heat) is absorbed and changed into chemical potential energy.
In exothermic changes, chemical potential energy is released as heat or other kinetic energy.

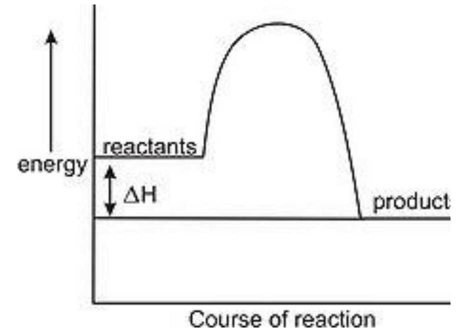
Sometimes the products have more energy stored in their bonds than the reactants had to start with. This means that the reaction started with less hidden energy than we had at the end. Where did this extra energy come from? In these reactions, heat or other forms of energy are absorbed by the reactants from the surroundings to supply some of this hidden bond energy. These reactions are called endothermic reactions. **Endothermic reactions** absorb heat or other forms of energy from their surroundings.

Sometimes the products have less energy stored in their bonds than the reactants had to start with. This means that the reaction started with more hidden energy than we had at the end. Where did this extra energy go? In these reactions, heat or other forms of energy are released by the reactants from the surroundings to give off some of this extra hidden bond energy. These reactions are called exothermic reactions. **Exothermic reactions** release heat or other forms of energy into their surroundings.

The last time you went camping, you might have lit a campfire. The burning of the wood in the fire pit released energy into the environment. But like lighting the fire at the camp, chemical reactions need a minimum amount of energy in order for a reaction to begin. Your campfire won't start until you supply a bit of energy, the initial match, to get it started. Before any reaction can occur, reactant bonds need to be broken. A minimum amount of energy, the **activation energy**, must be supplied before any reaction can take place. This is different than labeling a reaction as endothermic or exothermic. Although sometimes it is necessary to give a little energy to a reaction to get it to start, if more energy is given off than you put in to get it started, the reaction is overall exothermic. If you give the reaction more energy than it gives you, the reaction is endothermic.



The activation energy is the bump on the graph. This reaction is endothermic, because overall, energy must be added to turn reactants into products and the products are higher than the reactants on the graph.



In an Energy Diagram of an exothermic reaction, the products are lower than the reactants. (Source: Therese Forsythe. CC-BY-SA).

Energy changes are frequently shown by drawing an energy diagram. Energy diagrams show the stored/hidden energy of the reactants and products as well as the activation energy. If, on an energy diagram, the products have more stored energy than the reactants started with, the reaction is endothermic. You had to give the reaction energy. If, on the energy diagram, the products have less stored energy than the reactants started with, the reaction is exothermic.

Another way of classifying a reaction as endothermic or exothermic was already presented to you in chapter 7. Remember what was said about ΔH? If ΔH had a positive value, the reaction is endothermic. This means that energy must be added to the reactants in order for the reaction to occur. If ΔH has a negative value, the reaction is exothermic and energy is produced with the rest of the products and given off into the surroundings.

Consider the following equation:

$$2\ HgO \rightarrow 2\ Hg + O_2 \qquad \Delta H = +181.7\ kJ$$

The positive sign (+) on the ΔH tells us that the reaction is endothermic, that more energy had to be added to the reaction and that there is less energy stored in the bonds of the reactant (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.

Contrast the previous reaction to the next reaction:

$$NaCl + AgNO_3 \rightarrow AgCl + NaNO_3 \qquad \Delta H = -166\ kJ$$

This is an example of a chemical reaction in which energy is released. This means that there is less energy stored in the bonds of the products than there was in the bonds in the reactants. Therefore, extra energy was left over when the reactants become the products. The negative sign (-) on the ΔH tells us that this reaction had extra energy. This reaction is **exothermic,** meaning that energy is released. Another way to think of this is that energy is a product, it is something produced, or released, in the reaction.

Example: Label each of the following processes as endothermic or exothermic.
a) water boiling
b) gasoline burning
c) ice forming on a pond
d) $2\ SO_2 + O_2 \rightarrow 2\ SO_3$ $\Delta H = -46.8$ kJ/mol

Solution:
a) endothermic – you must put a pan of water on the stove and give it heat in order to get water to boil. Because you are adding heat/energy, the reaction is endothermic.
b) exothermic – when you burn something, it feels hot to you because it is given off heat into the surroundings
c) exothermic – think of ice forming in your freezer instead. You put water into the freezer, which takes heat out of the water, to get it to freeze. Because heat is being pulled out of the water, it is exothermic. Heat is leaving.
d) This reaction is exothermic, because ΔH is negative.

Lesson Summary

- All chemical reactions involve changes in energy. This may be a change in heat, electricity, light, or other forms of energy.
- Reactions that absorb energy are endothermic.
- Reactions that release energy are exothermic.

Vocabulary

- Potential energy: The energy of position or stored energy, including bond energy.
- Endothermic: reactions in which energy is absorbed
- Exothermic: reactions in which energy is released

Further Reading / Supplemental Links

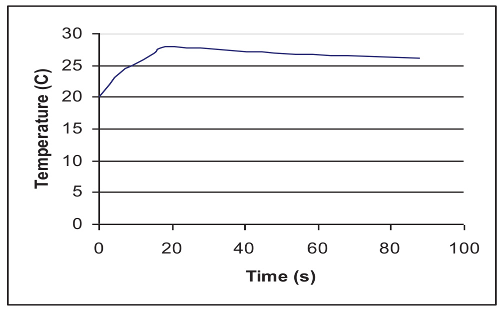
- http://en.wikipedia.org
- Fire: http://www.pbs.org/wgbh/nova/fire/onfire.html#

9.1: Review Questions

1) Define endothermic and exothermic reactions.

Label each of the following processes as endothermic or exothermic
2) natural gas burning
3) melting chocolate
4) fireworks exploding
5) Steam condensing
6) Photosynthesis (plants using light to make sugar)
7) Sugar is dissolved in water in a test tube and the test tube feels cold.
8) Gasoline is burned in a car engine.
9) Water is converted to steam according to the equation: $H_2O_{(l)}$ + heat → $H_2O_{(g)}$

10) Two solutions were mixed and the temperature of the resulting solution was measured over time. Given the following graph, was the reaction exothermic or endothermic?
Explain.



9.3: Oxidation – Reduction

Objectives

- Identify the substance being oxidized and the substance being reduced in an oxidation-reduction reaction
- Identify the anode and the cathode given a diagram of an electrolysis apparatus that includes the compound being electrolyzed.
- Describe how batteries can produce electrical energy

Introduction

Electricity is an important form of energy that you use every day. It runs your calculators, cell phones, dishwashers, and watches. This form of energy involves moving electrons through a wire and using the energy of these electrons.

Batteries are one way of producing this type of energy. Many important chemical reactions involve the exchange of one or more electrons, and, therefore we can use this movement of electrons as electricity. These reactions are called oxidation-reduction (or "redox") reactions.

Oxidation and Reduction

Reactions in which electrons are transferred are called oxidation-reduction (or "redox") reactions. There are two parts to these changes: one atom must lose electrons and another atom must gain them. These two parts are described by the terms "oxidation" and "reduction".

Originally, a substance was said to be oxidized when it reacted with oxygen. Today, the word "oxidized" is still used for those situations, but now we have a much broader second meaning for these words. Today, the broader sense of the word **oxidation** is defined as losing electrons. When a substance loses electrons, its charge will increase. This may feel

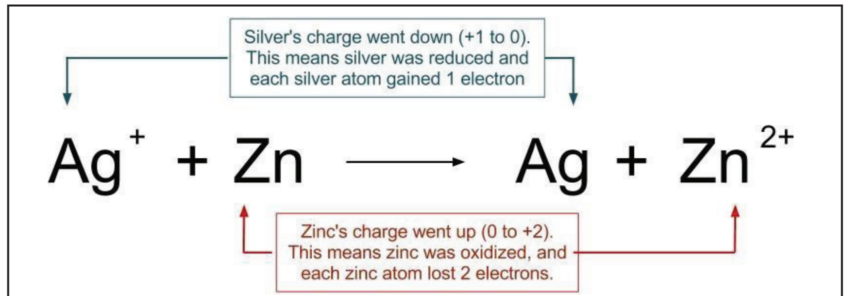
a bit backwards, but remember that electrons are negative. If an atom loses electrons, it is losing negative particles so its charge will increase.

The other half of this process, the gaining of electrons, also needs a name. When an atom or an ion gains electrons, the charge on the particle goes down. For example, if a sulfur atom whose charge is zero (0) gains two electrons, its charge becomes -2 and if an Fe^{3+} ion gains an electron, its charge changes from +3 to +2. In both cases the charge on the particle is reduced by the gain of electrons. Remember that electrons have a negative charge, so gaining electrons will result in the charge decreasing. The word **reduction** is defined to mean gaining electrons and the reduction of charge.

In chemical systems, these two processes (oxidation and reduction) must occur simultaneously and the number of electrons lost in the oxidation must be the same as the number of electrons gained in the reduction. In oxidation-reduction reactions, electrons are transferred from one substance to another. Here's an example of an oxidation – reduction reaction.

$$2\,Ag^{+}(aq) + Cu(s) \rightarrow 2\,Ag(s) + Cu^{2+}(aq)$$

In this reaction, the silver ions are gaining electrons to become silver atoms. Therefore, the silver ions are being reduced and the charge of silver is decreasing. The copper atoms are losing electrons to become copper +2 ions and are therefore, being oxidized and the charge of copper is increasing. Whenever, a chemical reaction involves electrons being transferred from one substance to another, the reaction is an **oxidation – reduction reaction** (or a redox reaction).



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Half-equations are very helpful in discussing and analyzing processes but half-reactions cannot occur as they appear. The half-reactions for the reaction above would be:

$$2\,Ag^{+}(aq) + 2e^{-} \rightarrow 2\,Ag(s)$$
$$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^{-}$$

Both oxidation and reduction must occur at the same time so the electrons are donated and absorbed nearly simultaneously. The two half-reactions may be added together to represent a complete reaction. In order to add the half-reactions, the number of electrons donated and the number of electrons accepted must be equal.

Example: For each reaction, identify what is oxidized and what is reduced.
a) $Zn + HCl \rightarrow H_2 + ZnCl_2$
b) $Fe + O_2 \rightarrow Fe_2O_3$
c) $NaBr + I_2 \rightarrow NaI + Br_2$

Solution: In order to determine what is being oxidized and reduced, we must look at charges of atoms and see if they increase or decrease. (Remember, elements have no charge. In a compound, we can use our periodic table and what we learned in chapter 4 to assign charges). If the charge increases, the atom was oxidized. If the charge decreases, the atom was reduced.

a) This reaction written with charges is:

$$Zn^0 + H^+Cl^- \rightarrow H_2^0 + Zn^{2+}Cl_2^-$$

Zn is oxidized because it went from 0 to +2. H is reduced because it went from +1 to 0. Cl was neither oxidized nor reduced.

b) This reaction written with charges is:

$$Fe^0 + O_2^0 \rightarrow Fe^{3+}O^{2-}$$

Fe is oxidized because it went from 0 to +3. O is reduced because it went from 0 to -2.

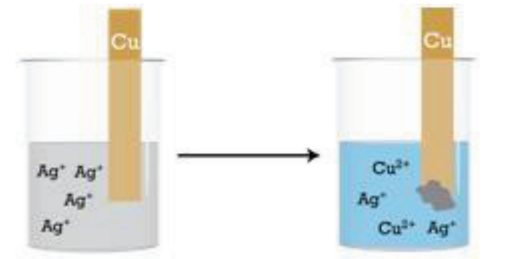
c) This reaction written with charges is:

$$Na^+Br^- + I_2^0 \rightarrow Na^+I^- + Br_2^0$$

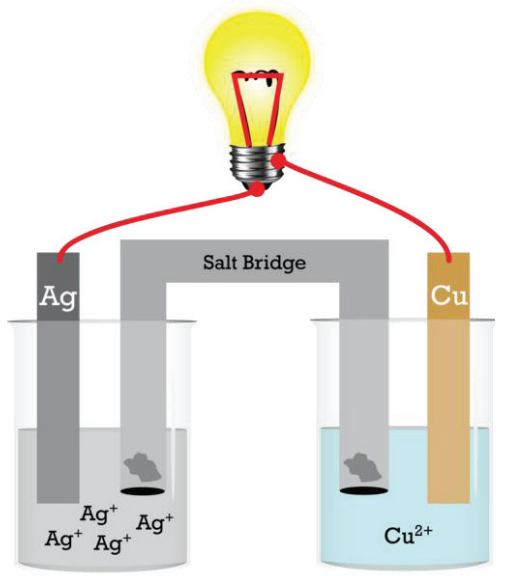
Br is oxidized because it went from -2 to 0. I is reduced because it went from 0 to -1. Na was neither oxidized nor reduced as it stayed +1 the whole time.

Batteries

Batteries are devices use chemical reactions to produce electrical energy. These reactions occur because the products contain less potential energy in their bonds than the reactants. The energy produced from excess potential energy not only allows the reaction to occur but also often gives off energy to the surroundings. Some of these reactions can be physically arranged so the energy given off is given off in the form of an electric current. 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.**



Redox reaction in one beaker does not produce electricity. (Source: Richard Parsons. CC-BY-SA)



In a battery, the oxidation and reduction half-reactions occur in separate containers, allowing the electrons to flow through a wire producing electricity.

If a strip of copper is placed in a solution of silver nitrate, the following reaction takes place:

$$2\ Ag^+(aq) + Cu(s) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$$

In this reaction, copper atoms are donating electrons to silver ions so the silver ions are reduced to silver atoms and 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.

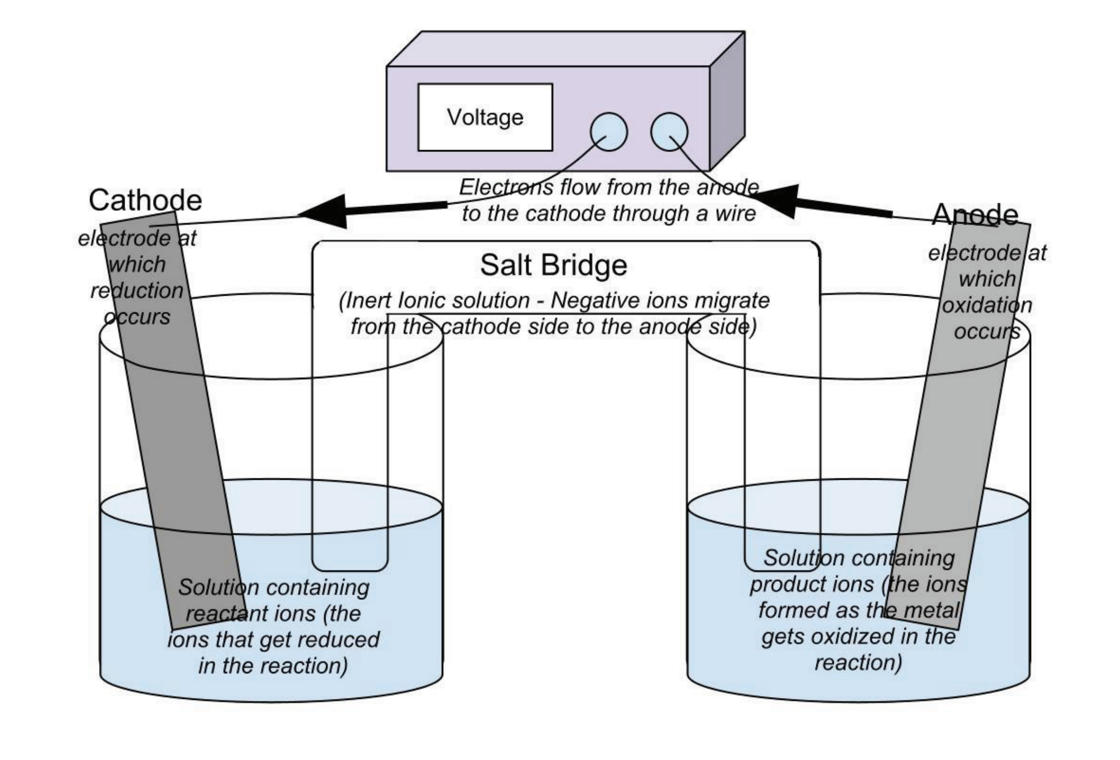
The reaction we just described is not set up in such as way to produce electricity. It is true that electrons are being transferred, but to produce electricity we need electrons flowing through a wire so we can use the energy of these electrons. This reaction, $2\ Ag^+(aq) + Cu(s) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$, is

one that could be arranged to produce electricity. To do this, the two half-reactions (oxidation and reduction) must occur in separate compartments and the separate compartments must remain in contact through an ionic solution and an external wire.

In this electrochemical cell, the copper metal must be separated from the silver ions to avoid a direct reaction. Each electrode in its solution could be represented by a half-reaction.

$$Cu \rightarrow Cu^{2+} + 2\ e-$$
$$2\ Ag^+ + 2\ e- \rightarrow 2\ Ag$$

The wire connects the two halves of the reaction, allowing electrons to flow from one metal strip to the other. In this particular example, electrons will flow from the copper electrode (which is losing electrons) into the silver electrode (which is where the silver ions gain the electrons). The cell produces electricity through the wire and will continue to do so as long as there are sufficient reactants (Ag^+ and Cu) to continue the reaction.



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Electrochemical cells will always have two electrodes, the pieces of metal where electrons are gained or lost. (In this example, the strip of Ag metal and Cu metal are the electrodes.) The electrode where reduction occurs and electrons are gained is called the **cathode.** The electrode where oxidation occurs and the electrons are lost is called the **anode.** Electrons will always move from the anode to the cathode. The electrons that pass through the external circuit can do useful work such as lighting lights, running cell phones, and so forth.

If the light bulb is removed from the circuit with the electrochemical cell and replaced with a voltmeter, 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.

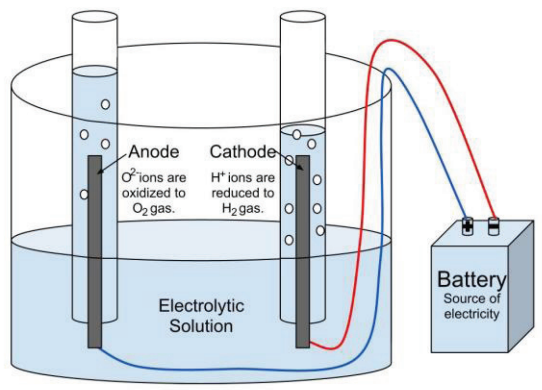
It may seem complicated to construct an electrochemical cell because of all their complexities. Electrochemical cells are actually easy to make and sometimes even occur accidentally. If you take two coins of different denomination and push them part way through the peel of a whole lemon and then connect the two coins with a wire, a small electric current will flow.

Electrolysis

So far we have discussed how electricity can be produced from chemical reactions in batteries. Some reactions will, instead, *use* electricity to get a reaction to occur. In these reactions, electrical energy is given to the reactants causing them to react to form the products. These reactions have many uses.

Electrolysis is a process that involves forcing electricity through a liquid or solution to cause a reaction to occur. Electrolysis reactions will not run unless energy is put into the system from outside. In the case of electrolysis reactions, the energy is provided by the battery.

If electrodes connected to battery terminals are placed in liquid sodium chloride, the sodium ions will migrate toward the negative electrode and be reduced while the chloride ions migrate toward the positive electrode and are oxidized. The processes that occur at the electrodes can be represented by what are called half-equations.



In this image, an electrical current is passed through water, splitting the water into hydrogen and oxygen gases.
CC Tracy Poulsen

Reduction occurs at the positive electrode:

$$Na^+ + e^- \rightarrow Na$$

Oxidation occurs at the negative electrode:

$$2\ Cl^- \rightarrow Cl_2 + 2\ e^-$$

The overall reaction for this reaction is:

$$2\ Na^+ + 2\ Cl^- \rightarrow 2\ Na + Cl_2$$

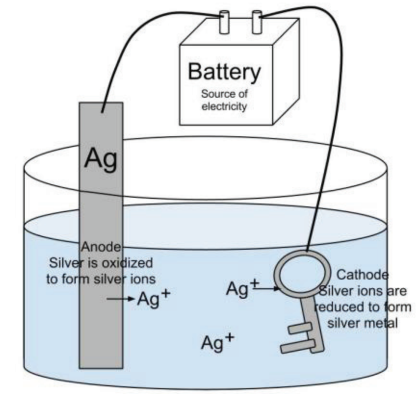
With appropriate treatment from the battery, it is possible to get the metal being reduced in an electrolysis process to adhere strongly to the electrode. The use of electrolysis to coat one material with a layer of metal is called **electroplating.** Usually, electroplating is used to cover a cheap metal with a layer of more expensive and more attractive metal. Many girls buy jewelry that is plated in gold. Sometimes, electroplating is used to get a surface metal that is a better conductor of electricity. When you wish to have the surface properties of gold (attractive, corrosion resistant, or good conductor) but you don't want to have the great cost of making the entire object out of solid gold, the answer may be to use cheap metal to make the object and then electroplate a thin layer of gold on the surface.

To silver plate an object like a spoon (silverware that's plated is less expensive than pure silver), the spoon is placed in the position of the cathode in an electrolysis set up with a solution of silver nitrate. When the current is turned on, the silver ions will migrate through

the solution, touch the cathode (spoon) and adhere to it. With enough time and care, a layer of silver can be plated over the entire spoon. The anode for this operation would often be a large piece of silver from which silver ions would be oxidized and these ions would enter the solution. This is a way of ensuring a steady supply of silver ions for the plating process.

Half-reaction at the cathode: $Ag^+ + e^- \rightarrow Ag$

Half-reaction at the anode: $Ag \rightarrow Ag^+ + e^-$



Electroplating uses electricity to coat one metal with another metal.
CC Tracy Poulsen

Some percentage of the gold and silver jewelry sold is electroplated. The connection points in electric switches are often gold plated to improve electrical conductivity and most of the chromium pieces on automobiles are chromium plated.

Lesson Summary

- Reactions in which there is a transfer of electrons is said to be an oxidation-reduction reaction or a redox reaction.
- A substance that loses electrons is said to be oxidized, and the substance the gains electrons is said to be reduced.
- Redox reactions can be used in electrochemical cells to produce electricity.
- Electrochemical cells are composed of an anode and cathode in two separate solutions. These solutions are connected by a salt bridge and a conductive wire.
- An electric current consists of a flow of charged particles.
- The electrode where oxidation occurs is called the anode and the electrode where reduction occurs is called the cathode.
- In electroplating, the object to be plated is made the cathode

Vocabulary

- Oxidation: a loss of electrons, resulting in an increased charge or oxidation number
- Reduction: gaining electrons, resulting in a decreased charge or oxidation number
- Battery: A group of two or more cells that produces an electric current.
- Anode: The electrode at which oxidation occurs.
- Cathode: electrode at which reduction occurs.
- Electrochemical cell: An arrangement of electrodes and ionic solutions in which a redox reaction is used to make electricity (a.k.a., a battery)
- Electrolysis: A chemical reaction brought about by an electric current.
- Electroplating: A process in which electrolysis is used as a means of coating an object with a layer of metal.

Further Reading / Supplemental Links

- Wikipedia: How Batteries Work
  http://en.wikipedia.org/wiki/Battery_(electricity)#How_batteries_work
- Battery simulation: Galvanic Cells
  http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/galvan5.swf

- 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 Busy Electron***.
- The principles of electrochemical cell design are explained through batteries, sensors, and a solar-powered car. The Busy Electron (http://www.learner.org/vod/vod_window.html?pid=807)
- *Dictionary of Scientific and Technical Terms,* Sybil P. Parker, Editor in Chief, McGraw-Hill, 1994.
- http://academic.pgcc.edu/~ssinex/E_cells.pdf
- http://academic.pgcc.edu/~ssinex/E_cells.pdf
- http://en.wikipedia.org
- Therese Forsythe. *Illustration of an Endothermic Reaction.*. CC-BY-SA.
- Richard Parsons. *Sodium chloride crystal, NaCl.*. CC-BY-SA.

9.2: Review Questions

Identify the element oxidized and the element reduced in the following reactions:

1) $Cu + 2\ HCl \rightarrow CuCl_2 + H_2$
2) $2\ H_2 + O_2 \rightarrow 2\ H_2O$
3) $2\ Al + 6\ HCl \rightarrow 2AlCl_3 + 3\ H_2$
4) $Zn + 2\ HCl \rightarrow ZnCl_2 + H_2$
5) $CuCl_2 + Al \rightarrow AlCl_3 + Cu$

6) What is electricity or electrical current?

7) What is an anode?

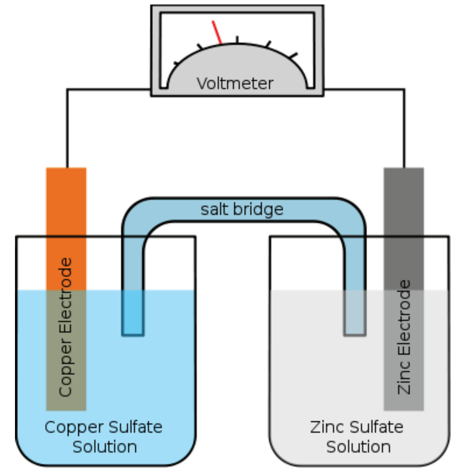
8) What is a cathode?

Consider the battery diagrammed. The overall reaction that occurs in this battery is:

$$Cu^{2+} + Zn \rightarrow Cu + Zn^{2+}$$

Answer each of the following about the battery:

9) What is being oxidized?
10) What is being reduced?
11) Label the anode. How do you know this is the anode?
12) Label the cathode. How do you know this is the cathode?
13) Which direction will electrons flow through the wire and voltmeter? (From left to right or right to left?)



Obtained from: http://en.wikipedia.org/wiki/File:Galvanic_Cell.svg

Chapter 10: Nuclear Changes

10.1: Discovery of Radioactivity

Objectives

- List the most common emissions from naturally radioactive nuclei.
- Compare the energy released per gram of matter in nuclear reactions to that in chemical reactions.
- Express the relationship between nuclear stability and the nuclei's binding energy per nucleon ratio.

Introduction

No one could have known in the 1800s that the discovery of the fascinating science and art form of photography would eventually lead to the splitting of the atom. The basis of photography is the fact that visible light causes certain chemical reactions. If the chemicals are spread thinly on a surface but protected from light by a covering, no reaction occurs. When the covering is removed, however, light acting on the chemicals causes them to darken. With millions of cameras in use today we do not think of it as a strange phenomenon, but at the time of its discovery photography was a strange and wonderful thing.

Even stranger was the discovery by Roentgen, that radiation other than visible light could expose photographic film. He found that film wrapped in dark paper would react when x-rays went through the paper and struck the film.

Becquerel and Radioactivity

When Becquerel heard about Roentgen's discovery, he wondered if his fluorescent minerals would give the same x-rays. Becquerel placed some of his rock crystals on top of a well-covered photographic plate and sat them in the sunlight. The sunlight made the crystals glow with a bright fluorescent light, but when Becquerel developed the film he was very disappointed. He found that only one of his minerals, a uranium salt, had fogged the photographic plate. He decided to try again, and this time, to leave them out in the sun for a longer period of time. Fortunately, the weather didn't cooperate and Becquerel had to leave the crystals and film stored in a drawer for several cloudy days. Before continuing his experiments, Becquerel decided to check one of the photographic plates to make sure the chemicals were still good. To his amazement, he found that the plate had been exposed in spots where it had been near the uranium containing rocks and some of these rocks had not been exposed to sunlight at all. In later experiments, Becquerel confirmed that the radiation from the uranium had no connection with light or fluorescence, but the amount of radiation was directly proportional to the concentration of uranium in the rock. Becqueral had discovered **radioactivity**.

Bequerel's radioactive rock sitting on an envelope containing photographic film.

The Curies and Radium

One of Becquerel's assistants, a young Polish scientist named Maria Sklowdowska (to become Marie Curie after she married Pierre Curie), became interested in the phenomenon of radioactivity. With her husband, she decided to find out if chemicals other than uranium were radioactive. The Austrian government was happy to send the Curies a ton of pitchblende from the mining region of Joachimstahl because it was waste material that had to be disposed of anyway. The Curies wanted the pitchblend because it was the residue of uranium mining. From the ton of pitchblend, the Curies separated 0.10 g of a previously unknown element, radium, in the form of the compound, radium chloride. This radium was many times more radioactive than uranium.

Marie Sklodowska before she moved to Paris. (Source: http://chemistry.about.com/od/historyofchemistry/ig/Pictures-of-Famous-Chemists/Marie.--bs.htm. Public Domain)

By 1902, the world was aware of a new phenomenon called radioactivity and of new elements which exhibited natural radioactivity. For this work, Becquerel and the Curies shared the 1903 Nobel Prize and for subsequent work, Marie Curie received a second Nobel Prize. She is the only person ever to receive two Nobel Prizes in science.

Further experiments provided information about the characteristics of the penetrating emissions from radioactive substances. It was soon discovered that there were three common types of radioactive emissions. Some of the radiation could pass easily through aluminum foil while some of the radiation was stopped by the foil. Some of the radiation could even pass through foil up to a centimeter thick. The three basic types of radiation were named alpha, beta, and gamma radiation. The actual composition of the three types of radiation was still not known.

Eventually, scientists were able to demonstrate experimentally that the alpha particle, α, was a helium nucleus (a particle containing two protons and two neutrons), a beta particle, β, was a high speed electron, and gamma rays, γ, were a very high energy form of light (even higher energy than x-rays).

Although these scientists didn't know it at the time, all of us are subjected to a certain amount of radiation every day. This radiation is called **background radiation** and comes from a variety of natural and artificial radiation sources. Approximately 82% of background radiation comes from natural sources. These include 1) sources in the earth, including naturally occurring radioactive elements which are incorporated in building materials and also in the human body, 2) sources from space in the form of cosmic rays, and 3) sources in the atmosphere such as radioactive radon gas released from the earth and radioactive atoms like carbon-14 produced in the atmosphere by bombardment from high-energy cosmic rays.

Unstable Nuclei May Disintegrate

A nucleus (with one exception, hydrogen-1) consists of some number of protons and neutrons pulled together in an extremely tiny volume. Since protons are positively charged and like charges repel, it is clear that protons cannot remain together in the nucleus unless there is a powerful force holding them there. The force which holds the nucleus together is generated by **nuclear binding energy**.

A nucleus with a large amount of binding energy per nucleon (proton or neutron) will be held together tightly and is referred to as stable. These nuclei do not break apart. When there is too little binding energy per nucleon, the nucleus will be less stable and may disintegrate (come apart). Such disintegrations are referred to as **natural radioactivity**. It is also possible for scientists to smash nuclear particles together and cause nuclear reactions between normally stable nuclei. These disintegrations are referred to **artificial radioactivity**. None of the elements above #92 on the periodic table occur on earth naturally . . . they are all products of artificial radioactivity (man-made).

When nuclei come apart, they come apart violently accompanied by a tremendous release of energy in the form of heat, light, and radiation. This energy comes from some of the nuclear binding energy. In nuclear changes, the energy involved comes from the nuclear binding energy. However, in chemical reactions, the energy comes from electrons moving energy levels. A typical nuclear change (such as fission) may involve millions of times more energy per atom changing compared to a chemical changes (such as burning)!

Lesson Summary

- Henri Becquerral, Marie Curie, and Pierre Curie shared the discovery of radioactivity.

Vocabulary

- Alpha particle: An alpha particle is a helium-4 nucleus, composed of 2 protons and 2 neutrons
- Beta particle: A beta particle is a high speed electron, specifically an electron of nuclear origin.
- Gamma ray: Gamma radiation is the highest energy on the spectrum of electromagnetic radiation.

10.2: Types of Radiation

Objectives

- Compare qualitatively the ionizing and penetration power of alpha particles (α), beta particles (β), and gamma rays (γ).
- Express the changes in the atomic number and mass number of a radioactive nuclei when an alpha, beta, or gamma particle is emitted
- Write nuclear equations for alpha and beta decay reactions.

Introduction

Many nuclei are radioactive; that is, they decompose by emitting particles and in doing so, become a different nucleus. In our studies up to this point, atoms of one element were unable to change into different elements. That is because in all other types of changes we have talked about only the electrons were changing. In these changes, the nucleus, which contains the protons which dictate which element an atom is, is changing. All nuclei with 84 or more protons are radioactive and elements with less than 84 protons have both stable and unstable isotopes. All of these elements can go through nuclear changes and turn into different elements.

Types of Radioactive Decay

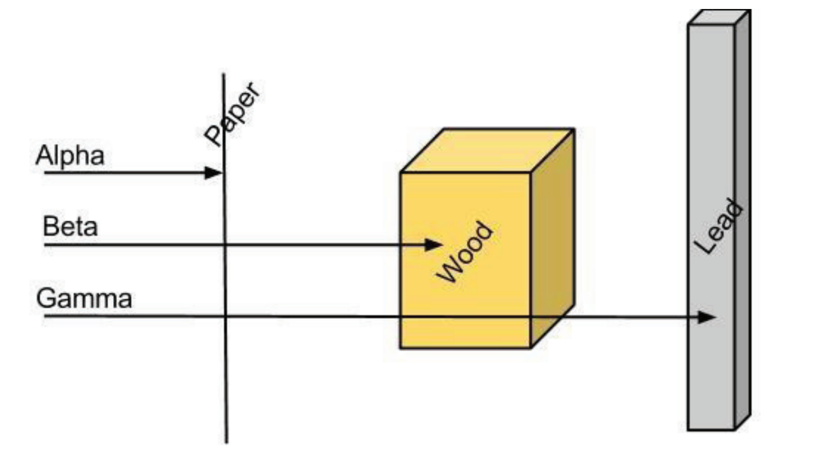
In natural radioactive decay, three common emissions occur. When these emissions were originally observed, scientists were unable to identify them as some already known particle and so named them **alpha particles** (α), **beta particles** (β), and **gamma rays** (γ) using the first three letters of the Greek alphabet. Some later time, alpha particles were identified as helium-4 nuclei, beta particles were identified as electrons, and gamma rays as a form of electromagnetic radiation like x-rays except much higher in energy and even more dangerous to living systems.

The Ionizing and Penetration Power of Radiation

With all the radiation from natural and man-made sources, we should quite reasonably be concerned about how all the radiation might affect our health. The damage to living systems is done by radioactive emissions when the particles or rays strike tissue, cells, or molecules and alter them. These interactions can alter molecular structure and function; cells no longer carry out their proper function and molecules, such as DNA, no longer carry the appropriate information. Large amounts of radiation are very dangerous, even deadly. In most cases, radiation will damage a single (or very small number) of cells by breaking the cell wall or otherwise preventing a cell from reproducing.

The ability of radiation to damage molecules is analyzed in terms of what is called **ionizing power.** When a radiation particle interacts with atoms, the interaction can cause the atom to lose electrons and thus become ionized. The greater the likelihood that damage will occur by an interaction is the ionizing power of the radiation.

Much of the threat from radiation is involved with the ease or difficulty of protecting oneself from the particles. How thick of a wall do you need to hide behind to be safe? The ability of each type of radiation to pass through matter is expressed in terms of **penetration power.** The more material the radiation can pass through, the greater the penetration power and the more dangerous they are. In general, the greater mass present the greater the ionizing power and the lower the penetration power.



Alpha particles can be blocked with paper. Beta particles penetrate through paper, but are blocked by wood. Gamma particles are capable of penetrating paper and wood but can be blocked by a few inches of lead.
CC Tracy Poulsen

Comparing only the three common types of ionizing radiation, alpha particles have the greatest mass. Alpha particles have approximately four times the mass of a proton or neutron and approximately 8,000 times the mass of a beta particle. Because of the large mass of the alpha particle, it has the highest ionizing power and the greatest ability to damage tissue. That same large size of alpha particles, however, makes them less able to penetrate matter. They collide with molecules very quickly when striking matter, add two electrons and become a harmless helium atom. Alpha particles have the least penetration power and can be stopped by a thick sheet of paper or even a layer of clothes. They are also stopped by the outer layer of dead skin on people. This may seem to remove the threat from alpha particles but only from external sources. In a situation like a nuclear explosion or some sort of nuclear accident where radioactive emitters are spread around in the environment, the emitters can be inhaled or taken in with food or water and once the alpha emitter is inside you, you have no protection at all.

Beta particles are much smaller than alpha particles and therefore, have much less ionizing power (less ability to damage tissue), but their small size gives them much greater penetration power. Most resources say that beta particles can be stopped by a one-quarter inch thick sheet of aluminum. Once again, however, the greatest danger occurs when the beta emitting source gets inside of you.

Gamma rays are not particles but a high energy form of electromagnetic radiation (like x-rays except more powerful). Gamma rays are energy that has no mass or charge. Gamma rays have tremendous penetration power and require several inches of dense material (like lead) to shield them. Gamma rays may pass all the way through a human body without striking anything. They are considered to have the least ionizing power and the greatest penetration power.

Particle	*Symbol*	*Mass*	*Penetrating Power*	*Ionizing Power*	*Shielding*
Alpha	α	4 amu	Very Low	Very High	Paper Skin
Beta	β	1/2000 amu	Intermediate	Intermediate	Aluminum
Gamma	γ	0 (energy only)	Very High	Very Low	2 inches lead

The safest amount of radiation to the human body is zero. It isn't possible to be exposed to no ionizing radiation so the next best goal is to be exposed to as little as possible. The two best ways to minimize exposure is to limit time of exposure and to increase distance from the source.

Alpha Decay

The nuclear disintegration process that emits alpha particles is called alpha decay. An example of a nucleus that undergoes alpha decay is uranium-238. The alpha decay of U-238 is

$$^{238}_{92}U \rightarrow {}^{4}_{2}He + {}^{234}_{90}Th$$

In this nuclear change, the uranium atom ($^{238}_{92}U$) transmutated into an atom of thorium ($^{234}_{90}Th$) and, in the process, gave off an alpha particle. Look at the symbol for the alpha particle: $^{4}_{2}He$. Where does an alpha particle get this symbol? The bottom number in a nuclear symbol is the number of protons. That means that the alpha particle has two protons in it which were lost by the uranium atom. The two protons also have a charge of +2. The top number, 4, is the mass number or the total of the protons and neutrons in the particle.

Because it has 2 protons, and a total of 4 protons and neutrons, alpha particles must also have two neutrons. Alpha particles always have this same composition: two protons and two neutrons.

These types of equations are called nuclear equations. When writing these, there are some general rules that will help you:

- The sum of the mass numbers (top numbers) on the reactant side equal the sum of the atomic numbers on the product side
- The atomic numbers (bottom numbers) on the two sides of the reaction will also be equal.

In this equation,

mass number: 238 = 4 + 234.

atomic number: 92 = 2 + 90

Another alpha particle producer is thorium-230.

$$^{230}_{90}Th \rightarrow {}^{4}_{2}He + {}^{226}_{88}Ra$$

Confirm that this equation is correctly balanced by adding up the reactants' and products' atomic and mass numbers. Also, note that because this was an alpha reaction, one of the products is the alpha particle, $^{4}_{2}He$.

Beta Decay

Another common decay process is beta particle emission, or beta decay. A beta particle is simply a high energy electron that is emitted from the nucleus. It may occur to you that we have a logically difficult situation here. Nuclei do not contain electrons and yet during beta decay, an electron is emitted from a nucleus. At the same time that the electron is being ejected from the nucleus, a neutron is becoming a proton. It is tempting to picture this as a neutron breaking into two pieces with the pieces being a proton and an electron. That would be convenient for simplicity, but unfortunately that is not what happens; more about this at the end of this section. For convenience sake, though, we will treat beta decay as a neutron splitting into a proton and an electron. The proton stays in the nucleus, increasing the atomic number of the atom by one. The electron is ejected from the nucleus and is the particle of radiation called beta.

In order to insert an electron into a nuclear equation and have the numbers add up properly, an atomic number and a mass number had to be assigned to an electron. The mass number assigned to an electron is zero (0) which is reasonable since the mass number is the number of protons plus neutrons and an electron contains no protons and no neutrons. The atomic number assigned to an electron is negative one (-1), because that allows a nuclear equation containing an electron to balance atomic numbers. Therefore, the nuclear symbol representing an electron (beta particle) is

$$^{0}_{-1}e \text{ or } {}^{0}_{-1}\beta$$

Thorium-234 is a nucleus that undergoes beta decay. Here is the nuclear equation for this beta decay.

$$^{234}_{90}Th \rightarrow {}^{0}_{-1}e + {}^{234}_{91}Pa$$

Note that both the mass numbers and the atomic numbers add up properly:

mass number: 234 = 0 + 234

atomic number: 90= -1 + 91

The mass numbers of the original nucleus and the new nucleus are the same because a neutron has been lost, but a proton has been gained and so the sum of protons plus neutrons

remains the same. The atomic number in the process has been increased by one since the new nucleus has one more proton than the original nucleus. In this beta decay, a thorium-234 nucleus has become a protactinium-234 nucleus. Protactinium-234 is also a beta emitter and produces uranium-234.

$$^{234}_{91}Pa \rightarrow {}^{0}_{-1}e + {}^{234}_{92}U$$

Once again, the atomic number increases by one and the mass number remains the same; confirm that the equation is correctly balanced.

Gamma Radiation

Frequently, gamma ray production accompanies nuclear reactions of all types. In the alpha decay of U-238, two gamma rays of different energies are emitted in addition to the alpha particle.

$$^{238}_{92}U \rightarrow {}^{4}_{2}He + {}^{234}_{90}Th + 2{}^{0}_{0}\gamma$$

Virtually all of the nuclear reactions in this chapter also emit gamma rays, but for simplicity the gamma rays are generally not shown. Nuclear reactions produce a great deal more energy than chemical reactions. Chemical reactions release the difference between the chemical bond energy of the reactants and products, and the energies released have an order of magnitude of $1\text{x}10^3$ kJ/mol. Nuclear reactions release some of the binding energy and may convert tiny amounts of matter into energy. The energy released in a nuclear reaction has an order of magnitude of $1\text{x}10^8$ kJ/mol. That means that nuclear changes involve almost a million times more energy per atom than chemical changes!!! That's a lot.

Example: Complete the following nuclear reaction by filling in the missing particle.

$$^{210}_{86}Rn \rightarrow {}^{4}_{2}He + ?$$

Solution: This reaction is an alpha decay. We can solve this problem one of two ways:

Solution 1: When an atom gives off an alpha particle, its atomic number drops by 2 and its mass number drops by 4 leaving: $^{206}_{84}Po$. We know the symbol is Po, for polonium, because this is the element with 84 protons on the periodic table.

Solution 2: Remember that the mass numbers on each side must total up to the same amount. The same is true of the atomic numbers. 4+?

Mass numbers: 210 = 4 + ?

Atomic numbers: 86 = 2 + ?

We are left with $\boldsymbol{^{206}_{84}Po}$

Example: Write each of the following nuclear reactions.

a) Carbon-14, used in carbon dating, decays by beta emission.

b) Uranium-238 decays by alpha emission.

Solution:

a) Beta particles have the symbol $^{0}_{-1}e$. Emitting a beta particle causes the atomic number to increase by 1 and the mass number to not change. We get atomic number and symbols for elements using our periodic table. We are left with the following reaction:

$$\boldsymbol{^{14}_{6}C \rightarrow {}^{0}_{-1}e + {}^{14}_{7}N}$$

b) Alpha particles have the symbol $^{4}_{2}He$. Emitting an alpha particle causes the atomic number to decrease by 2 and the mass number to decrease by 4. We are left with:

$$\boldsymbol{^{238}_{92}U \rightarrow {}^{4}_{2}He + {}^{234}_{90}Th}$$

Decay Series

The decay of a radioactive nucleus is a move toward becoming stable. Often, a radioactive nucleus cannot reach a stable state through a single decay. In such cases, a series of decays will occur until a stable nucleus is formed. The decay of U-238 is an example of this. The U-238 decay series starts with U-238 and goes through fourteen separate decays to finally reach a stable nucleus, Pb-206. There are similar decay series for U-235 and Th-232. The U-235 series ends with Pb-207 and the Th-232 series ends with Pb-208.

Several of the radioactive nuclei that are found in nature are present there because they are produced in one of the radioactive decay series. That is to say, there may have been radon on the earth at the time of its formation, but that original radon would have all decayed by this time. The radon that is present now is present because it was formed in a decay series.

Lesson Summary

- A nuclear reaction is one that changes the structure of the nucleus of an atom.
- The atomic numbers and mass numbers in a nuclear equation must be balanced.
- Protons and neutrons are made up of quarks.
- The two most common modes of natural radioactivity are alpha decay and beta decay.
- Most nuclear reactions emit energy in the form of gamma rays.

Vocabulary

- Alpha decay: Alpha decay is a common mode of radioactive decay in which a nucleus emits an alpha particle (a helium-4 nucleus).
- Beta decay: Beta decay is a common mode of radioactive decay in which a nucleus emits beta particles. The daughter nucleus will have a higher atomic number than the original nucleus.
- Quark: particles that form one of the two basic constituents of matter. Various species of quarks combine in specific ways to form protons and neutrons, in each case taking exactly three quarks to make the composite particle.

Further Reading / Supplementary Links

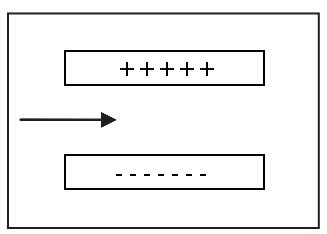
- Tutorial: Vision Learning: Nuclear Chemistry: http://visionlearning.com/library/module_viewer.php?mid=59&l=&c3=
- Radioactive Decay: http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/radioa7.swf

10.2: Review Questions

Put the letter of the matching phrase on the line preceding the number.

1) alpha particle	a. high energy electromagnetic radiation
2) beta particle	b. a high speed electron
3) gamma ray	c. a helium nucleus

4) Because alpha, beta, and gamma particles have different charges, they will interact differently through an electric field. If the arrow represents the original path of alpha, beta, and gamma particles, complete arrows (3 total) showing the movement of each type of radiation through the electric field. Label each of your arrows.



5) When a nucleus gives off a beta particle, what effect does this have on the number of protons, neutrons, and total number of particles in the nucleus?
6) When a nucleus gives off an alpha particle, what effect does this have on the number of protons, neutrons, and total number of particles in the nucleus?
7) Which of the three common emissions from radioactive sources requires the heaviest shielding?
8) Which type of radiation is most dangerous to living tissues?

Complete the following nuclear equations by supplying the missing particle.

9) $^{208}_{84}Po \rightarrow {}^{4}_{2}He + ?$
10) $? \rightarrow {}^{210}_{84}Po + {}^{0}_{-1}e$

Write each of the following nuclear reactions as a nuclear equation.

11) The radioactive isotope iodine-131 (used to study thyroid function) decays by beta emission.
12) Thorium-230 decays by alpha emission.
13) ^{234}Th decays by beta emission.
14) Hydrogen-3 decays by beta emission.
15) The alpha decay of radon-198.
16) The beta decay of uranium-237.

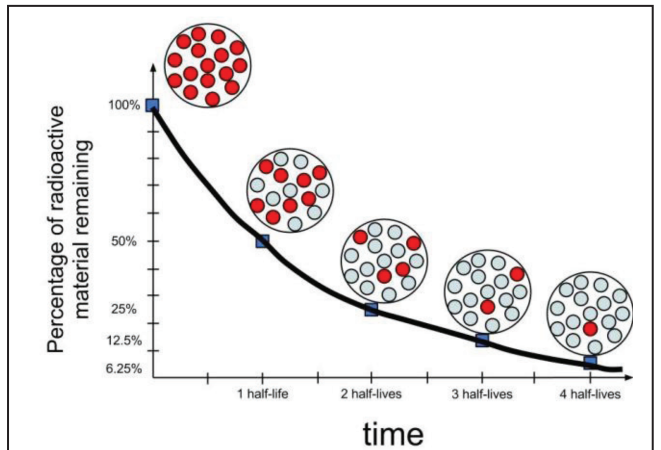
10.3: Half-life & Rate of Radioactive Decay

Objectives

- Describe what is meant by the term half-life and what factors affect half-life.
- Calculate the amount of radioactive material that will remain after an integral number of half-lives.
- Find the half-life of an isotope given graphical or other data
- Describe how carbon-14 is used to determine the age of carbon containing objects.

Rate of Radioactive Decay

During natural radioactive decay, not all atoms of an element are instantaneously changed to atoms of another element. The decay process takes time and there is value in being able to express the rate at which a process occurs. A useful concept is **half-life,** which is the time required for half of the starting material to change or decay. Half-lives can be calculated from



After one half-life, half of the radioactive atoms have decayed. After another half-life, half of the remaining atoms have changed. This pattern continues, with half of the atoms changing over each half-life time.
CC Tracy Poulsen

measurements on the change in mass of a nuclide and the time it takes to occur. The only thing we know is that in the time of that substance's half-life, half of the original nuclei will disintegrate. Although chemical changes were sped up or slowed down by changing factors such as temperature, concentration, etc, these factors have no effect on half-life. Each radioactive isotope will have its own unique half-life that is independent of any of these factors.

The half-lives of many radioactive isotopes have been determined and they have been found to range from extremely long half-lives of 10 billion years to extremely short half-lives of fractions of a second.

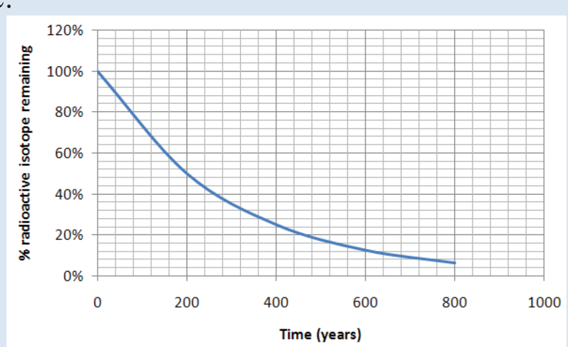
Table of Selected Half-lives

Element	Mass Number	Half-Life	Element	Mass Number	Half-Life
Uranium	238	4.5 billion years	Californium	251	800years
Neptunium	240	1 hour	Nobelium	254	3seconds
Plutonium	243	5 hours	Carbon	14	5730years
Americium	246	25 minutes	Carbon	16	0.74seconds

The quantity of radioactive nuclei at any given time will decrease to half as much in one half-life. For example, if there were 100g of Cf-251in a sample at some time, after 800 years, there would be 50 g of Cf-251 remaining and after another 800 years (1600 years total), there would only be 25 g remaining.

Remember, the half-life is the time it takes for half of your sample, not matter how much you have, to remain. Each half-life will follow the same general pattern as Cf-251. The only difference is the length of time it takes for half of a sample to decay.

Example: Using the graph, what is the half-life of an isotope that produces the following graph of decay over time:



Solution: We know that the half-life is the time it takes for half of a sample to change. How long did it take for half of our isotope to change? It took approximately 200 years for 100% of our sample leave only 50% (half of the original amount) to remain. **The half-life is 200 years.**
*Notice that after another 200 years (400 years total), 25% remains (half of 50%)

Look carefully at the graph in the previous example. All types radioactive decay makes a graph of the same general shape. The only difference is the scale and units of the x-axis, as the half-life time will be different.

Example: If there are 60 grams of Np-240 present, how much Np-240 will remain after 4 hours? (Np-240 has a half-life of 1 hour)

Solution: Np-240 with a half-life of only 1 hour.

Amount of Np-240 present	Amount of time passed
60 g	0 (this is the amount before any time has passed)
30 g	1 hour (1 half-life)
15 g	2 hours (2 half-lives)
7.5 g	3 hours
3.75 g	4 hours

After 4 hours, only **3.75 g** of our original 60 g sample would remain the radioactive isotope Np-240.

Example: A sample of Ac-225 originally contained 80 grams and after 50 days only 2.55 grams of the original Ac-225 remain. What is the half-life of Ac-225?

Solution: We are going to tackle this problem similar to the last problem. The difference is that we are looking for the half-life time. Let's set up a similar table, though:

Amount of Ac-225 present	Amount of time passed
80 g	0
40 g	1 half-life
20 g	2 half-lives
10 g	3 half-lives
5 g	4 half-lives
2.5 g	5 half-lives

We know that 50 days is the same as 5 half-lives. Therefore, 1 half-life is 10 days. The half-life of Ac-225 is 10 days.

Radioactive Dating

An ingenious application of half-life studies established a new science of determining ages of materials by half-life calculations. For geological dating, the decay of U-238 can be used. The half-life of U-238 is $4.5\text{x}10^9$ years. The end product of the decay of U-238 is Pb-206. After one half-life, a 1.00 gram sample of uranium will have decayed to 0.50 grams of U-238 and 0.43 grams of Pb-206. By comparing the amount of U-238 to the amount of Pb-206 in a sample of uranium mineral, the age of the mineral can be estimated. Present day estimates for the age of the Earth's crust from this method is at 4 billion years.

Organic material (material made from things that were once living, such as paper and fabric) is radioactively dated using the long-lived nuclide of carbon, carbon-14. This method of determining the age of organic material (or once living materials) was given the name radiocarbon dating. The carbon dioxide consumed by living systems contains a certain concentration of ${}^{14}CO_2$. When an organism dies, the acquisition of carbon-14 stops but the decay of the C-14 in the body continues. As time goes by, the ratio of C-14 to C-12 decreases at a rate determined by the half-life of C-14. Using half-life equations, the time since the organism died can be calculated. These procedures have been used to determine the age of organic artifacts and determine, for instance, whether art works are real or fake.

Lesson Summary
- The half-life of an isotope is used to describe the rate at which the isotope will decay and give off radiation.
- Using the half-life, it is possible to predict the amount of radioactive material that will remain after a given amount of time.
- C-14 dating procedures have been used to determine the age of organic artifacts. Its half-life is approximately 5700 years.

Vocabulary
- Background radiation: Radiation that comes from environment sources including the earth's crust, the atmosphere, cosmic rays, and radioisotopes. These natural sources of radiation account for the largest amount of radiation received by most people.
- Half-life: The half-life of a radioactive substance is the time interval required for a quantity of material to decay to half its original value.

Further Reading / Supplementary Links
- The Dating Game: Carbon-dating tutorial: http://www.pbs.org/wgbh/nova/first/radiocarbon.html
- Half-life: http://www.colorado.edu/physics/2000/isotopes/radioactive_decay3.html

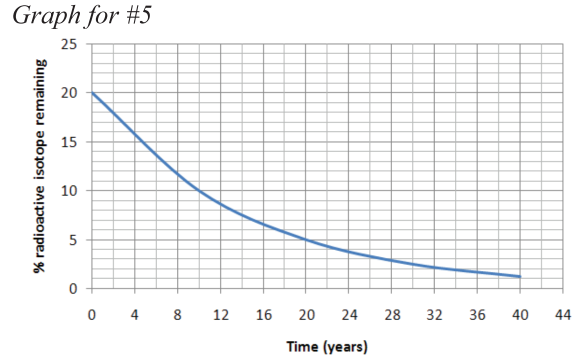
10.3: Questions
1. The half-life of radium-226 is about 1600 years. How many grams of a 2.00 gram sample will remain after 4800 years?

2. Sodium-24 has a half-life of about 15 hours. How much of a 16.0 gram sample of sodium-24 will remain after 60.0 hours?

3. A radioactive isotope decayed from 24.0 grams to 0.75 grams in 40.0 years. What is the half-life of the isotope?

4. The half-life of C-14is about 5,700 years. An organic relic is found to contain C-14 and C-12 in a ratio that is about one-eighth as great as the ratio in the atmosphere, meaning only 1/8 of the original amount of C-14 remains. What is the approximate age of the relic?

Graph for #5



5. What is the half-life of the isotope which produced the following graph over time?

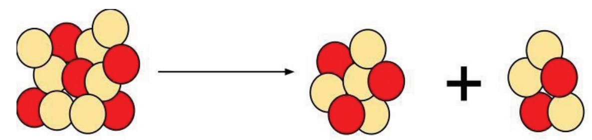
10.4: Applications of Nuclear Changes

Objectives
- Define and give examples of fission and fusion.
- Classify nuclear reactions as fission or fusion.
- List some medical uses of nuclear energy.

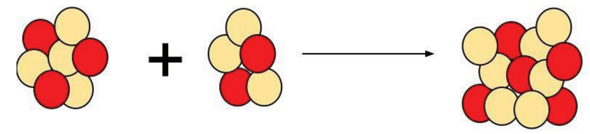
Introduction

Nuclei that are larger than iron-56 may undergo nuclear reactions in which they break up into two or more smaller nuclei. These reactions are called **fission** reactions.

Conversely, nuclei that are smaller than iron-56 become larger nuclei in order to be more stable. These nuclei undergo a nuclear reaction in which smaller nuclei join together to form a larger nucleus. Such nuclear reactions are called **fusion** reactions.



Nuclear fission, in which one large nucleus is split into two or more smaller nuclei.



Nuclear fusion happens when two small nuclei combine to make one larger nucleus.
CC Tracy Poulsen

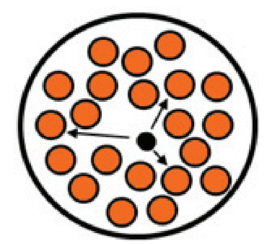
Fission and Chain Reactions

In both fission and fusion, large amounts of energy are given off in the form of heat, light, and gamma radiation. Nuclear fission was discovered in the late 1930s when U-235 nuclides were bombarded with neutrons and were observed to split into two smaller-mass nuclei.

$${}^{1}_{0}n + {}^{235}_{92}U \rightarrow {}^{141}_{56}Ba + {}^{92}_{36}Kr + 3{}^{1}_{0}n$$

The products shown are only one of many sets of products from the disintegration of a U-235 nucleus. Over 35 different elements have been observed in the fission products of U-235.

When a neutron strikes a U-235 nucleus and the nucleus captures a neutron, it undergoes fission producing two lighter nuclei and three free neutrons. The production of the free neutrons makes it possible to have a self-sustaining fission process – a nuclear **chain reaction.** If at least one of the neutrons goes on to cause another U-235 disintegration, the fission will be self-sustaining.



In the fission of U-235, free neutrons are released which may strike the nucleus of other atoms causing them to fission. If enough U-235 is present, a self-sustaining chain reaction may result. Source: Richard Parsons. CC-BY-SA

Fission Reactors

Fission reactions can be used in the production of electricity if we control the rate at which the fission occurs. The great majority of all electrical generating systems (whether coal burning power plants, hydroelectric plants or nuclear power plants) all follow a reasonably simple design. The electricity is produced by spinning a coil of wire inside a magnetic field. When a fluid (air, steam, water) is forced through the pipe, it spins the fan blades which in turn spin the axle. To generate electricity, the axle of a turbine is attached to the loop of wire in a generator. When a fluid is forced through the turbine, the fan blades turn, the turbine axle turns, and the loop of wire inside the generator turns, thus generating electricity.

The essential difference in various kinds of electrical generating systems is the method used to spin the turbine. For a wind generator, the turbine is a windmill. In a geothermal generator, steam from a geyser is forced through the turbine. In hydroelectric generating plants, water falling over a dam passes through the turbine and spins it. In fossil fuel (coal, oil, natural gas) generating plants, the fossil fuel is burned and the heat is used to boil water into steam and then the steam passes through the turbine and makes it spin. In a fission reactor generating plant, a fission reaction is used to boil the water into steam and the steam passes through the turbine to make it spin. Once the steam is generated by the fission reaction, a nuclear power plant is essentially the same as a fossil fuel plant.

Naturally occurring uranium is composed almost totally of two uranium isotopes. It contains more than 99% uranium-238 and less than 1% uranium-235. It is the uranium-235, however, that is **fissionable** (will undergo fission). In order for uranium to be used as fuel in a fission reactor, the percentage of uranium-235 must be increased, usually to about 3%. (Uranium in which the U-235 content is more than 1% is called **enriched uranium**.)

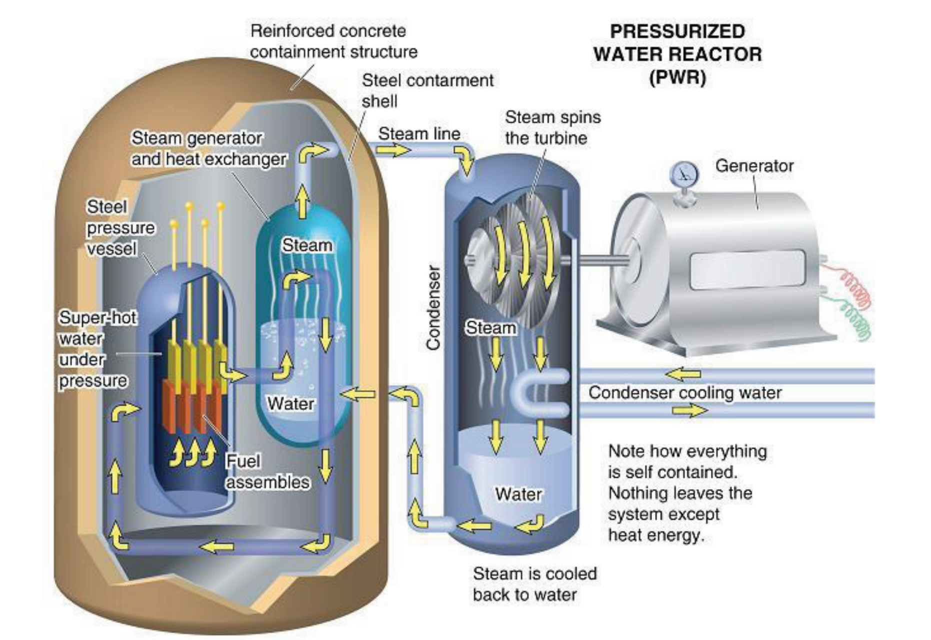
Once the supply of U-235 is acquired, it is placed in a series of long cylindrical tubes called fuel rods**.** These fuel cylinders are bundled together with **control rods** made of neutron-absorbing material. The amount of U-235 in all the fuel rods taken together is adequate to carry on a chain reaction but is less than the critical mass. (In the United States, all public nuclear power plants contain less than a critical mass of U-235 and therefore, could never produce a nuclear explosion.) The amount of heat generated by the chain reaction is controlled by the rate at which the nuclear reaction occurs. The rate of the nuclear reaction is dependent on how many neutrons are emitted by one U-235 nuclear disintegration *and* strike a new U-235 nucleus to cause another disintegration. The purpose of the control rods is to absorb some of the neutrons and thus stop them from causing further disintegrations. The control rods can be raised or lowered into the fuel rod bundle. When the control rods are lowered all the way into the fuel rod bundle, they absorb so many neutrons that the chain reaction essentially stops. When more heat is desired, the control rods are raised so they catch fewer neutrons, the chain reaction speeds up and more heat is generated. The control rods are operated in a fail-safe system so that power is necessary to hold them up; and during a power failure, gravity will pull the control rods down into shut off position.

U-235 nuclei can capture neutrons and disintegrate more efficiently if the neutrons are moving slower than the speed at which they are released. Fission reactors use a moderator surrounding the fuel rods to slow down the neutrons. Water is not only a good coolant but also a good moderator so a common type of fission reactor has the fuel core submerged in a huge pool of water.

You can follow the operation of an electricity-generating fission reactor in the figure. The reactor core is submerged in a pool of water. The heat from the fission reaction heats the

water and the water is pumped into a heat exchanger container where the heated water boils the water in the heat exchanger. The steam from there is forced through a turbine which spins a generator and produces electricity. After the water passes through the turbine, it is condensed back to liquid water and pumped back to the heat exchanger.

In the United States, heavy opposition to the use of nuclear energy was mounted in the late 1960s and early 1970s. Every environmentalist organization in the US opposed the use of nuclear energy and the constant pressure from environmentalist groups brought increased public fear and therefore, opposition. This is not true today; at least one environmental leader has published a paper in favor of nuclear powered electricity generation.



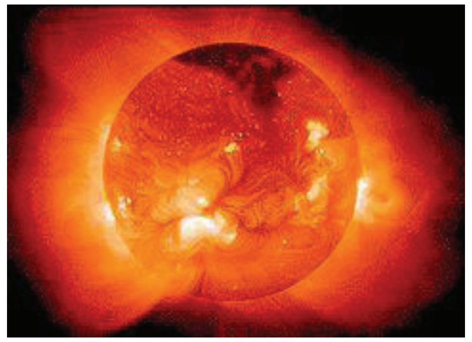
A fission reactor.

In 1979, a reactor core meltdown at Pennsylvania's Three Mile Island nuclear power plant reminded the entire country of the dangers of nuclear radiation. The concrete containment structure (six feet thick walls of reinforced concrete), however, did what it was designed to do – prevent radiation from escaping into the environment. Although the reactor was shut down for years, there were no injuries or deaths among nuclear workers or nearby residents. Three Mile Island was the only serious accident in the entire history of 103 civilian power plants operating for 40 years in the United States. There has never been a single injury or death due to radiation in any public nuclear power plant in the U.S. The accident at Three Mile Island did, however, frighten the public so that there has not been a nuclear plant built in the U.S. since the accident.

The 103 nuclear power plants operating in the U.S. deliver approximately 19.4% of American electricity with zero greenhouse gas emission. There are 600 coal-burning electric plants in the US delivering 48.5% of American electricity and producing 2 billion tons of CO_2 annually, accounting for 40% of U.S. CO_2 emissions and 10% of global emissions. These coal burning plants also produce 64% of the sulfur dioxide emissions, 26% of the nitrous oxide emissions, and 33% of mercury emissions.

Fusion

Nuclear reactions, in which two or more lighter-mass nuclei join together to form a single nucleus, are called **fusion** reactions or nuclear fusions. Of particular interest are fusion reactions in which hydrogen nuclei combine to form helium. Hydrogen nuclei are positively charged and repel each other. The closer the particles come, the greater is the force of repulsion. In order for fusion reactions to occur, the hydrogen nuclei must have extremely high kinetic energies so the velocities can overcome the forces of repulsion. These kinetic energies only occur at extreme temperatures such as those that occur in the cores of the sun and other stars. Nuclear fusion is the power source for the stars where the necessary temperature to ignite the fusion reaction is provided by massive gravitational pressure. In stars more massive than our sun, fusion reactions involving carbon and nitrogen are possible. These reactions produce more energy than hydrogen fusion reactions.


The energy that comes from the sun and other stars is produced by fusion. (Source: http://commons.wikimedia.org/wiki/File:Sun-in-X-ray. Public Domain)

Intensive research is now being conducted to develop fusion reactors for electricity generation. The two major problems slowing up the development is finding a practical means for generating the intense temperature needed and developing a container than won't melt under the conditions of a fusion reaction. Electricity-producing fusion reactors are still a distant dream.

Uses of Nuclear Radiation

It is unfortunate that when the topics of radioactivity and nuclear energy come up, most thoughts probably go to weapons of war. The second thought might be about the possibility of nuclear energy contributing to the solution of the energy crisis. Nuclear energy, however, has many applications beyond bombs and the generation of electricity. Radioactivity has huge applications in scientific research, several fields of medicine both in terms of imaging and in terms of treatment, industrial processes, some very useful appliances, and even in agriculture.

The field of nuclear medicine has expanded greatly in the last twenty years. A great deal of the expansion has come in the area of imaging. This section will focus on nuclear medicine involving the types of nuclear radiation introduced in this chapter. The x-ray imaging systems will not be covered.

Radioiodine (I-131) Therapy involves imaging and treatment of the thyroid gland. The thyroid gland is a gland in the neck that produces two hormones that regulate metabolism. In some individuals, this gland becomes overactive and produces too much of these hormones. The treatment for this problem uses radioactive iodine (I-131) which is produced for this purpose in research fission reactors or by neutron bombardment of other nuclei.

The thyroid gland uses iodine in the process of its normal function. Any iodine in food that enters the bloodstream is usually removed by, and concentrated in the thyroid gland. When a patient suffering from an overactive thyroid swallows a small pill containing

radioactive iodine, the I-131 is absorbed into the bloodstream just like non-radioactive iodine and follows the same process to be concentrated in the thyroid. The concentrated emissions of nuclear radiation in the thyroid destroy some of the gland's cells and control the problem of the overactive thyroid.

Smaller doses of I-131 (too small to kill cells) are also used for purposes of imaging the thyroid. Once the iodine is concentrated in the thyroid, the patient lays down on a sheet of film and the radiation from the I-131 makes a picture of the thyroid on the film. The half-life of iodine-131 is approximately 8 days so after a few weeks, virtually all of the radioactive iodine is out of the patient's system. During that time, they are advised that they will set off radiation detectors in airports and will need to get special permission to fly on commercial flights.

Positron Emission tomography or PET scan is a type of nuclear medicine imaging. Depending on the area of the body being imaged, a radioactive isotope is either injected into a vein, swallowed by mouth, or inhaled as a gas. When the radioisotope is collected in the appropriate area of the body, the gamma ray emissions are detected by a PET scanner (often called a gamma camera) which works together with a computer to generate special pictures, providing details on both the structure and function of various organs. PET scans are used to:

- detect cancer
- determine the amount of cancer spread
- assess the effectiveness of treatment plans
- determine blood flow to the heart muscle
- determine the effects of a heart attack
- evaluate brain abnormalities such as tumors and memory disorders
- map brain and heart function

External Beam Therapy (EBT) is a method of delivering a high energy beam of radiation to the precise location of a patient's tumor. These beams can destroy cancer cells and with careful planning, NOT kill surrounding cells. The concept is to have several beams of radiation, each of which is sub-lethal, enter the body from different directions. The only place in the body where the beam would be lethal is at the point where all the beams intersect. Before the EBT process, the patient is three-dimensionally mapped using CT scans and x-rays. The patient receives small tattoos to allow the therapist to line up the beams exactly. Alignment lasers are used to precisely locate the target. The radiation beam is usually generated with a linear accelerator. EBT is used to treat the following diseases as well as others:

- breast cancer
- colorectal cancer
- head and neck cancer
- lung cancer
- prostate cancer

Lesson Summary

- Naturally radioactive elements exist in the earth and are either alpha or beta emitters.
- Artificial transmutation of elements can be accomplished by bombarding the nuclei of some elements with alpha or subatomic particles.
- Nuclear fission refers to the splitting of atomic nuclei.

- Nuclear fusion refers to the joining together to two or more smaller nuclei to form a single nucleus.
- The fission of U-235 or Pu-239 is used in nuclear reactors.
- Nuclear radiation also has many medical uses.

Vocabulary
- Chain reaction: A multi-stage nuclear reaction that sustains itself in a series of fissions in which the release of neutrons from the splitting of one atom leads to the splitting of others.
- Critical mass: The smallest mass of a fissionable material that will sustain a nuclear chain reaction at a constant level.
- Fission: A nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy.
- Fusion: A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy.
- Control rods : Control rods are made of chemical elements capable of absorbing many neutrons and are used to control the rate of a fission chain reaction in a nuclear reactor.

Further Reading / Supplementary Links
- A short animation of nuclear fission can be viewed at http://www.classzone.com/cz/books/woc_07/resources/htmls/ani_chem/chem_flash/popup.html?layer=act&src=qtiwf_act129.1.xml
- A short animation of nuclear fusion can be viewed at http://www.classzone.com/cz/books/woc_07/resources/htmls/ani_chem/chem_flash/popup.html?layer=act&src=qtiwf_act130.1.xml
- *Nuclear Power VS. Other Sources of Power,* Neil M. Cabreza, Department of Nuclear Engineering, University of California, Berkeley, NE-161 Report. Available at http://www.nuc.berkeley.edu/thyd/ne161/ncabreza/sources.
- *Chemistry, A Modern Course*, Chapter 28: Nuclear Chemistry, Robert C. Smoot, Jack Price, and Richard G. Smith, Merrill Publishing Co., 1987.
- www.hps.org/publicinformation/ate/cat10.html
- www.sciencemag.org/cgi/content/full/309/5732/233
- www.hrd.qut.edu.au/toolkit/Faqs/radiation.jsp
- www.radiationnetwork.com/RadiationNetwork.htm
- http://www.iaea.org/NewsCenter/Features/Chernobyl-15
- http://www.doh.wa.gove/ehp/rp/factsheets-pd/fs10
- http://www.radscihealth.org/rsh/About
- http://nrc.gov/reading-rm/doc-collections
- http://www.world-nuclear.org/info/Chernobyl
- http://www.iaea.org/NewsCenter/Features/Chernobyl-15
- http://www.nuclearweaponarchive.org/Russia/Tsarbomba.html
- http://www.atomicarchive.com/effects/index.shtml
- http://www.en.wikipedia.org

10.4: Review Questions
1) What is fission?
2) What is fusion?
3) Compared to ordinary chemical reactions (such as burning wood), how much energy is given off in nuclear reactions?
4) What is the primary physical difference between a nuclear electricity generating plant and a coal-burning electricity generating plant?
5) What do the control rods in a nuclear reactor do and how do they do it?
6) Is it possible for a nuclear explosion to occur in a nuclear reactor? Why or why not?

For each of the following, indicate which type of nuclear change is used (fission, fusion, or nuclear decay/radiation):

7) Nuclear power plants
8) Biological tracers
9) Energy from the stars
10) Cancer treatment
11) Nuclear warheads

10.5: Big Bang Theory

Objectives
- Describe the statement "Big Bang Theory"
- Give evidence for the Big Bang Theory

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 best available measurements the big bang occurred about 13.75 billion years ago.

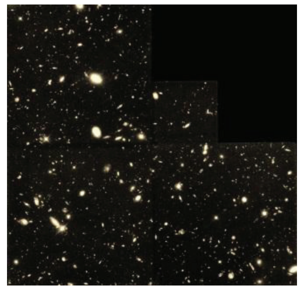
According to the theory the Universe would have cooled sufficiently to allow energy to be converted into subatomic particles (protons, neutrons, and electrons and many other particles). While protons and neutrons would have formed the first atomic nuclei only a few minutes after the Big Bang, it would then have taken thousands of years for electrons loose enough energy to form neutral atoms. The first element produced would be hydrogen. Giant clouds of these primordial elements would then form stars and galaxies. Other elements were formed by fusion 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 theory of the expansion of the Universe and proposed his 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 recession of the nebulae (later shown to be galaxies) was 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 1900's. 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 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. 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 remains in the lightest of the elements, hydrogen and helium. The heavier elements, from helium to iron were formed from fusion within stars. 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 transuranium (artificial or man-made) elements have been formed during the 20th century.

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

Further Reading / Supplementary Links

- http://www.en.wikipedia.org
- 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)

10.5: Review Questions

1) What is red shift? What causes it to occur?
2) What does redshift indicate?
3) How old is the universe, according to the Big Bang Theory?
4) Why is the abundance of hydrogen and helium so important in accepting Big Bang Theory?
5) What evidence exists that the Big Bang did occur? How do these evidences support the theory?
6) Earth (and the other inner planets) contains large amounts of elements heavier than carbon. Where did these elements come from?
7) 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?

Unit 3: Gases

11.1: Gases and Kinetic Theory

Objectives

- Compare the properties of gases, liquids and solids
- Convert between units of volume, pressure, and pressure
- State the relationship between temperature and kinetic energy

Introduction

The **Kinetic Molecular Theory** allows us to explain the existence of the three phases of matter: solid, liquid, and gas. In addition, it helps explain the physical characteristics of each phase and how phases change from one to another. The Kinetic Molecular Theory is essential for the explanations of gas pressure, compressibility, diffusion, and mixing. Our explanations for reaction rates and equilibrium also rest on the concepts of the Kinetic-Molecular Theory.

Gases are tremendously compressible, can exert massive pressures, expand nearly instantaneously into a vacuum, and fill every container they are placed in regardless of size. All of these properties of gases are due to their molecular arrangement.

Volume of Gases

In dealing with gases, we lose the meaning of the word "full." A glass of water may be 1/4 full or 1/2 full or full, but a container containing a gaseous substance is always full. The same amount of gas will fill a quart jar, or a gallon jug, a barrel, or a house. The gas molecules separate farther from each other and spread out uniformly until they fill whatever container they are in. Gases can be compressed to small fractions of their original volume and expand to fill virtually any volume. If gas molecules are pushed together to the point that they touch, the substance would then be in the liquid form. One method of converting a gas to a liquid is to cool it and another method is to compress it.

The two most common ways of expressing volume are using mL and L. You will need to be able to convert between these two units. The relationship is as follows:

1000 mL = 1 L

Pressure of Gases

The constant random motion of the gas molecules causes them to collide with each other and with the walls of their container. These collisions of gas molecules with their surroundings exert a pressure on the surroundings. When you blow up a balloon, the air particles inside the balloon push against the elastic sides, the walls of the balloon are pushed outward and kept firm. This pressure is produced by air molecules pounding on the inside walls of the balloon.

There are three two units of pressure commonly used in chemistry. Pressure is commonly measured on a device called a monometer, similar to the barometer which a meteorologist uses. Pressures in monometers are typically recorded in units of millimeters of mercury, abbreviated mmHg. Pressure is defined as the force exerted divided by the area over which the force is exerted.

$$pressure = \frac{force}{area}$$

The air molecules in our atmosphere exert pressure on every surface that is in contact with air. The air pressure of our atmosphere at sea level is approximately 15 pounds/in^2. This pressure is unnoticed, because the air is not only outside the surfaces but also inside allowing the atmospheric air pressure to be balanced. The pressure exerted by our atmosphere will become quickly noticed, however, if the air is removed or reduced inside an object. A common demonstration of air pressure makes use of a one-gallon metal can. The can has a few drops of water placed inside and is then heated to boiling. The water inside the can vaporizes and expands to fill the can pushing the air out. The lid is then tightly sealed on the can. As the can cools, the water vapor inside condenses back to liquid water leaving the inside of the can with a lack of air molecules. As the water vapor condenses to liquid water, the air pressure outside the can slowly crushes the can flat.

Barometer

People, of course, also have atmospheric pressure pressing on them. An averaged sized person probably has a total force exerted on them from the atmosphere in excess of 25,000 pounds. Fortunately, people also have air inside them to balance the force. A device to measure atmospheric pressure, the barometer, was invented in 1643 by an Italian scientist named Evangelista Torricelli (1608 – 1647) who had been a student of Galileo. Torricelli's barometer was constructed by filling a glass tube, open at one end and closed at the other, with liquid mercury and then inverting the tube in a dish of mercury.

The mercury in the tube fell to a height such that the difference between the surface of the mercury in the dish and the top of the mercury column in the tube was 760 millimeters. The volume of empty space above the mercury in the tube was a vacuum. The explanation for why the mercury stays in the tube is that there are air molecules pounding on the surface of the mercury in the dish and there are no air molecules pounding on the top of the mercury in the tube. The weight of the mercury in the tube divided by the area of the opening in the tube is exactly equal to the atmospheric pressure.

The height to which the mercury is held would only be 760 millimeters when air pressure is normal and at sea level. The atmospheric pressure changes due to weather conditions and the height of the mercury in the barometer will change with it. Atmospheric pressure also varies with altitude. Higher altitudes have lower air pressure because the air is "thinner" – fewer air molecules per unit volume. In the mountains, at an altitude of 9600 feet, the normal atmospheric pressure will only support a mercury column of 520 mmHg.

For various reasons, chemistry has many different units for measuring and expressing gas pressure. You will need to be familiar with most of them so you can convert them into preferred units. Because instruments for measuring pressure often contain a column of mercury, the most commonly used units for pressure are based on the height of the mercury column that the gas can support. The original unit in chemistry for gas pressure was *mmHg* (millimeters of mercury). Standard atmospheric pressure at sea level is 760. mmHg. This unit is something of a problem because while it is a pressure unit, it looks a lot like a length unit. Students, in particular, occasionally leave off the *Hg* and then it definitely appears to be a length unit. To eliminate this problem, the unit was given another name. It was called the torr in honor of Torricelli. 760 torr is exactly the same as 760 mmHg. For certain work, it became convenient to express gas pressure in terms of multiples of normal atmospheric pressure at sea level and so the unit atmosphere (atm) was introduced. The conversion you need to know between various pressure units are:

1.00 atm = 760. mmHg = 760. torr

Example: Convert 425 mmHg to atm.
Solution The conversion factor is 760. mmHg = 1.00 atm $425\ mmHg \times \frac{1\ atm}{760\ mmHg} = \mathbf{0.559\ atm}$

This example shows how to perform this conversion using dimensional analysis. If you are the memorizing type, you can just memorize that to convert from mmHg to atm you must divide by 760.

Gas Temperature and Kinetic Energy

Kinetic energy is the energy of motion and therefore, all moving objects have kinetic energy. The mathematical formula for calculating the kinetic energy of an object is KE=1/2 mv^2, where *m* is the mass and *v* is the velocity of the object or particle. This physics formula applies to all objects in exactly the same way whether we are talking about the moon moving in its orbit, a baseball flying toward home plate, or a gas molecule banging around in a bottle. All of these objects have kinetic energy and their kinetic energies can all be calculated with the same formula. The kinetic energy of a molecule would be calculated in exactly this same way. You should note that if the mass of an object is doubled while its velocity remains the same, the kinetic energy of the object would also be doubled. If, on the other hand, the velocity is doubled while the mass remains the same, the kinetic energy would be quadrupled because of the square in the formula.

When you measure the temperature of a group of molecules, what you are actually measuring is their average kinetic energy. They are the same thing but expressed in different units. The formula for this relationship is KE_{ave}=3/2RT where R is the gas constant and T is the absolute temperature, measured in Kelvin. When a substance is heated, the average kinetic energy of the molecules is increased. Since the mass of the molecules cannot be increased by heating, it is clear that the velocity of the molecules is increasing.

Remember, the motion of molecules is related to their temperature. If you think of the average kinetic energy of a group of molecules and temperature measured in degrees Kelvin, the relationship is a direct proportion. That means that if the temperature, in Kelvin, is doubled the kinetic energy of the particles is also doubled. It is absolutely vital that you keep in mind that the mathematical relationship between the temperature and the average kinetic energy of molecules only exists when the temperature is expressed in the Kelvin scale. In order for the direct proportion to exist, the molecules must have zero kinetic energy when the temperature is zero. The temperature at which molecular motion stops is 0 K (-273°C). It is surely apparent to you that molecules do NOT have zero kinetic energy at 0°C. Balloons and automobile tires do not go flat when the outside temperature reaches 0°C. If temperature is measured in Kelvin degrees, then the average kinetic energy of a substance at 100 K is exactly double the average kinetic energy of a substance at 50 K. *Make sure all the*

calculations you do dealing with the kinetic energy of molecules is done with Kelvin temperatures.

Some important principles can be derived from this relationship:

1. All gases at the same temperature have the same average kinetic energy.
2. Heavier gases must move more slowly in order to have the same kinetic energy as lighter gases.

Example: If molecules of H_2, O_2, and N_2 are all placed in the same container at the same temperature, which molecules will have the greatest velocity?
Solution: Because they are at the same temperature, they will have the same energy. However, lighter particles must move faster in order to have the same kinetic energy. We must, therefore, look at their masses. Use your periodic table: Mass of H_2 = 2(1.008 g/mol) = 2.016 g/mol Mass of O_2 = 2(16.00 g/mol) = 32.00 g/mol Mass of N_2 = 2(14.01 g/mol) = 28.02 g/mol Because H_2 is the lightest, it must have the greatest velocity in order to have the same energy (the same temperature) as the other gases.

Section Summary

- The collisions between molecules are perfectly elastic. The phrase "perfectly elastic collision" comes from physics and means that kinetic energy is conserved in collisions.
- The molecules of an ideal gas have no attraction or repulsion for each other.
- At any given moment, the molecules of a gas have different kinetic energies. We deal with this variation by considering the average kinetic energy of the molecules. The average kinetic energy of a group of molecules is measured by temperature.
- Molecules of a gas are so far apart, on average, that the volume of the molecules themselves in negligible compared to the volume of the gas.
- Molecular collisions with container walls cause the gas to exert pressure.
- Because of the molecular motion of molecules, they possess kinetic energy at all temperatures above absolute zero.
- Temperature is directly proportional to the average kinetic energy of gas molecules.
- Lighter gases will have higher velocities than heavier gases, at the same temperature and pressure.
- In the Kelvin scale, 0 K means the particles have no kinetic energy. Doubling the temperature in Kelvin doubles the kinetic energy of particles.
- Real gases tend to deviate from ideal gases at high pressures and low temperatures, as the attractive forces between molecules and the volume of gas molecules becomes significant

Vocabulary

- **Kelvin temperature**: The absolute temperature scale where 0 K is the theoretical absence of all thermal energy (no molecular motion).
- **Kinetic energy**: Kinetic energy is the energy a body possesses due to it motion, KE=1/2mv^2.

- **Kinetic theory:** used to explain how properties of gases
- **Pressure:** a measure of the force with which gas particles collide with the walls of their containers
- **Temperature:** a measurement of the kinetic energy of particles

11.2: Gas Laws

Objectives

- Predict effect on pressure, volume, or temperature if one of the other variables are changed.
- Solve problems using the combined gas law

?maybe do combined first and get other gas laws after

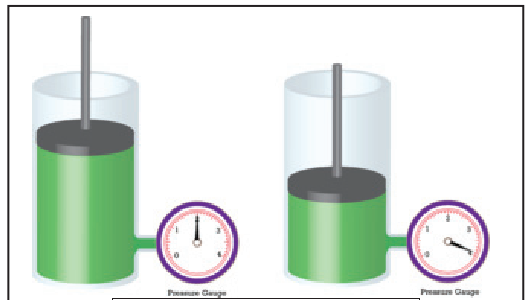
Introduction

Gases are often characterized by their volume, temperature, and pressure. These characteristics, however, are not independent of each other. Gas pressure is dependent on the force exerted by the molecular collisions and the area over which the force is exerted. The force exerted by the molecular collisions is dependent on the absolute temperature and so forth. The relationships between these characteristics can be determined both experimentally and logically from their mathematical definitions.

The gas laws are mathematical relationships that exist for gases between the volume, pressure, temperature, and quantity of gas present. They were determined experimentally over a period of 100 years. They are logically derivable from our present day definitions of pressure, volume, and temperature.

Boyle's Law: Pressure vs. Volume

The relationship between the pressure and volume of a gas was first determined experimentally by an Irish chemist named Robert Boyle (1627-1691). The relationship between the pressure and volume of a gas is commonly referred to as **Boyle's Law**.



PV Data

Trial	Volume	Pressure
1	8.0 L	1.0 atm
2	4.0 L	2.0 atm
3	2.0 L	4.0 atm
4	1.0 L	8.0 atm

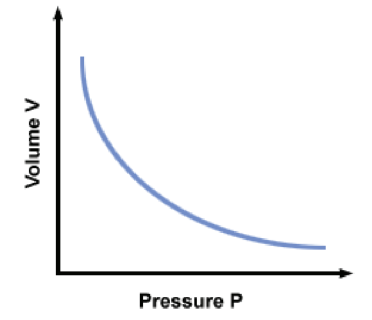
Volume and pressure data for a gas sample.

When we wish to observe the relationship between two variables, it is absolutely necessary to keep all other variables constant so that the change in one variable can be directly related to the change in the other. Therefore, when the relationship between gas volume and gas pressure is investigated, the quantity of gas and its temperature must be held constant so these factors do not contribute to any observed changes.

You may have noticed that when you try to squeeze a balloon, the resistance to squeezing is greater as the balloon becomes smaller. That is, the pressure inside the balloon becomes greater when the volume is reduced. This phenomenon can be studied more carefully with an apparatus like that in Figure 9. This is a cylinder tightly fitted with a piston that can be raised or lowered. There is also a pressure gauge fitted to the cylinder so that the gas pressure inside

the cylinder can be measured. The amount of gas inside the cylinder cannot change and the temperature of the gas is not allowed to change.

In the picture on the right, the volume of the gas is 4.0 L and the pressure exerted by the gas is 2.0 atm. If the piston is pushed down to decrease the volume of the gas to 2.0 L, the pressure of the gas is found to increase to 4.0 atm. The piston can be moved up and down to positions for several different volumes and the pressure of the gas read at each of the volumes.



The relationship between volume and pressure is an inverse relationship.

We might note from casual observation of the data that doubling volume is associated with the pressure being reduced to half and if we move the piston to cause the pressure to double, the volume is halved. The data show that the relationship is an inverse relationship, meaning that as volume increases the pressure decreases. The opposite is also true.

Boyle's Law can be summarized in the following equation:

$$P_1V_1 = P_2V_2$$

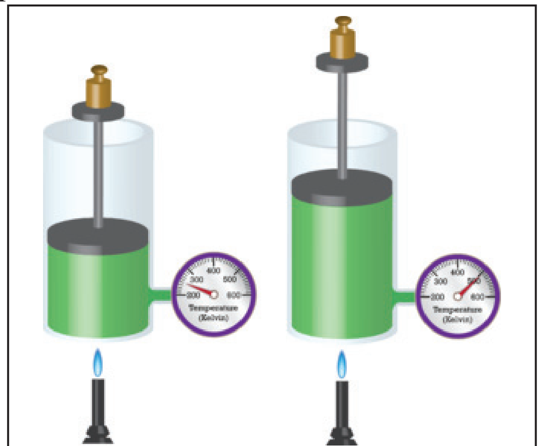
Where:

- P_1=the initial pressure
- V_1=the initial volume
- P_2=the final pressure
- V_2=the final volume

For this equation, the units used for pressure are unimportant, as long as both pressures have the same unit (either mmHg or atm) and each volume has the same unit (either mL or L).

Charles's Law: Temperature and Volume

The relationship between the volume and temperature of a gas was investigated by a French physicist, Jacques Charles (1746-1823). (As a piece of trivia, Charles was also the first person to fill a large balloon with hydrogen gas and take a solo balloon flight.) The relationship between the volume and temperature of a gas is often referred to as **Charles's Law.**

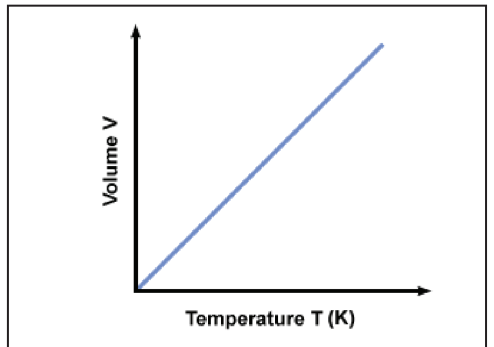


The picture on the left shows the volume of a sample of gas at 250. K and the picture on the right shows the volume when the temperature has been raised to 500 K.

An apparatus that can be used to study the relationship between the temperature and volume of a gas is shown in the picture to the right. Once again, we have a sample of gas trapped inside a cylinder so no gas can get in or out. Thus we have a constant mass of gas. We also have a mass set on top of a moveable piston to keep a constant force pushing against the gas. This guarantees that the gas pressure in the cylinder will be constant because if the pressure inside increases, the piston will be pushed up expanding inside volume until the inside pressure becomes equal to outside pressure again.

Similarly, if the inside pressure decreases, the outside pressure will push the cylinder down, decreasing volume, until the two pressures again become the same. This system guarantees constant gas pressure inside the cylinder.

This relationship is a direction relationship. If the temperature, in Kelvin, doubles, so does the volume. This relationship would also be expected when we recognize that we are increasing the total force of molecular collisions with the walls by raising the temperature and the only way to keep the pressure from increasing is to increase the area over which that larger force is exerted. This mathematical relationship is known as a direct proportionality. When one variable is increased, the other variable also increases by exactly the same factor. An equation to show how these values are related is given by:



Volume and temperature (in K) are directly related.

$$\frac{\boldsymbol{V_1}}{\boldsymbol{V_2}} = \frac{\boldsymbol{T_1}}{\boldsymbol{T_2}}$$

This relationship is ONLY true if the temperature is measured in Kelvin. However, the units of volume are irrelevant, as long as the two volumes are measured in the same units.

Gay-Lussac's Law: Temperature and Pressure

The relationship between temperature and pressure was investigated by the French chemist, Joseph Gay-Lussac (1778-1850). In an apparatus used for this investigation, the cylinder does not have a moveable piston because it is necessary to hold the volume constant as well as the quantity of gas. This apparatus allows us to alter the temperature of a gas and measure the pressure exerted by the gas at each temperature.

After a series of temperatures and pressures have been measured, a data table like the others can be produced.

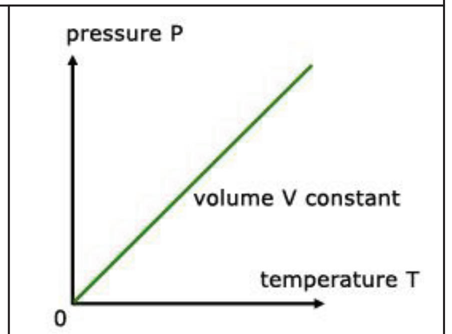
Pressure vs. Temperature Data

Trial	Temperature	Pressure
1	200. K	600. mmHg
2	300. K	900 mmHg
3	400. K	1200 mmHg
4	500. K	1500 mmHg

Temperature and pressure data. Note that if the temperature doubles from 200. K to 400. K, the pressure also doubles.

Temperature and pressure are also directly related, meaning that if the temperature, in Kelvin, doubles, so does the pressure. This relationship is also logical since by increasing temperature, we are increasing the force of molecular collision and keeping the area over which the force is exerted constant requires that the pressure increases.

$$\frac{\boldsymbol{P_1}}{\boldsymbol{P_2}} = \frac{\boldsymbol{T_1}}{\boldsymbol{T_2}}$$



Pressure and temperature (in K) are directly related in Gay-Lussac's Law

Standard Temperature and Pressure (STP)

It should be apparent by now that expressing a quantity of gas simply by stating its volume is totally inadequate. Ten liters of oxygen gas could contain any mass of oxygen from 4000g to 0.50g depending on the temperature and pressure of the gas. Chemists have found it useful to have a standard temperature and pressure with which to express gas volume. The standard conditions of temperature and pressure (STP) were chosen to be 0°C (273 K) and 1.00 atm (760 mmHg). You will commonly see gas volumes expressed as 1.5L at STP. Once you know the temperature and pressure conditions of a volume of gas, you can calculate the volume at other conditions and you can also calculate the mass of the gas if you know the formula.

The Combined Gas Law

Boyle's Law shows how the volume of a gas changes when its pressure is changed (temperature held constant) and Charles's Law shows how the volume of a gas changes when the temperature is changed (pressure held constant). Is there a formula we can use to calculate the change in volume of a gas if both pressure and temperature change? The answer is "yes", we can use a formula that combines Boyle's Law and Charles's Law.

This equation is most commonly written in the from shown below and is known as the Combined Gas Law.

$$\frac{\boldsymbol{P_1 V_1}}{\boldsymbol{T_1}} = \frac{\boldsymbol{P_2 V_2}}{\boldsymbol{T_2}}$$

As in the other laws, when solving problems with the combined gas law, temperatures must always be in Kelvin. The units for pressure and volume may be any appropriate units but the units for each value of pressure must be the same and the units for each value of volume must be the same.

Another interesting point about the combined gas law is that all the other gas laws (Charles', Gay-Lussac's, and Boyle's) can be derived from this equation. To do this, you simply cancel out the variable that was held constant in the reaction. For example, temperature is constant in Boyle's Law. If you cancel the temperature's out of Boyle's Law, you get:

$$P_1 V_1 = P_2 V_2$$

Although the other equations are not as obvious, the same method can be used to derive the other equations. If you are able to derive the other equations, you will not have to memorize them.

Example: A sample of gas has a volume of 400. liters when its temperature is 20.°C and its pressure is 300. mmHg. What volume will the gas occupy at STP?

Solution:

Step 1: Identify the given information & check units. Temperature must be in Kelvin. Volume units must match and pressure units must match.

P_1=300 mmHg

V_1=400. L

T_1=293 K (remember, ALL temperatures must be in Kelvin)

P_2=760 mmHg (standard pressure)

V_2=?

T_2=273 K
Step 2: Solve the combined gas law for the unknown variable.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(300\ mmHg)(400L)}{293\ K} = \frac{(760\ mmHg)V_2}{273K}$$

V_2=147 L

Example: A sample of gas occupies 1.00 under standard conditions. What temperature would be required for this sample of gas to occupy 1.50 L and exert a pressure of 2.00 atm?

Solution:
Step 1: Identify the given information & check units. Temperature must be in Kelvin. Volume units must match and pressure units must match.
P_1=1.00 atm (standard pressure)
V_1=1.00 L
T_1=273 K (standard temperature, remember, ALL temperatures must be in Kelvin)
P_2=2.00 atm
V_2=1.50 L
T_2=?
Step 2: Solve the combined gas law for the unknown variable.

$$\frac{(1.00\ atm)(1.00L)}{273\ K} = \frac{(2.00\ atm)(1.50\ L)}{T_2}$$

T_2=819 K

Example: A sample of gas has a volume of 500.mL under a pressure of 500.mmHg. What will be the new volume of the gas if the pressure is reduced to 300.mmHg at constant temperature?

Solution:
Step 1: Identify the given information & check units. Temperature must be in Kelvin. Volume units must match and pressure units must match.
P_1=500. mmHg
V_1=500. mL
P_2=300. mmHg
V_2=?
Temperature is constant, so it cancels out of the combined gas law.
Step 2: Solve the combined gas law for the unknown variable. (Or, recognize this is Boyle's Law and start with that equation.)

$$P_1V_1 = P_2V_2$$

$$(500\ mmHg)(500\ mL) = (300\ mmHg)V_2$$

V_2=833 mL

Avogadro's Law

Avogadro's Law was known as Avogadro's hypothesis for the first century of its existence. Since Avogadro's hypothesis can now be demonstrated mathematically, it was decided that it should be called a law instead of a hypothesis. Avogadro's Law postulates

that *equal volumes of gas under the same conditions of temperature and pressure contain the same number of molecules.*

This relationship is important for a couple of reasons. It means that all gases under the same conditions behave the same way: all of these equations work equally well for carbon dioxide, helium, or a mixture of gases. Furthermore, we will be able to use this relationship again when we deal with balanced reactions. The volume of two gases at the same temperature and pressure are directly related to the number of molecules (or moles) of the gases involved in a chemical reaction.

Section Summary

- For a fixed sample of ideal gas at constant temperature, volume is inversely proportional to pressure.
- For a fixed sample of ideal gas at constant pressure, volume in directly proportional to temperature.
- For a fixed sample of ideal gas at constant volume, pressure is directly proportional to temperature.
- The volume of a mass of gas is dependent on the temperature and pressure. Therefore, these conditions must be given along with the volume of a gas.
- Standard conditions of temperature and pressure are 0°C and 1.0 atm.
- Avogadro's Law: Equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules.

Further Reading / Supplemental Links

- Section 7-6 is on the Combined Gas Law. http://www.fordhamprep.org/gcurran/sho/sho/Sections/Section31.htm
- http://en.wikipedia.org/wiki/Kinetic_theory;
- http://www.chm.davidson.edu/chemistryapplets/kineticmoleculartheory/basicconcepts.html

11.3: Ideal Gas Law

Objectives

- Solve problems using the ideal gas law, PV=nRT.

Introduction

The individual gas laws and the combined gas law all require that the quantity of gas remain constant. The Universal Gas Law (also sometimes called the Ideal Gas Law) allows us to make calculations on different quantities of gas as well.

The Universal Gas Law Constant

We have considered four laws that describe the behavior of gases: Boyle's Law, Charles's Law, Avogadro's Law, and Gay-Lussac's Law. These three relationships, which show how the volume of a gas depends on pressure, temperature, and the number of moles of gas, can be combined to form the **ideal gas law**:

$$\boldsymbol{PV = nRT}$$

Where each variable and its units are:

- P=pressure (atm)
- V=volume (L)
- n=number of moles of gas (mol)
- T=temperature (K)
- R=ideal gas constant = 0.0821 atm·L/mol·K

Up to this point in gas law calculations, we haven't worried too much about which unit you use for pressure and volume as long as the units matched. Notice that the gas constant, R, has specific units. Your units of pressure and volume must be in atm and L, respectively, because they must match the appropriate units in the constant, R. Moles, of course, always have the unit moles and temperature must always be Kelvin. You can convert the value of R into values for any set of units for pressure and volume, if you wanted, but the numerical value of R would also change.

Example: A sample of nitrogen gas, N_2, has a volume of 5.56 L at 0°C and 1.50 atm pressure. How many moles of nitrogen are present in this sample?

Solution:

Step 1: Identify the given information & check units. Temperature must be in Kelvin. Volume and pressure units must match R.

P=1.50 atm
V=5.56 L
n=?
T=273 K (must be in K)

Step 2: Solve the ideal gas law for the unknown variable.

$$PV = nRT$$

$$(1.50\ atm)(5.56L) = n\left(0.0821\frac{atm \cdot L}{mol \cdot K}\right)(273K)$$

n=0.372 mol

Example: 2.00 mol of methane gas, CH_4, are placed in a rigid 500. mL container and heated to 100.°C. What pressure will be exerted by the methane?

Solution:

Step 1: Identify the given information & check units. Temperature must be in Kelvin. Volume and pressure units must match R.

P=?
V=500 mL = 0.500 L
n=2.00 mol
T=100□ C = 373 K

Step 2: Solve the ideal gas law for the unknown variable.

$$PV = nRT$$

$$P(0.500L) = (2.00\ mol)\left(0.0821\frac{atm \cdot L}{mol \cdot K}\right)(373K)$$

P=122 atm

Section Summary

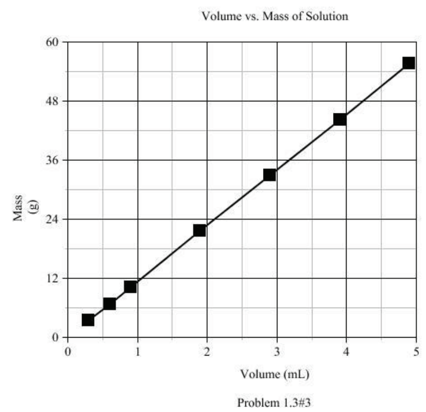
- The Universal Gas Law: PV=nRT
- At STP, one mole of any gas occupies 22.4
- The universal gas law is often used along with laboratory data to find the molar mass of an unknown substance.

Answers to Selected Problems

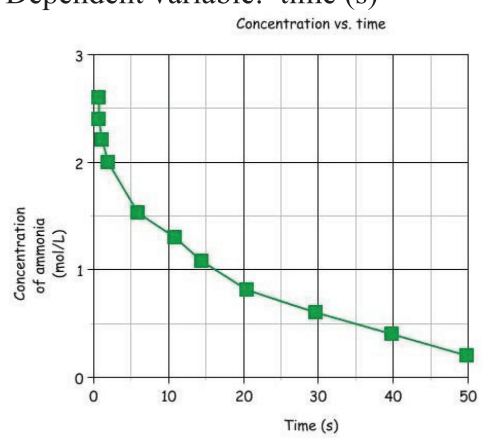
Section 1.1
1) B
2) A
3) D
4) D
5) B
6) Whether or not the plants received fertilizer
7) Growth (height) of plants
8) Amount of sun, amount of water, type of plant (corn)

Section 1.2
1) D
2) A
3) D
4) A
5) A
6) D
7) D
8) F

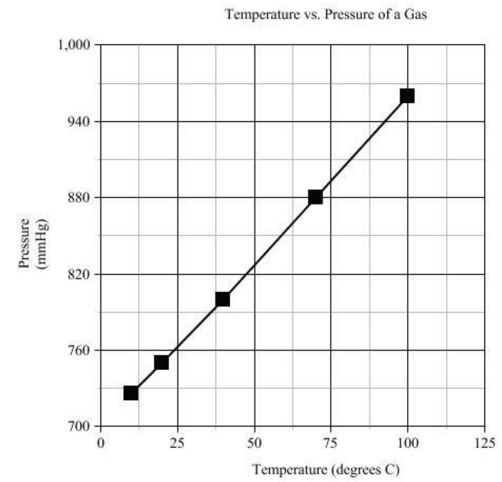
Section 1.3
1) The independent variable is the label of the first column. The dependent variable is the label of the last column(s).
2) The independent variable is the label of the x-axis or the key. The dependent variable is the label of the y-axis.
3) Exact graphs and answers may vary, but should look similar to the following
   a) Independent variable: volume of solution (mL); dependent: mass of solution (g)



   b)
   c) 11.3 g/mL
   d) 27 g
   e) 5.2 mL
4) Exact graphs and answers may vary, but should look similar to the following
   a) Independent variable: concentration of ammonia (mol/L). Dependent variable: time (s)



   b)
   c) About 47 seconds
   d) About 1.45 mol/L
5) Exact graphs and answers may vary, but should look similar to the following
   a) Independent variable: temperature (°C); dependent variable: pressure (mmHg)



   b)
   c) 2.6 mmHg/°C
   d) About 830 mmHg
   e) About 1000 mmHg

Section 2.1
1) C
2) Dalton had experimental evidence to support his claims. Democritus did not.
3) No! Inaccurate theories give scientists an idea to build from and a way to test other ideas and develop experiments. Most current ideas are adaptations of previous ideas.

Section 2.2
1) F
2) T
3) T
4) F
5) D
6) A
7) E
8) C
9) B
10) Nucleus
11) Repelled by…attracted to
12) Location of positive mass in atoms
13) C
14) If the particles hit the positive central mass they would bounce off. If they missed the central positive part, they would go straight through.

Section 2.3
1) T
2) F
3) F
4) T
5) B
6) C
7) A
8) T
9) T
10) T
11) F
12) F
13) Ru
14) Zn
15) Kr
16) B
17) E
18) A
19) D
20) C
21) $^{58}_{26}Fe$
22) $^{19}_{9}F$
23) $p^+=2$, $n^0=2$, $e^-=2$
24) $p^+=11$, $n^0=12$, $e^-=11$
25) $p^+=1$, $n^0=0$, $e^-=1$
26) $p^+=26$, $n^0=29$, $e^-=26$
27) $p^+=17$, $n^0=20$, $e^-=17$
28) $p^+=5$, $n^0=6$, $e^-=5$
29) $p^+=92$, $n^0=146$, $e^-=92$
30) $p^+=92$, $n^0=143$, $e^-=92$

Section 2.4
1) 63.55 amu
2) 35.49 amu

Section 2.5
1) Red
2) As the energy of a wave increases, frequency increases. As the energy of a wave increases, the wavelength decreases.
3) The wave on the left has more energy, because it has a shorter wavelength.
4) Radio, infrared, visible, UV, gamma

5) Red, orange, yellow, green, blue, violet

Section 2.6

1) Quantized means to have specific amounts of energy. Bohr said electrons can have only specific amounts of energy and are, therefore, quantized.
2) Because each element has a different spectrum, you can use it to identify which elements are present
3) The sun gives off the specific pattern of light unique to helium. No other element produces that pattern of light.
4) Electrons gain energy and move to higher energy levels. When electrons lose energy they move to lower energy levels.
5) The electrons give off the extra energy as light.
6) Each element has different possible energy levels which its electrons can occupy, so there are different possible “drops” electrons can make, producing different spectra.
7) Electrons start at ground state (lowest energy level possible). When they are given energy as light, heat, or electricity, the electrons may move up to a higher energy level (excited state). The electrons will drop back to lower energy levels releasing the extra energy as a photon (or piece of light).

Section 2.7

1) 2
2) p
3) d
4) 10
5) 2p
6) 3
7) Al
8) 4d
9) 4s
10) $1s^2\ 2s^2\ 2p^6\ 3s^2$
11) $1s^2\ 2s^2\ 2p^3$
12) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^1$
13) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^2$
14) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6$
15) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 3d^{10}\ 6p^6$
16) $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^6\ 5s^2\ 4d^{10}\ 5p^6\ 6s^2\ 4f^{14}\ 3d^{10}\ 6p^6\ 7s^2\ 5f^{14}$
17) $[He]\ 2s^2\ 2p^5$
18) $[Ne]\ 3s^2\ 3p^1$
19) $[Ar]\ 4s^2\ 3d^2$
20) $[Ar]4s^2\ 3d^{10}\ 4p^3$
21) $[Kr]5s^1$
22) $[He]\ 2s^2\ 2p^2$

Section 3.1

1) Mendeleev first put the elements in order from lightest mass to heaviest mass. Then he put elements with similar properties in the same group.
2) Mendeleev left room for undiscovered elements, he didn’t force elements into groups which didn’t have similar properties even if the mass didn’t follow his original pattern
3) Yes, unlike other methods of organization. As new elements were discovered there was room for them in Mendeleev’s table.
4) By looking how the properties such as melting point, density, etc, changed as he went down a group/family or across a row of his periodic table, he predicted what numbers would fit the pattern.
5) Sometimes the next heaviest element didn’t fit according to properties in the next available place. He either traded the order of the neighboring element (such as what he did for I and Te) or he left blank spaces to put elements in the appropriate group (such as leaving holes where Ga and Ge currently are placed).

6) Each element has a different number of protons, and elements are now arranged in order of increasing atomic number instead of increasing atomic mass.
7) There are more elements, the table is in order of increasing atomic number instead of mass, the family of noble gases has been added, the table has been turned sideways from its original form.

Section 3.2

1) Nonmetal
2) Nonmetal
3) Nonmetal
4) Metal
5) Metal
6) Nonmetal
7) Metal
8) Metalloid
9) Nonmetal
10) Metal
11) Nonmetal
12) Metal
13) Metalloid
14) Mercury has properties of metals (such as being malleable, lustrous, conductivity, ductile, etc) and bromine has properties of nonmetals (such as brittle, insulator, etc). Even though they are both liquids, their other properties place them as metal for mercury and nonmetal for bromine.

Section 3.3

1) 5
2) 3
3) 5
4) 5
5) 1
6) 1
7) 3
8) 3
9) 8
10) 8

Section 3.4

1) C
2) B
3) C
4) Halogen
5) Noble gas
6) Alkaline earth metal
7) Alkali metal
8) Halogens
9) Halogens
10) Noble gas
11) Transition metals
12) Alkaline earth metals
13) Alkali metals

Section 3.5

1) B
2) D
3) A
4) Ga
5) K
6) Ti
7) Iodine has electrons in a higher energy level further from the nucleus than bromine.
8) B, Al, Ga, In, Tl
9) C, Ga, Sn
10) The energy required to remove the electron furthest from the nucleus
11) Na, Mg, S, Ar
12) The relative attraction for electrons in a bond (how hard an atom pulls on electrons in a bond)
13) Li
14) Na
15) K
16) Mg

Section 4.1

1) Compound
2) Mixture
3) Element
4) Compound
5) Mixture
6) Mixture

7) Compound
8) Element
9) Mixture
10) Mixture

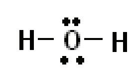
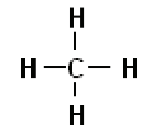
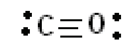
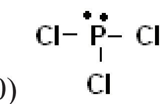
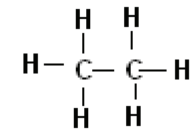
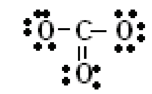
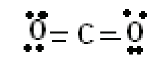
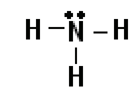
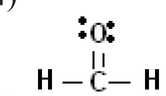
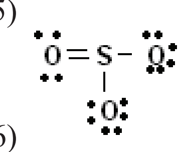
Section 4.2
1) Elements are most stable with eight valence electrons.
2) Nonmetals can form covalent bonds. Metals cannot.
3) Covalent
4) Ionic
5) Covalent
6) Ionic
7) Ionic
8) Covalent
9) Ionic
10) Metallic
11) Covalent
12) Covalent
13) Covalent
14) Metallic
15) Ionic
16) Ionic
17) Metallic
18) 1=sucrose; 2=sodium chloride, 3=zinc

Section 4.3
1) An atom or group of atoms with a charge
2) Positive, metal atoms lose electrons to form positive ions
3) Negative, nonmetals will gain electrons to form negative ions
4) -1, chloride
5) -1, bromide
6) -3, nitride
7) -2, oxide
8) +2, calcium
9) -1, fluoride
10) +2, magnesium
11) +1, lithium
12) -1, iodide
13) +1, sodium
14) +1, potassium
15) +3 aluminum
16) The names of transition ions include the charge of the ion, because they can form ions with more than one charge.
17) Copper(II)
18) Cobalt(II)
19) Cobalt(III)
20) Copper(I)
21) Nickel(II)
22) Chromium(III)
23) Iron(II)
24) Iron(III)
25) A group of atoms which together hold a charge
26) Nitrate
27) Acetate
28) Hydroxide
29) Phosphate
30) Sulfate
31) Carbonate

Section 4.4
1) $NaNO_3$
2) $Ca(NO_3)_2$
3) $Fe(NO_3)_3$
4) Na_2SO_4
5) $CaSO_4$
6) $Fe_2(SO_4)_3$
7) $NaCl$
8) $CaCl_2$
9) $FeCl_3$
10) Na_3PO_4
11) $Ca_3(PO_4)_2$
12) $FePO_4$
13) $NaOH$
14) $Ca(OH)_2$
15) $Fe(OH)_3$
16) Na_2CO_3
17) $CaCO_3$
18) $Fe_2(CO_3)_3$
19) MgS
20) $Pb(NO_3)_2$
21) Na_2O
22) $Ca(OH)_2$
23) K_2CO_3
24) $AlBr_3$
25) $Fe(NO_3)_3$

26) $FeCl_2$
27) $Cu(NO_3)_2$
28) MgO
29) CaO
30) $CuBr$
31) Al_2S_3
32) H_2CO_3
33) $KMnO_4$
34) $Cu_2Cr_2O_7$
35) $FeCl_3$
36) $FeSO_4$

Section 4.5
1) Potassium chloride
2) Magnesium oxide
3) Copper(II) sulfate
4) Sodium chloride
5) Cobalt(II) bromide
6) Magnesium fluoride
7) Nickel(II) hydroxide
8) Sodium acetate
9) Copper(II) oxide
10) Iron(II) chloride
11) Lithium chloride
12) Magnesium bromide
13) Ammonium hydroxide
14) Copper(I) oxide
15) Calcium fluoride
16) Potassium carbonate
17) Sodium chloride
18) Lead(II) oxide
19) Calcium nitrate
20) Magnesium hydroxide
21) Tin(IV) oxide

Section 4.6
1) Covalent
2) Ionic
3) Ionic
4) Covalent
5) Ionic
6) Ionic
7) 
8) 
9) 
10) 
11) 
12) 
13) 
14) 
15) 
16) 

Section 4.7
1) Tetrahedron
2) Tetrahedron
3) Tetrahedron
4) Bent
5) Trigonal pyramid
6) Tetrahedron

Section 4.8
1) In a nonpolar covalent bond, electrons are evenly shared between atoms. In a polar covalent bond electrons are not evenly shared resulting in a partial positive and partial negative side. In an ionic bond, electrons are shared at all but one atom loses electrons to another atom forming particles with full charges.
2) P-Cl is more polar than S-Cl, because there is a bigger difference in

electronegativities. P is less electronegative than S, so Cl is able to pull the electrons further from P than S making it more polar.
3) Electrons are not evenly shared
4) N, O, or F
5) Hydrogen bonding is a stronger attraction between molecules with partial charges than the attraction between polar molecules.
6) Hydrogen bonding is a strong attraction between neighboring molecules in which H is bonded to N, O, or F.
7) Polar
8) Nonpolar
9) Polar
10) Polar and hydrogen bonding
11) Polar and hydrogen bonding
12) Nonpolar
13) Ammonia
14) Water
15) Ammonia
16) Ammonia

Section 5.1
1) Based on the decimal (10) system; used internationally; units are based on physical constants
2) 1kg
3) 10 g
4) 100cg
5) 1000
6) cL
7) 100
8) Scientists need to use the same unit of measurement so they can share information, data, and calculations more effectively.
9) Meter (m)
10) Kilogram (kg)
11) Liter (L); the liter is the volume of a 10cm x 10cm x 10cm container or 1dm x 1 dm x 1 dm
12) NO! This is only 2°C (almost as cold as ice water).

Section 5.2
1) When working with really large or really small numbers and measurements
2) 4.79×10^{-5}
3) 4.26×10^{3}
4) 2.51×10^{9}
5) 2.06×10^{-3}
6) 23000
7) 0.0009156
8) .0072
9) 8,255,000
10) 7.3(EE)14
11) 6.01(EE)(-)6
12) 7.98(EE)5
13) 6.0×10^{7}
14) 6.67×10^{-2} or 0.067
15) 1.4×10^{-2} or 0.014
16) 9.13×10^{-5}

Section 5.3
1) 22.9 cm
2) 48 min
3) 296 g
4) $5.48
5) Answers vary. A 120 lb person has a mass of 5.45×10^{4} g. 150 lb is 6.82×10^{4} g. 175 lbs is 7.95×10^{4} g.
6) 7.5 miles
7) 249 min
8) 3.64×10^{6} g
9) 15.7 km/L
10) 3.2×10^{4} miles/hr

Section 5.4
1) 1.5×10^{23} molecules H_2O
2) 2.71×10^{21} molecules $Al_2(CO_3)_3$
3) 1.66×10^{-4} mol H_2O
4) 8.3×10^{-15} mol C
5) 18.02 g/mol H_2O
6) 40.0 g/mol NaOH
7) 53.49 g/mol NH_4Cl
8) 98.08 g/mol H_2SO_4
9) 234.0 g/mol $Al_2(CO_3)_3$
10) 239.2 g/mol PbO_2

11) 1.5 mol NaOH
12) 0.058 mol H_2SO_4
13) 0.051 mol NH_4Cl
14) 0.042 mol PbO_2
15) 4.40 g CO_2
16) 48.05 g $(NH_4)_2CO_3$
17) 17.48 g NaOH
18) 54.06 g H_2O
19) 5.68×10^{21} molecules Na_2CO_3
20) 3.34×102^{25} molecules H_2O
21) 0.67 g H_2O
22) 169.85 g NaCl
23) 5.71 g NaOH

Section 6.1
1) Solutions have extremely small particles that allow light to go through. Colloids have larger particles which scatter light. Suspensions separate into layers upon standing.
2) Solutions will allow light to go through and colloids will not.
3) C
4) A

Section 6.2
1) Polar (or ionic) compounds will dissolve in polar compounds. Nonpolar compounds will dissolve in nonpolar compounds.
2) LiCl is ionic and CCl4 is nonpolar. Ionic compounds do not dissolve in nonpolar compounds, because the ionic compound has charged particles which are not attracted to the nonpolar solvent.
3) A
4) The oil is nonpolar and does not mix with the polar water.

Section 6.3
1) Ppm and ppb are convenient for very, very small concentrations.
2) 0.60 M
3) 0.22 m
4) 0.18 M
5) 1.35 M
6) 273 g
7) 5.68 m
8) 0.26 M
9) 0.17 M
10) 0.35 m
11) 0.37 m

Section 6.4
1) A, D
2) The salt lowers the freezing point of water, making it so it must be colder before the water will freeze into ice.
3) The salt raises the boiling temperature of the water, cooking the spaghetti at a higher temperature.
4) Ionic compounds split into separate ions when they dissolve, but covalent compounds stay as whole formulas.
5) Ionic, 2
6) Covalent, 1
7) Covalent, 1
8) Ionic, 3
9) Ionic, 4
10) Ionic, 2
11) 0.2 m $CaCl_2$
12) 0.1 m KI
13) 0.2m NaCl
14) 0.076°C
15) -1.86°C
16) -3.01°C

Section 7.1
1) Physical
2) Chemical
3) Physical
4) Chemical
5) Chemical
6) Chemical
7) Chemical
8) Physical
9) Chemical
10) Chemical

Section 7.2
1) A, B, C

2) A
3) A
4) B
5) C
6) Crushed ice
7) Sugar crystals
8) Wood shavings
9) If the surface area is higher, there are more collisions between reacting particles. The greater the frequency of effective collisions, the faster the reaction.
10) Slower
11) Faster
12) Slower

Section 7.3

1) $Ca + H_2O \rightarrow Ca(OH)_2 + H_2$
2) $NaOH + Cl_2 \rightarrow NaCl + NaClO + H_2O$
3) $Fe + S \rightarrow FeS$
4) $Al + H_2SO_4 \rightarrow H_2 + Al_2(SO_4)_3$
5) $Al + Fe_2O_3 \rightarrow Al_2O_3 + Fe$
6) $F_2 + NaOH \rightarrow NaF + O_2 + H_2O$
7) $Fe + CuNO_3 \rightarrow Fe(NO_3)_2 + Cu$

Section 7.4

1) $2\ Cu + O_2 \rightarrow 2\ CuO$
2) $2\ H_2O \rightarrow 2\ H_2 + O_2$
3) $2\ Fe + 3\ H_2O \rightarrow 3\ H_2 + Fe_2O_3$
4) $2\ NaCl \rightarrow 2\ Na + Cl_2$
5) $2\ AsCl_3 + 3\ H_2S \rightarrow As_2S_3 + 6\ HCl$
6) $CaCO_3 \rightarrow CaO + CO_2$
7) $H_2S + 2\ KOH \rightarrow 2\ HOH + K_2S$
8) $XeF_6 + 3\ H_2O \rightarrow XeO_3 + 6\ HF$
9) $Cu + 2\ AgNO_3 \rightarrow 2\ Ag + Cu(NO_3)_2$
10) $4\ Fe + 3\ O_2 \rightarrow 2\ Fe_2O_3$
11) $2\ Al(OH)_3 + Mg_3(PO_4)_2 \rightarrow 2\ AlPO_4 + 3\ Mg(OH)_2$
12) $2\ Al + 3\ H_2SO_4 \rightarrow 3\ H_2 + Al_2(SO_4)_3$
13) $H_3PO_4 + 3\ NH_4OH \rightarrow 3\ HOH + (NH_4)_3PO_4$
14) $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$
15) $4\ Al + 3\ O_2 \rightarrow 2\ Al_2O_3$
16) $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$
17) 5
18) 9
19) Changing the subscripts changes which compounds are involved in the chemical reaction, while changing the coefficients only changes how many of a specific substances are involved in the reaction.

Section 7.5

1) Synthesis
2) Decomposition
3) Single replacement
4) Double replacement
5) Single replacement
6) Decomposition
7) Double replacement
8) Decomposition
9) Single replacement
10) Combustion
11) Synthesis combines two or more substances into one product, whereas decomposition splits one reactant into more than one product.
12) $CO_2 + H_2O$
13) Single replacement; $2\ Fe + 3\ H_2O \rightarrow 3\ H_2 + Fe_2O_3$
14) Double replacement; H_2O (or HOH) + $(NH_4)_3PO_4$
15) Combustion; $CO_2 + H_2O$
16) Synthesis; Al_2O_3
17) Double replacement; $BaSO_4 + NaCl$
18) Single replacement; $CaCl_2 + H_2$
19) Double replacement; $FeCl_2 + H_2S$
20) Single Replacement; $NaBr + I_2$

Section 7.6

1) 3 mol H_2O
2) 0.31 mol Bi_2O_3
3) 15 mol LiCl
4) 1.05 mol SiO_2
5) 0.15 mol $Ca_3(PO_4)_2$
6) 0.36 mol Fe_2O_3
7) 3.28 mol FeS
8) 2.79 g HNO_3
9) 2.12 g I_2
10) 0.78 g LiOH

11) 66.5 g Cs
12) 297 L O_2
13) 4.3×10^{22} molecules CO
14) 12.7 g O_2
15) 180.2 g H_2O

Section 7.7

1) PCl_5 is reacting to form PCl_3 and Cl_2 at the same time and at the same speed that PCl_3 and Cl_2 are recombining to form PCl_5.
2) Between 2.0 and 2.5 minutes the reaction reaches equilibrium, because the concentration is no longer changing after this time.
3) The rate of the forward reaction is equal to the rate of the reverse reaction.
4) F
5) T
6) T
7) F
8) T

Section 7.8

1) Solids and liquids
2) $K = \frac{[H_2O]^2}{[H_2]^2[O_2]}$
3) $K = \frac{[NOBr]^2}{[NO]^2[Br_2]}$
4) $K = \frac{[O_2][NO_2]}{[NO][O_3]}$
5) $K = \frac{[CO][H_2]^3}{[CH_4][H_2O]}$
6) $K = \frac{[CH_3OH]}{[CO][H_2]^2}$
7) $K = \frac{[CO_2]^4[H_2O]^6}{[C_2H_6]^2[O_2]^7}$
8) $K = \frac{[C_2H_4][H_2]}{[C_2H_6]}$
9) $K = \frac{[HgI_2]}{[I_2][Hg]}$
10) $K = \frac{[CO_2]^2}{[CO]^2}$
11) $K = [Cu^{2+}][OH^-]^2$
12) Products are favored over reactants; there is a greater concentration of products than reactants at equilibrium.
13) Reactants are favored over products; there is a greater concentration of reactants than products at equilibrium.
14) 0.64
15) 2.89 M
16) $K = \frac{[H_2][CO_2]}{[CO][H_2O]}$; K=1.34
17) $K = \frac{[H_2]^2[S_2]}{[H_2S]^2}$; $K=1.1 \times 10^{-4}$
18) $a)\ K = \frac{[SO_3]^2}{[SO_2]^2[O_2]}$; b)more products are present at equilibrium because K>1; c)$[SO_3]$=0.103 M

Section 7.9

1) The equilibrium shifts toward the products.
2) The equilibrium shifts toward the reactants.
3) More products are formed
4) More reactants are formed
5) More reactants are formed
6) More products are formed
7) More reactants are formed
8) More products are formed
9) More products are formed
10) More reactants are formed
11) More reactants are formed
12) More products are formed
13) More products are formed
14) More reactants are formed
15) More reactants are formed
16) More products are formed
17) More reactants are formed
18) More products are formed
19) More products are formed
20) More products are formed
21) More reactants are formed
22) More reactants are formed

Section 8.1

1) Acids
2) Bases
3) Acids
4) Bases
5) Acids
6) Bases

7) Acids react to form H^+ ions in water

Section 8.2

1) 2.5×10^{-2} M; 12.4; basic
2) 1.7×10^{-8} M; 6.22; slightly acidic
3) 3.16×10^{-4} M; 3.5; acidic
4) 3.16×10^{-7} M; 6.5; slightly acidic
5) 6.6
6) 8.2
7) The solution with the lower pH has 100,000 times greater concentration of H^+ ions.

Section 8.3

11) $HNO_3 + KOH \rightarrow H_2O + KNO_3$
12) $HClO_4 + NH_4OH \rightarrow H_2O + NH_4ClO_4$
13) $H_2SO_4 + 2\ NaOH \rightarrow 2\ H_2O + Na_2SO_4$
14) $HNO_3 + NH_4OH \rightarrow H_2O + NH_4NO_3$
15) $HF + NH_4OH \rightarrow H_2O + NH_4F$
16) $HC_2H_3O_2 + KOH \rightarrow H_2O + KC_2H_3O_2$
17) $HCl + KOH \rightarrow H_2O + KCl$
18) $Mg(OH)_2 + 2\ HCl \rightarrow 2\ H_2O + MgCl_2$
19) $2\ HCl + Ba(OH)_2 \rightarrow 2\ H_2O + BaCl_2$
20) $NaOH + HClO_4 \rightarrow H_2O + NaClO_4$

Section 8.4

1) Indicators are weak acids that change color when they react with a base. They are used in a titration to show when all of the acid or base has reacted.
2) When the number of moles of acid is equal to the number of moles of base
3) 0.1176 M NaOH
4) 0.1708 M HClO
5) 33.2 mL HCl
6) 62.5 mL HI
7) 1.4 M NaOH
8) 0.30 M $HC_2H_3O_2$

Section 9.1

1) Endothermic reactions absorb (take in) energy; exothermic reactions release energy.
2) Exothermic
3) Endothermic
4) Exothermic
5) Exothermic
6) Exothermic
7) Endothermic
8) Exothermic
9) Endothermic
10) Exothermic; the temperature rises initially meaning that heat was given off to the surroundings.

Section 9.2

1) Oxidized: Cu; Reduced: H
2) Oxidized: H; Reduced: O
3) Oxidized: Al; Reduced: H
4) Oxidized: Zn; Reduced: H
5) Oxidized: Al; Reduced: Cu
6) Energy of moving electrons
7) The anode is where electrons are lost (where oxidation occurs)
8) The cathode is where electrons are gained (where reduction occurs)
9) Zinc
10) The anode is the zinc electrode, because the zinc is being oxidized
11) The anode is the copper electrode, because copper ions are being reduced
12) Electrons will flow from the anode (zinc electrode) to the cathode (copper electrode)

Section 10.2

1) C
2) B
3) A
4) Alpha particles will move toward the negative plate; beta particles will move toward the positive plate; gamma particles will continue in a straight path.
5) The number of protons increases by 1, the number of neutrons decreases by 1, and the total number of particles in t he nucleus does not change.
6) The number of protons decreases by 2, the number of neutrons decreases by 2,

and the total number of particles in the nucleus decreases by 4
7) Gamma
8) Alpha
9) ${}^{204}_{82}Pb$
10) ${}^{210}_{83}Bi$
11) ${}^{131}_{53}I \rightarrow {}^{0}_{-1}e + {}^{131}_{54}Xe$
12) ${}^{230}_{90}Th \rightarrow {}^{4}_{2}He + {}^{226}_{88}Ra$
13) ${}^{234}_{90}Th \rightarrow {}^{0}_{-1}e + {}^{234}_{91}Pa$
14) ${}^{3}_{1}H \rightarrow {}^{0}_{-1}e + {}^{3}_{2}He$
15) ${}^{198}_{86}Th \rightarrow {}^{4}_{2}He + {}^{194}_{84}Po$
16) ${}^{237}_{92}U \rightarrow {}^{0}_{-1}e + {}^{237}_{93}Np$

Section 10.3

1) 0.25g
2) 1.0 g
3) 8.0 years
4) 17,100 years
5) 10 years

Section 10.4

1) The process by which a large nucleus is split into two or more smaller nuclei
2) The process by which two small nuclei combine to make one larger nuclei
3) Nuclear changes involve much more energy then chemical changes (frequently about 1 million times more energy per atom)
4) The manner in which the heat is produced that heats the water to turn the turbine
5) Control the speed at which the fission reaction occurs by absorbing many of the free neutrons which start fission reactions
6) No, the power plants in the US contain less than the critical mass of the fissionable isotopes so are unable to cause a nuclear explosion
7) Fission
8) Nuclear decay
9) Fusion
10) Nuclear decay
11) Fission and fusion

Glossary

A	
Absolute Zero:	the temperature at which molecules stop moving and therefore, have zero kinetic energy
Alkali earth metals:	group 2A of the periodic table
Alkali metals:	group 1A of the periodic table
Alpha decay:	Alpha decay is a common mode of radioactive decay in which a nucleus emits an alpha particle (a helium-4 nucleus).
Alpha particle:	An alpha particle is a helium-4 nucleus, composed of 2 protons and 2 neutrons
Anion:	negative ion; formed by gaining electrons
Anode:	The electrode at which oxidation occurs.
Arrhenius acid:	a substance that produces H+ ions in solution
Arrhenius base:	a substance that produces OH- ions in a solution
Atom:	Democritus' word for the tiny, indivisible, solid objects that he believed made up all matter in the universe
Atomic mass unit (amu):	a unit of mass equal to one-twelfth the mass of a carbon-twelve atom
Atomic mass:	the weighted average of the masses of the isotopes of an element
Atomic number:	the number of protons in the nucleus of an atom
Avogadro's number:	The number of objects in a mole; equal to 6.02×10^{23}.
B	
Background radiation:	Radiation that comes from environment sources including the earth's crust, the atmosphere, cosmic rays, and radioisotopes. These natural sources of radiation account for the largest amount of radiation received by most people.
Balanced chemical equation:	a chemical equation in which the number of each type of atom is equal on the two sides of the equation
Battery:	A group of two or more cells that produces an electric current.
Beta decay:	Beta decay is a common mode of radioactive decay in which a nucleus emits beta particles. The daughter nucleus will have a higher atomic number than the original nucleus.
Beta particle:	A beta particle is a high speed electron, specifically an electron of nuclear origin.
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 elevation:	the amount the boiling point of a solution increases from the boiling point of a pure solvent
C	
Catalyst:	A substance that increases the rate of a chemical reaction but is, itself, left unchanged, at the end of the reaction; lowers activation energy
Cathode:	electrode at which reduction occurs.
Cation:	positive ion; formed by losing electrons
Chain reaction:	A multi-stage nuclear reaction that sustains itself in a series of fissions in which the release of neutrons from the splitting of one atom leads to the splitting of others.
Chemical changes:	changes that occur when one substance is turned into another substance; different types of molecules are present at the beginning and end of the change.
Chemical reaction:	the process in which one or more substances are changed into one or more new substances

Coefficient:	a small whole number that appears in front of a formula in a balanced chemical equation
Colligative property:	a property that is due only to the number of particles in solution and not the type of the solute
Colloid:	type of mixture in which the size of the particles is between 1×10^3 pm and 1×10^8 pm
Combustion reaction:	a reaction in which oxygen reacts with another substance to produce carbon dioxide and water.
Compound:	a substance that is made up of more than one type of atom bonded together
Concentrated:	a solution in which there is a large amount of solute in a given amount of solvent
Concentration:	the measure of how much of a given substance is mixed with another substance
Control rods :	made of chemical elements capable of absorbing many neutrons and are used to control the rate of a fission chain reaction in a nuclear reactor.
Controlled experiment:	An experiment that compares the results of an experimental sample to a control sample
Conversion factor:	a ratio used to convert one unit of measurement into another.
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 by atoms.
Covalent compound:	two or more atoms (typically nonmetals) forming a molecule in which electrons are being shared between atoms.
Critical mass:	The smallest mass of a fissionable material that will sustain a nuclear chain reaction at a constant level.
D	
Dalton's Atomic Theory:	the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom
Decomposition reaction:	a reaction in which one reactant breaks down to form two or more products
Dilute:	a solution in which there is a small amount of solute in a given amount of solvent
Double replacement reaction:	a reaction in which two reactants form products by having the cations exchange places with the anions
Ductile:	can be drawn out into thin wires
E	
Electrochemical cell:	An arrangement of electrodes and ionic solutions in which a redox reaction is used to make electricity; a battery
Electrolysis:	A chemical reaction brought about by an electric current.
Electron configuration:	a list that represents the arrangement of electrons of an atom.
Electron:	a negatively charged subatomic particle, responsible for chemical bonding
Electronegativity:	The tendency of an atom in a molecule to attract shared electrons to itself.
Electronegativity:	the ability of an atom in a molecule to attract shared electrons
Electroplating:	A process in which electrolysis is used as a means of coating an object with a layer of metal.
electrostatic attraction:	The force of attraction between opposite electric charges.
electrostatic attraction:	the attraction of oppositely charged particles
Element:	a substance that is made up of only one type of atom Subatomic particles: particles that are smaller than the atom
Endothermic:	reactions in which energy is absorbed, heat can be considered as a reactant

Equilibrium constant (K):	A mathematical ratio that shows the concentrations of the products divided by concentration of the reactants.
Equilibrium:	A state that occurs when the rate of forward reaction is equal to the rate of the reverse reaction.
Equivalence point:	the point in the titration where the number of moles of acid equals the number of moles of base
Exothermic reaction:	A reaction in which heat is released, or is a product of a reaction.
Experiment:	A controlled method of testing a hypothesis.
Extrapolation:	the process of creating data points beyond the end of the graph line, using the basic shape of the curve as a guide
F	
Fission:	A nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy.
Formula unit:	the empirical formula of an ionic compound; shows the lowest possible ratio
Freezing point depression:	the amount the freezing point of a solution decreases from the freezing point of a pure solvent
Fusion:	A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy.
G	
Gamma ray:	Gamma radiation is the highest energy on the spectrum of electromagnetic radiation.
Graph:	a pictorial representation of patterns using a coordinate system
Group (family):	a vertical column in the periodic table, have similar chemcial properties
H	
Half-life:	the time interval required for a quantity of material to decay to half its original value.
Halogens:	group 7A of the periodic table; reactive non-metals
Hydrocarbon:	an organic substance consisting of only hydrogen and carbon
Hypothesis:	A tentative explanation that can be tested by further investigation.
I	
Immiscible:	liquids that do not have the ability to dissolve in each other
Indicator:	a substance that changes color at a specific pH and is used to indicate the pH of the solution
International System of Units (Le Système International d' Unites):	the internationally agreed upon standard metric system
Interpolation:	the process of estimating values between measured values
Ion:	An atom or group of atoms with an excess positive or negative charge, lost or gained electrons
ionic bond:	A bond between ions resulting from the transfer of electrons from one of the bonding atoms to the other and the resulting electrostatic attraction between the ions.
Ionic compound:	a positively charged particle (typically a metal) bonded to a negatively charged particle (typically a nonmetal) held together by electrostatic attraction
Ionic Formula:	includes the symbols and number of each ion (atom) present in a compound in the lowest whole number ratio
Ionization energy:	the energy required to remove the most loosely held electron from a gaseous atom or ion
Isotopes:	atoms of the same element that have the same number of protons but different numbers of neutrons, same atomic number but different mass number

L	
Le Châtelier's Principle:	Applying a stress to an equilibrium system causes the equilibrium position to shift to offset that stress and regain equilibrium.
M	
malleable:	can be hammered into thin sheets
Mass number:	the total number of protons and neutrons in the nucleus of an atom
Mass:	a measure of the amount of matter in an object
Mendeleev:	the Russian chemist credited with organizing the periodic table in the form we use today.
Metric system:	international decimal-based system of measurement.
Miscible:	liquids that have the ability to dissolve in each other
Mixture:	a combination of two or more elements or compounds which have not reacted to bond together; each part in the mixture retains its own properties
Molality:	
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.
Molarity:	the number of moles of solute per liter of solution
Mole ratio:	the ratio of the moles of one reactant or product to the moles of another reactant or product according to the coefficients in the balanced chemical equation
Mole:	An Avogadro's number of objects; 6.02×10^{23} particles
Molecular geometry:	The specific three-dimensional arrangement of atoms in molecules.
N	
Neutralization:	a reaction between an acid and a base that produces water and a salt
Neutron:	a subatomic particle with no charge
Noble gases:	group 8A of the periodic table; extremely non-reactive
Nuclear charge:	the number of protons in the nucleus
Nucleus:	the small, dense center of the atom
O	
Octet rule:	the tendency for atoms gain or lose the appropriate number of electrons so that the resulting ion has either completely filled or completely empty outer energy levels, or 8 valance electrons.
Oxidation:	a loss of electrons, resulting in an increased charge or oxidation number
P	
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.
Physical changes:	changes that do not alter the identity of a substance, the same types of molecules are present at the beginning and end of the change.
Polar covalent bond:	A covalent bond in which the electrons are not shared equally because one atom attracts them more strongly that the other.
Potential energy:	The energy of position or stored energy, including bond energy.
Products:	materials present at the end of a reaction, shown on the right of the arrow in a chemical equation
Proton:	a positively charged subatomic particle
Q	
Quark:	particles that form one of the two basic constituents of matter. Various species of

	quarks combine in specific ways to form protons and neutrons, in each case taking exactly three quarks to make the composite particle.
R	
Reactants:	the starting materials in a reaction, shown left of the arrow in a chemical equation
Reduction:	gaining electrons, resulting in a decreased charge or oxidation number
S	
Scientific notation:	a shorthand method of writing very large and very small numbers in terms of a decimal number between 1 and 10 multiplied by 10 to a power.
Shielding effect:	the inner electrons help "shield" the outer electrons and the nucleus from each other.
Single replacement reaction:	a reaction in which an element reacts with a compound to form products
Slope:	the ratio of the change in one variable with respect to the other variable.
Solute:	the substance in a solution present in the least amount, dissolved by the solute
Solution:	a homogeneous mixture of substances
Solvent:	the substance in a solution present in the greatest amount
Stoichiometry:	the calculation of quantitative relationships of the reactants and products in a balanced chemical equation
Subscripts:	part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element
Surface area to volume ratio:	The comparison of the volume inside a solid to the area exposed on the surface.
Suspension:	type of mixture in which the particles settle to the bottom of the container and can be separated by filtration
Synthesis reaction:	a reaction in which two or more reactants combine to make one product
T	
Temperature:	the average kinetic energy of the particles that make up a material
Theory:	A well-established explanation based on extensive experimental data
Titration:	the lab 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
V	
Valence electrons:	the electrons in the outermost energy level of an atom.
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.
W	
Weight:	the force of attraction between the object and the earth (or whatever large body it is resting on)

Proof

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