

Introduction to Chemistry

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Fr = weakest

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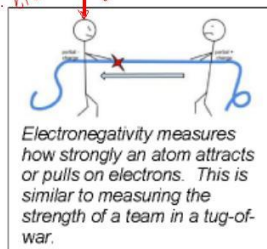
Write on the pictures. Summarize what you read on charts and pictures.

3.4: Periodic Trends

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.

more electronegative electrons pulled closer



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.

Stronger (more electronegative)

weaker

Stronger (more electronegative)

where are the noble gases?

1																	5	6	7	8	9
H																	B	C	N	O	F
2.20																	2.04	2.55	3.04	3.44	3.98
3	4															13	14	15	16	17	
Li	Be															Al	Si	P	S	Cl	
0.98	1.57															1.61	1.90	2.19	2.58	3.16	
11	12															31	32	33	34	35	
Na	Mg															K	Ca	Sc	Ti	V	
0.93	1.31															0.82	1.00	1.36	1.54	1.63	
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35					
K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br					
0.82	1.00	1.36	1.54	1.63	1.80	1.55	1.83	1.88	1.91	1.90	1.65	1.81	2.01	2.18	2.55	2.96					
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53					
Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I					
0.82	0.95	1.22	1.33	1.6	2.16	1.9	2.2	2.28	2.20	1.93	1.69	1.78	1.96	2.05	2.1	2.66					
55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85					
Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At					
0.79	0.89	1.1	1.3	1.5	2.36	1.9	2.2	2.20	2.28	2.54	2.00	1.62	2.33	2.02	2.0	2.2					
87	88																				
Fr	Ra																				
0.7	0.9																				

For this class, you do not need to memorize the electronegativity values. You do, however, need to understand the overall pattern that as you go from left to right across a period, the elements are more electronegative. As you move down a vertical column, the elements are less electronegative.

Electronegativity: The relative attraction an atom has for electrons in a bond.

what does this mean?

Add notes to example questions in the text about how the problem is being solved.

There is room to answer the practice questions in the book at the end of each section.

Vocabulary terms are listed at the bottom of most pages.

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Course Objectives by Chapter

Unit 1: Introduction to Chemistry and the Nature of Science

Nature of Science Goal—Science is based on observations, data, analysis and conclusions.

1. I can distinguish between hypothesis, theory, and law as these terms are used in science.
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.

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.

Unit 4: Describing Compounds

Nature of Science Goal—Vocabulary in science has specific meanings.

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.

Unit 5: Problem Solving and the Mole

Nature of Science Goal— Mathematics is a tool to increase scientific understanding.

1. I can describe the common measurements of the SI system of measurements
2. I can convert between standard notation and scientific notation.
3. I can convert between mass, moles, and atom or molecules using factor-label methods.

Unit 6: Mixtures and Their Properties

Nature of Science Goal-- Science provides predictable results.

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.
3. I can describe the relative amount a solute particles in concentrated and dilute solutions.
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: Chemistry & the Nature of Science



1.1: Experimentation

Why Experiment in Science?

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. 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.

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.



Aristotle believed logic could answer questions and didn't believe in experimentation.

Objectives:

- Describe scientific methods of solving problems.
- Identify independent, dependent, and controlled variables in experiments
- Describe the importance of having only one independent variable

1.1: Experimentation

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 "Scientific Method" was developed mainly by Francis Bacon (1561-1626)

The Steps in the Scientific Method

1. **Define the Problem:** Identify the problem or phenomenon that needs explaining.
2. **Make Observations:** Gather and organize data on the problem.
3. **Make a Hypothesis:** Suggest a possible solution or explanation.
4. **Test the Hypothesis:** Test the hypothesis by making new observations.
5. **Accept or Revise the Hypothesis:** 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.

Note that this should not be considered a "cookbook" for scientific research. Scientists do not sit down with their daily "to do" list and write down these steps. The steps may not be followed in order. But this does provide a general idea of how scientific research is usually done.

The Scientific Method: One of many methods used to answer questions in science

1.1: Experimentation

Experimental Design

Experimentation is the primary way through which science gathers evidence for ideas. It involves us causing 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 the conditions under which the phenomenon occurs. 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 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** (sometimes called the resultant 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.

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, give them the same amount of water, plant them in the same location, give them all the same pesticide treatment, etc. These would be controlled variables.

Controlled variable: Variable in an experiment that is held constant so it will not influence the outcome.

Independent (Manipulated) variable: Factor that is changed, or manipulated, by a researcher in a scientific experiment

Dependent (Resultant) variable: Factor in an experiment that is expected to change, or respond, when the manipulated variable changes

1.1: Experimentation

Examples of Experiments



It's exciting to roll down a skateboarding ramp, especially if you're going fast. The steeper the ramp, the faster you'll go. What else besides the steepness of a ramp influences how fast an object goes down it? You could do experiments to find out.

If you were to do an experiment to find out what influences the speed of an object down a ramp, what would be the dependent (resultant) variable? The responding variable would be the speed of the object. But there's another important question. What other variables might affect the speed? This is an important question to ask because

when designing an experiment to find how the slope of the ramp affects the speed, you would need to control other variables.

Let's assume you are sliding wooden blocks down a ramp in your experiment. You choose steepness of the ramp for your independent variable and the time it takes for the block to get to the bottom as your dependent variable. In other words, you are trying to find out how changing the steepness of a ramp affects the speed the block goes down the ramp. You decide to test two blocks on two ramps, one steeper than the other, and see which block reaches the bottom first. You use a shiny piece of varnished wood for one ramp and a rough board for the other ramp. You let go of both blocks at the same time and observe that the block on the ramp with the gentler slope reaches the bottom sooner. You're surprised, because you expected the block on the steeper ramp to go faster and get to the bottom first.

What explains your result? The problem is that you had more than one variable being changed. The ramps varied not only in steepness but also in smoothness. The block on the smoother ramp went faster than the block on the rougher ramp, even though the rougher ramp was steeper. Remember, in an experiment we want only one changing variable, the independent variable. Everything else should be kept the same.

Example: Suppose you wish to determine which brand of microwave popcorn leaves the fewest unpopped kernels. In your experiment: a) What is the independent variable, b) What is the dependent variable and c) What controlled variables would you need?

Solution:

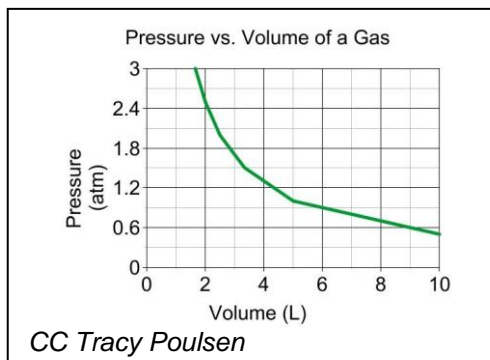
You will need a supply of various brands of microwave popcorn to test (independent variable). You will also need to use the same microwave for all of your bags of popcorn and cook them for the same amount of time. If you used different brands of microwave ovens with different brands of popcorn, the percentage of unpopped kernels could be caused by the different brands of popcorn, but it could also be caused by the different brands of ovens. You would not be able to conclude confidently whether the popcorn or the oven caused the difference. 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 changed

1.2 Data in Science

Why Make Graphs in Science

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.

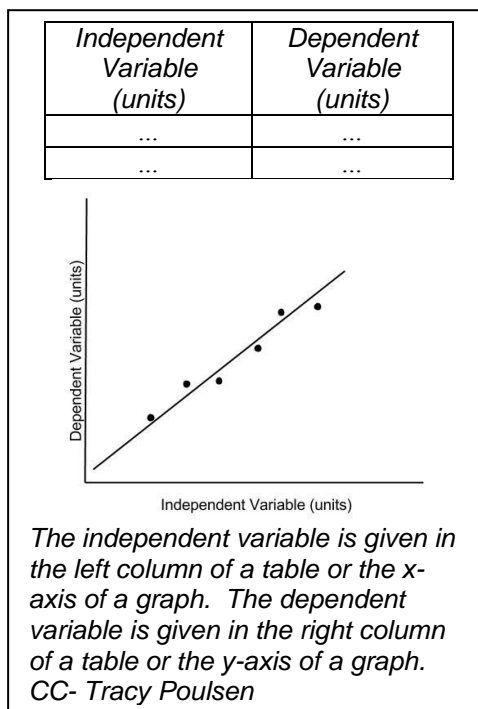
Volume (liters)	Pressure (atm)
10.0	0.50
5.0	1.00
3.33	1.50
2.50	2.00
2.00	2.50
1.67	3.00



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. 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.

Graphs and Experimental Variables

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).

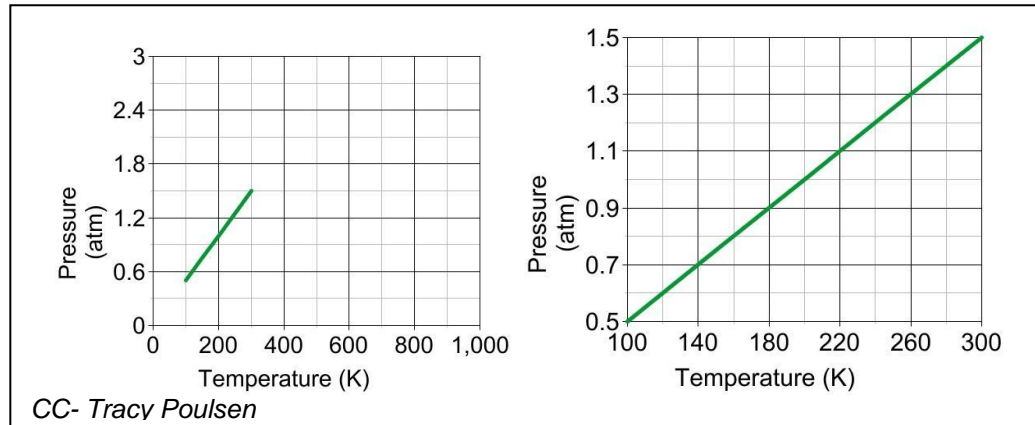
1.2 Data in Science

Drawing X-Y Scatter Plots

Objectives:

- Construct an appropriate graph of data
- Read a graph, correctly identifying the pattern of data, and interpolating and extrapolating the data
- Find the slope of a best-fit line of data

Reading information from a line graph is easier and more accurate as the size of the graph increases. In the two graphs shown, 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 left graph would give a less accurate result than using the right graph.



Features of X-Y Scatter plots

Hint: Try this website to make graphs:
<http://nces.ed.gov/nceskids/createagraph/default.aspx>

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

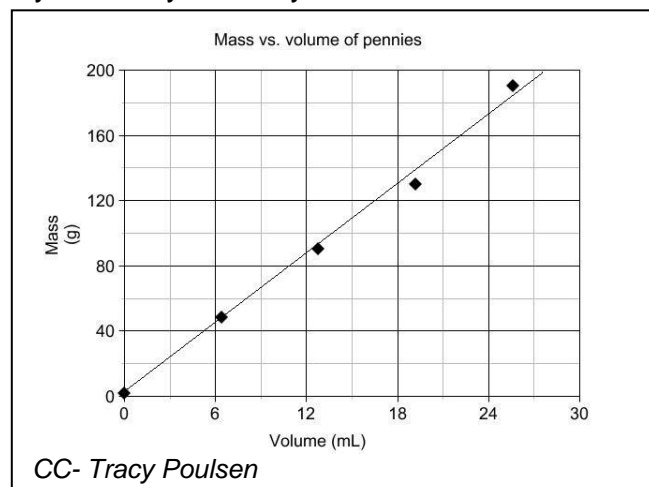
Scale: Take up as much of the graph paper as possible. The lowest x-value should be on the far left of the paper and the highest x-value should be on the far right side of the paper.

Your lowest y-value should be near the bottom of the graph and the highest y-value near the top. 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.

Labels: Both of your axes should be labeled, including units. What was measured along that axis and what unit was it measured in?

Best-fit-line: Draw a straight 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?

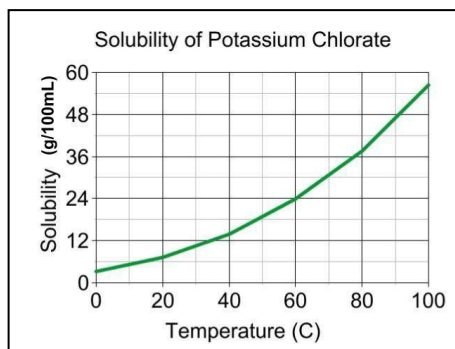


1.2 Data in Science

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 only a few data points, 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.

Temperature ($^{\circ}\text{C}$)	Solubility (g/100 mL H_2O)
0	3.3
20	7.3
40	13.9
60	23.8
80	37.5
100	56.3

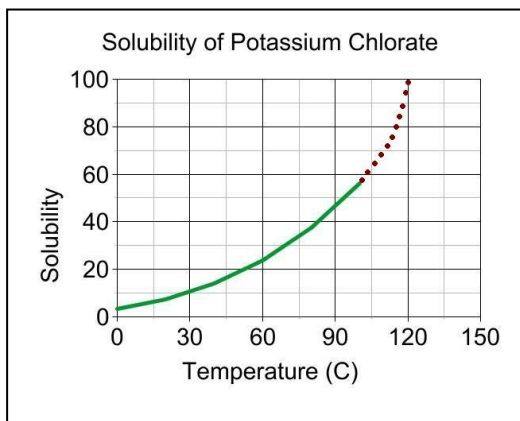


The table shows that there are exactly six known data points. When the data is graphed, however, the graph maker assumes that the relationship between the temperature and the solubility remains the same. The line is drawn by interpolating the data points between the actual data points.

We can now reasonably certainly read data from the graph for points that were not 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**.

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 . Extrapolation is more dangerous than interpolation in terms of possibly producing incorrect data.



1.2 Data in Science

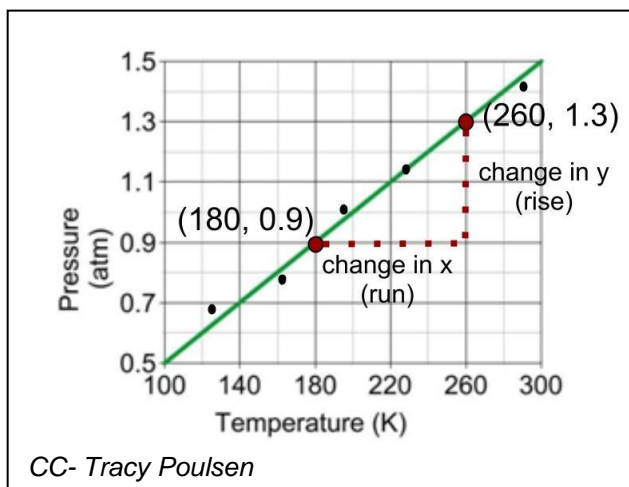
Finding the Slope of a Best Fit Line

As you may recall from math class, 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.

Consider the given graph. The relationship in this set of data is linear, that is, it produces a straight-line graph. The **slope** of a line is defined as the rise (change in vertical position) divided by the run (change in horizontal position).

$$\text{slope} = \frac{\text{rise}}{\text{run}} = \frac{(y_2 - y_1)}{(x_2 - x_1)}$$

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.



One other difference between what you may have done when calculating slope in math and how we calculate slope in science is that in science the slope is not just a number – it has units as well. The units are important because they tell us what measurements we are comparing with our slope. If the units were different, the slope would be different as well. The units of the slope are the y-axis units over the x-axis units.

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 example, the points selected are $(260, 1.3)$ and $(180, 0.9)$. The slope can then be calculated in the manner:

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

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.

Best-fit-line: a line or curve on a graph that goes as close to the data points as possible, showing the overall pattern of the data. Data points may or may not actually be on the line.

1.3 Science Terms

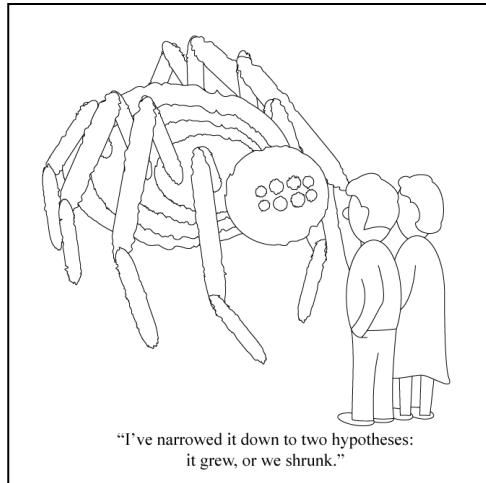
Hypotheses

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:

- 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 Suggested Explanation – hypotheses can be suggestions about why something is observed. A hypothesis is very tentative; it can be easily changed.

A Scientific Hypothesis Must Be Testable AND Falsifiable

For a hypothesis to be testable means that it is possible to make observations that agree or disagree with it. 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.” If a hypothesis cannot be tested by making observations, it is not scientific. Consider this statement:



Objectives:

- Describe the difference between hypothesis, theory, and law.
- Explain the limitations of models as scientific representations of reality.

"There are invisible creatures all around us that we can never observe in any way."

This statement may or may not be true, but it is not a scientific hypothesis, because it can't be tested. Given the nature of the hypothesis, there are no observations a scientist could make to test whether or not it is false.

A Scientific Hypothesis Must Be Falsifiable

A hypothesis may be testable, but even that isn't enough for it to be a scientific hypothesis. In addition, it must be possible to show that the hypothesis is false if it really is false. Consider this statement:

"There are other planets in the universe where life exists."

This statement is testable. If it is true, it is at least theoretically possible to find evidence showing that it's true. For example, a spacecraft could be sent from Earth to explore the universe and report back if it discovers an inhabited planet. If such a planet were found, it would prove the statement is true. However, the statement isn't a scientific hypothesis. Why? If it is false, it's not possible to show that it's false. The spacecraft may never find an inhabited planet, but that doesn't necessarily mean there isn't one. Given the vastness of the universe, we would never be able to check every planet for life!

Hypothesis: A reasonable prediction or suggested explanation.

1.3 Science Terms

Theories

The term “theory” is perhaps the most misunderstood term used in science. Most people think the word “theory” means a guess or suggestion. That is what the word means in ever-day-life, but not in science. As we just learned, the word “hypothesis” is a better word to use for a suggestion or reasonable guess. Scientists use the word “theory” much differently than most people do.

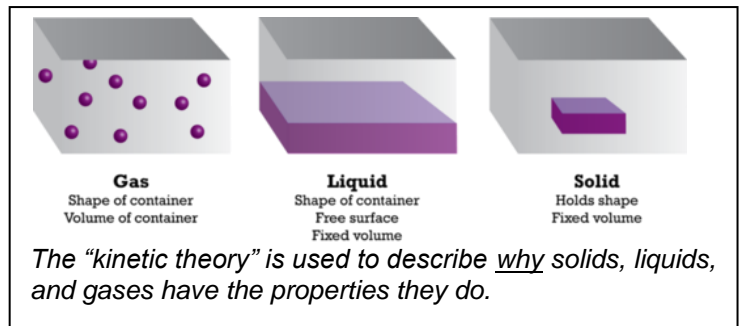
The *United States National Academy of Sciences* describes what a theory is as follows:

“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 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 observations or evidence which 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.

Theory: A well-established explanation

1.3 Science Terms

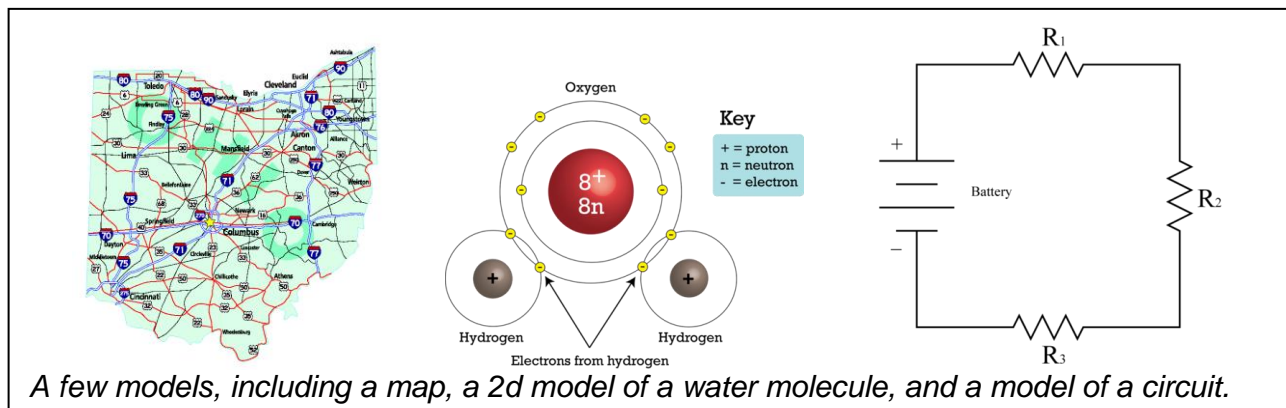
Models

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.

A **model** is a description, graphic, or 3-D representation of a 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 make 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.



Model: a description or 3d representation of a theory

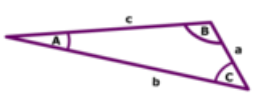
1.3 Science Terms

Laws

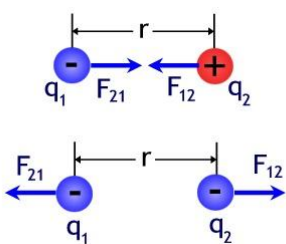
Scientific laws are similar to scientific theories in that they are principles that can be used to predict the behavior of the natural world. Both scientific laws and scientific theories are well-supported by observations and/or experimental evidence. Scientific **laws** refer to rules for how nature will behave under certain conditions, frequently written as an equation. Scientific theories, however, 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.

Scientific laws state *what* always happens. This can be very useful. It can let you predict what will happen under certain circumstances. For example, Newton's third law tells you that the harder you hit a softball with a bat, the faster and farther the ball will travel away from the bat.

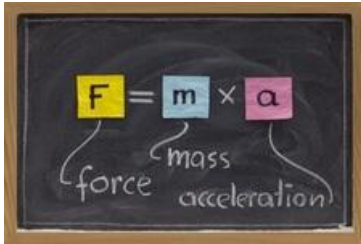
However, scientific laws have a basic limitation. They don't explain *why* things happen. "Why" questions are answered by scientific theories, not scientific laws.

$$\frac{\sin(A)}{a} = \frac{\sin(B)}{b} = \frac{\sin(C)}{c}$$

$$\frac{a}{\sin(A)} = \frac{b}{\sin(B)} = \frac{c}{\sin(C)}$$

The Law of Sines



Coulomb's Law



Newton's Second Law of Motion

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 was stated as Charles' 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 not 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.

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.

Law: descriptions of how nature will behave under certain conditions.

1.1: Experimentation

- Scientists use experimentation to test their ideas.
- In an experiment, it is important to change only one variable to see its effect on another variable.
- 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.
- The scientific method is one way of solving problems in science, but there are other methods that are acceptable.

1.2: Data in Science

- Two common methods of presenting data that aid in the search for regularities and trends are tables and graphs.
- In a data table, the independent variable is recorded in the first column and the dependent variable is recorded in the last column. On a graph, the independent variable is plotted on the x-axis and the dependent variable is plotted along the y-axis.
- When creating a graph, it is important to appropriately scale the axes to take up the entire space, label both axes with units, and draw a best fit line showing the pattern in the data.
- When we draw a scatter plot 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.
- Slope is calculated from two points on the best fit line using the equation: $\text{slope} = \frac{\text{rise}}{\text{run}} = \frac{(y_2 - y_1)}{(x_2 - x_1)}$
- The slope of a graph represents the rate at which one variable is changing with respect to the other variable. The slope of a graph will also have units.

1.3: Science Terms

- 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.
- 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).

1.1: Experimentation Review Questions

- 1) How many independent variables should any experiment have?
 - a) One
 - b) Many
 - c) It depends on the experiment

- 2) How many controlled variables should an experiment have?
 - a) One
 - b) Many
 - c) It depends on the experiment

- 3) What is another name for a dependent variable?

#4-7: Gary noticed that two tomato 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.

- 4) What is Gary's hypothesis?
 - a) Tomato plants grow tall.
 - b) Tomatoes should be planted in identical pots and receive identical amounts of water.
 - c) The tomato plants grew different heights because of the amount of sun they received.
 - d) Tomatoes are tasty.

- 5) 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) There is no possible scientific reason for having the same size containers.

- 6) 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 was conserving water.
 - d) There is no possible scientific reason for having the same amount of water for each plant every day.

- 7) What was the variable being tested in Gary's experiment (what is the independent variable)?
- The amount of water
 - The amount of plant food
 - The amount of sunshine
 - The type of soil
- 8) 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. He finds that the plants without detergent grew taller than the plants with detergent. What should be the conclusion from his experiment?
- Plants with detergent don't grow as tall as plants without detergent.
 - He cannot draw any valid conclusions, because he had more than one variable changing.
 - The more water a plant receives the taller the plant grows.
 - Corn grows taller than green beans.

A student is experimenting with Alka-Seltzer tablets. She puts one tablet in 0°C water, one tablet in 10°C water, another in 20°C water and another in 40°C water. For each test, she uses 100 mL of water placed in a 250 mL beaker. She then times how long it takes for the tablets to completely react.

9) What is the independent variable in this experiment?

10) What is the dependent variable in this experiment?

11) What variables are controlled in this experiment?

A student is conducting an experiment reacting Alka-seltzer tablets in water. She puts 100 mL of ice water in one beaker, 100 mL of room temperature water in a second beaker, 100 mL of luke-warm water in another beaker, and 100 mL of hot water in a fourth beaker. She puts one tablet in each of the beakers and times how long it takes until the Alka-seltzer tablet disappears. She records her data in a data table.

12) What is the independent variable in this experiment?

13) What is the dependent variable in this experiment?

14) What variables are controlled in this experiment?

- 15) Astronomers perform no scientific experiments. For example, they don't control variables to see how the mass of a star affects its life cycle. Because they don't experiment, they cannot follow the scientific method. Does this mean that when astronomers make claims about the life cycle of stars their claims are not valid, because they didn't follow the scientific method? Why or why not?

1.2: Data in Science Review Questions

- 16) On a data table, where is the independent variable typically listed? What about the dependent variable?

- 17) On a graph, how do you identify the independent variable and dependent variable?

Use the given data tables or graphs to answer each of the following questions.

#18-20: The following data was collected by measuring the intensity of light given off by a light bulb.

Distance (meters)	Intensity of light
1.0	0.29645
1.1	0.25215
1.2	0.20547
1.3	0.17462
1.4	0.15342
1.5	?
1.6	0.11450

- 18) What is the purpose of the experiment for which this data was collected? To see how _____ affects _____.

- 19) What would you expect the intensity of light to be at 1.5 m from the light source?

- 20) Complete the following sentence describe the pattern in this data: as the distance from the light increases, the intensity of the light _____.

#21-23: The following data was collected regarding a general education course at a large Midwestern university. (Source: *College Teaching*, Winter 2005, Vol. 53, Issue 1)

# if absences	Final Grade
0	89.2
1	86.4
2	83.5
3	81.1
4	78.2
5	73.9
6	64.3

- 21) What is the purpose of the experiment for which this data was collected? To see how _____ affects _____.

- 22) What would you expect the final grade to be for students who missed 2 classes?

- 23) What would you predict the final grade to be for students who missed 7 classes?

24) What is the purpose of the experiment for which this data was collected? To see how _____ affects _____.

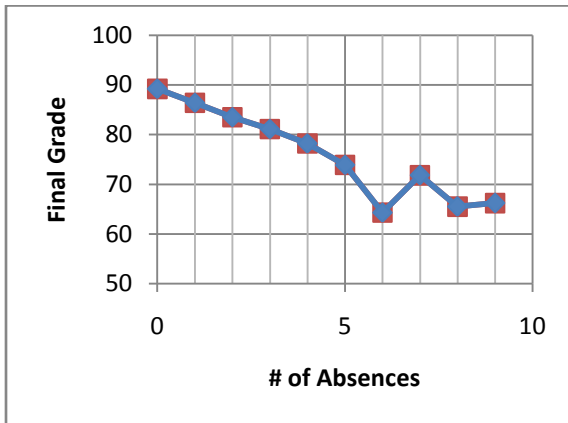
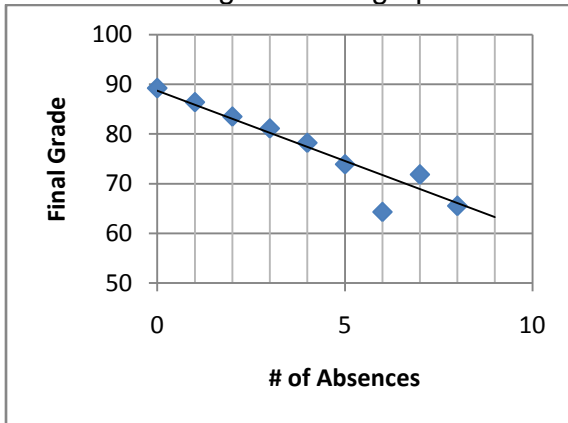
25) What gas mileage would you predict for a car that weighs 3000 lbs?

26) Complete the following sentence describing the pattern in this data: as weight increases, gas mileage _____.

#24-26 (Source: www.roadandtrack.com)

Car	Weight (pounds)	Miles per gallon
Dodge Neon	2580	27
Ford Focus	2655	26
Saturn Ion	2690	26
Pontiac Sunfire	2770	24
Chevrolet PT Cruiser	3100	21
Chevrolet Corvette	3180	19

Use the following data and graphs to answer the questions:



#27-28: The following data was collected regarding a general education course at a large Midwestern university. (Source: *College Teaching*, Winter 2005, Vol. 53, Issue 1)

# if absences	Final Grade
0	89.2
1	86.4
2	83.5
3	81.1
4	78.2
5	73.9
6	64.3

27) Which of the graph shows the appropriate trend-line?

28) Which of the following is the best way to calculate the slope of the best-fit line from the given graph and data?

$$\frac{86.4 - 73.9}{1 - 5} = -3.12$$

(#'s from data table)

$$\frac{86.4 - 73.9}{1 - 5} = \frac{12.5}{-4}$$

(#'s from data table)

$$\frac{80.0 - 70.0}{3 - 7} = -2.52$$

(#'s from best-fit line)

29) Andrew was completing his density lab for his chemistry class. He collected the given data for volume and mass.

- Identify the independent and dependent variables in this experiment.
- Draw a graph to represent the data, including a best-fit-line.
- If the graph is a straight line, calculate the slope, including units.
- What would you expect the mass of 2.5 mL of solution to have?
- What volume would you expect 60 g of the solution to occupy?

#29 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

30) Donna is completing an experiment to find the rate of a reaction by measuring how the concentration of ammonia changes over time.

- Identify the independent and dependent variables in this experiment.
- Draw a graph to represent the data, including a best-fit-line
- If the concentration of ammonia was 0.30 mol/L, how much time has passed?
- After 8 seconds, what will be the approximate concentration of ammonia?

#30 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

31) Consider the data table for an experiment on the behavior of gases.

- Identify the independent and dependent variables in this experiment.
- Draw a graph to represent the data.
- Calculate the slope, including units.
- What would be the pressure at 55°C?
- What would be the pressure at 120°C?

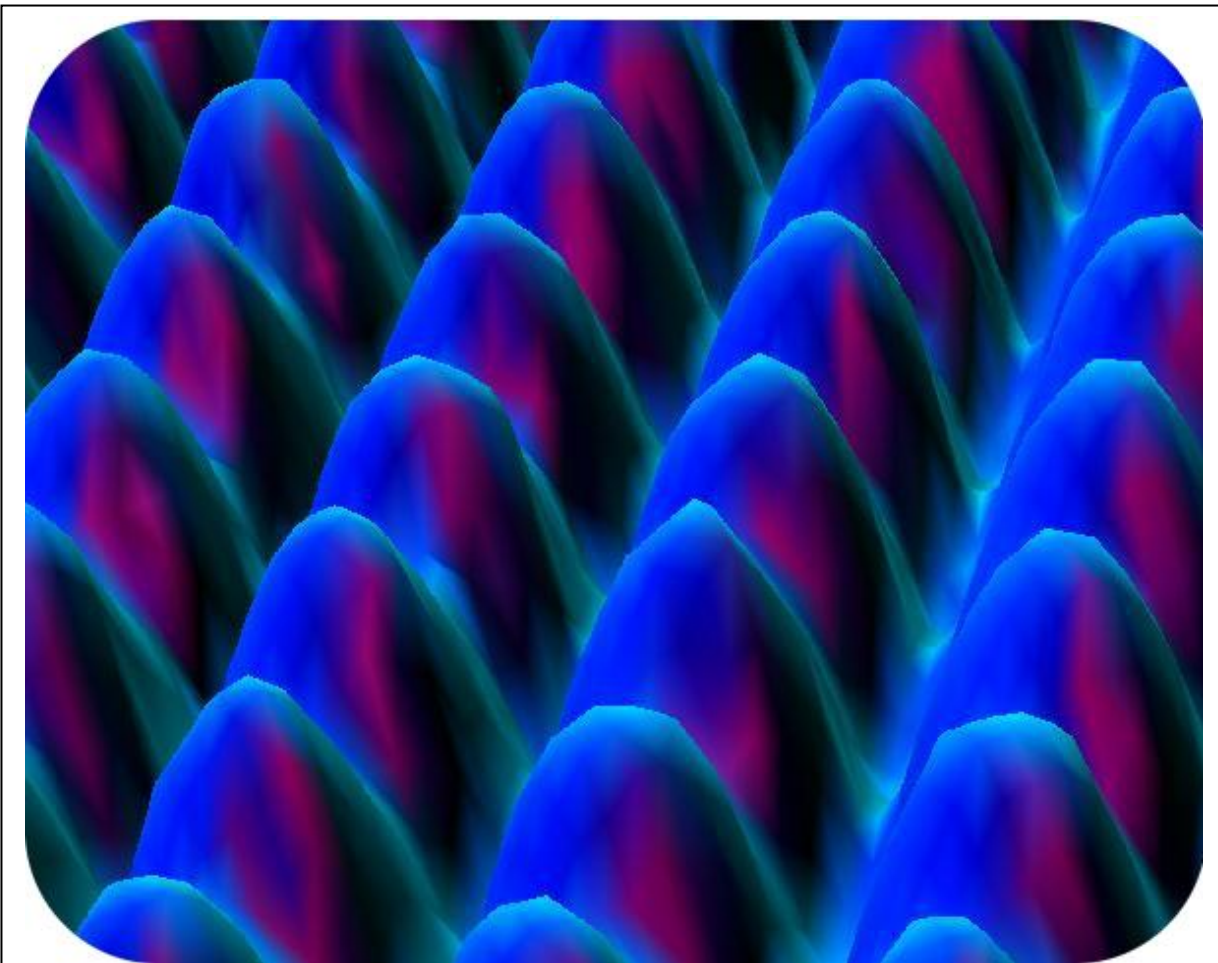
#31 data	
Temperature (°C)	Pressure (mmHg)
10	726
20	750
40	800
70	880
100	960

1.3: Science Terms Review Questions

- 32) 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.
- 33) 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.
- 34) 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.
- 35) 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.
- 36) A well-substantiated explanation of an aspect of the natural world is a:
- a) Theory.
 - b) Law.
 - c) Hypothesis.
 - d) None of these.
- 37) Which of the following words is closest to the same meaning as hypothesis?
- a) Fact
 - b) Law
 - c) Formula
 - d) Suggestion
 - e) Conclusion
- 38) 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.
- 39) True/False: When a theory has been known for a long time, it becomes a law.
- 40) Dalton thought that atoms are the smallest particles of matter. Scientists now know that atoms are composed of even smaller particles. Does this mean that the rest of Dalton's atomic theory should be thrown out?

Chapter 2

The Structure of the Atom



*What could this hilly blue surface possibly be? Do you have any idea?
The answer is row upon row of atoms of the metallic element nickel.
The picture was created using a scanning tunneling microscope.
No other microscope can make images of things as small as atoms.*

2.1: Early Ideas of Atoms

The Greek Philosophers

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.

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. 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 atoms)** – 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. 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.

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. The problem, of course, was that Greek philosophers didn't believe in experiments at all.

Greek philosophers didn't trust their senses; they only trusted the reasoning power of the mind. They 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. 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. 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.



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.



Greek philosophers tried to understand the nature of the world through reason and logic.

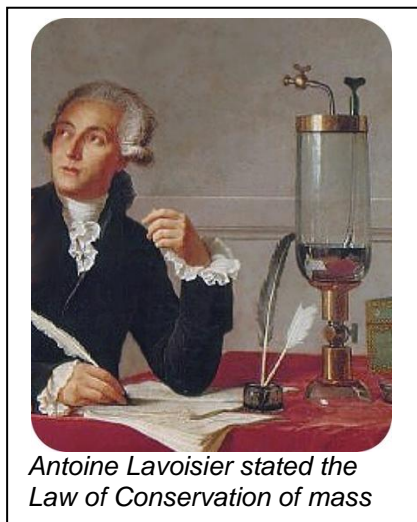
Objectives

- 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.

2.1: Early Ideas of Atoms

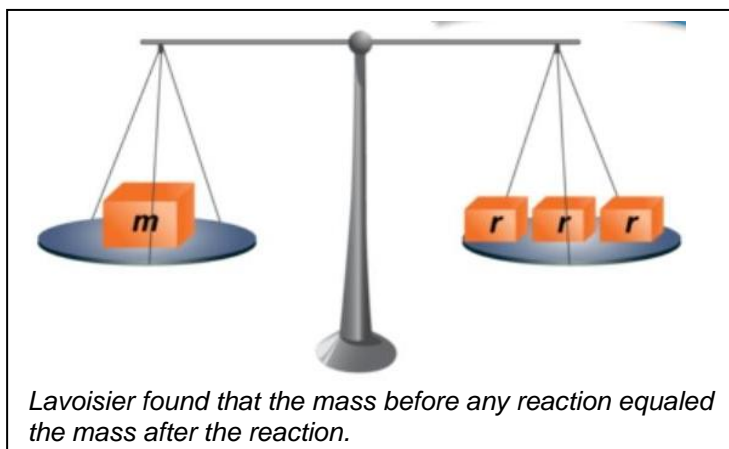
Evidence for Atoms

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. Remember, theories are written to explain what is observed. Let's first look at some of the observations and laws which the atomic theory was later able to explain.



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.



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

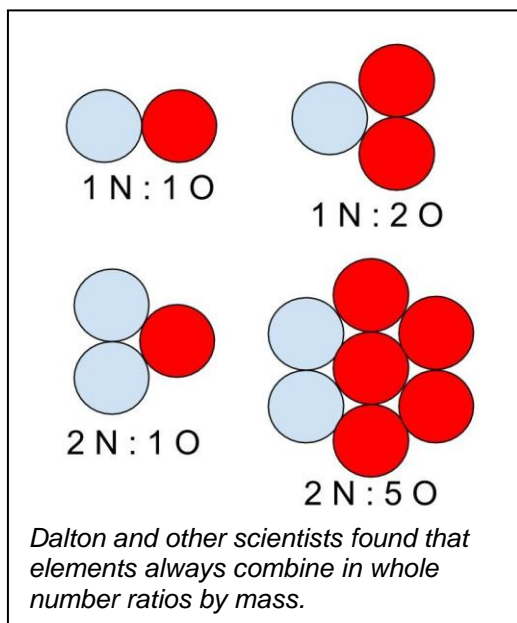
(hydrogen and oxygen) always reacted in the same proportions by mass. In other words, if 1 gram of hydrogen reacted with 8 grams of oxygen, then 2 grams of hydrogen would react with 16 grams of oxygen, and 3 grams of hydrogen would react with 24 grams of oxygen. The observation that hydrogen and oxygen always reacted in the "same proportions by

2.1: Early Ideas of Atoms

mass” was true of other compounds also. 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. This idea that elements always reacted in the same ratios is described in **law of definite 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?

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. The observation that elements always combined in whole number ratios came to be known as the **law of multiple proportions**.



All of these observations and laws are important. But, remember, laws summarize *what* is observed, but they don’t explain *why*. What could explain why these patterns were observed?

Law of Conservation of Mass: In any change, the mass of the starting materials equals the mass of the ending materials.

Law of Definite Proportions: In any compound, the elements always combine in the same ratio of elements

Law of Multiple Proportions: Elements always combine in whole number ratios.

2.1: Early Ideas of Atoms

Dalton's Atomic Theory

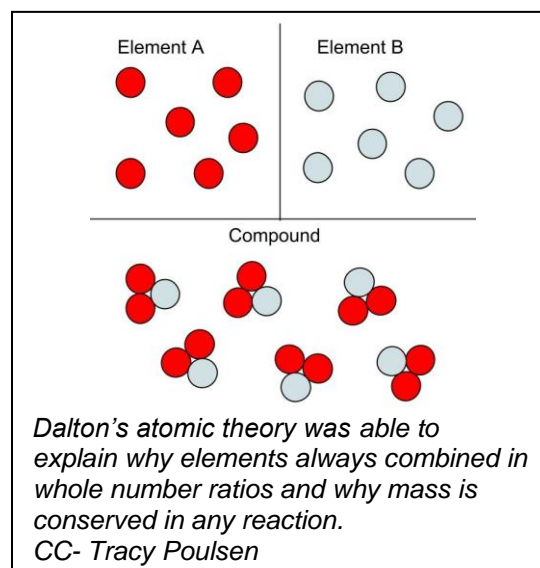
Dalton thought about his results and tried to find some theory that would explain his observations and the observations that other scientists were making. One way to explain the relationships that Dalton and others had observed was to suggest that materials like nitrogen, carbon and oxygen were composed of small, indivisible quantities which Dalton called "atoms" (in reference to Democritus' original idea). Dalton's theory would explain the law of conservation of mass by saying that the atoms (and mass) are not being created or destroyed in a reaction. The atoms are simply reacting with other atoms that are already present. Furthermore, elements always combined in whole number ratios because it was indivisible atoms that were combining together.



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

Dalton used this idea to generate what is now known as **Dalton's Atomic Theory** which stated the following:

1. All substances are made of atoms. Atoms are the smallest particles of matter. They cannot be divided into smaller particles, created, or destroyed.
2. All atoms of the same element are alike and have the same mass. Atoms of different elements are different and have different masses.
3. Atoms join together to form compounds, and a given compound always consists of the same kinds of atoms in the same proportions.



Dalton's Atomic Theory was able to explain many of the observations scientists made regarding the properties of matter. 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.

2.2: Further Understanding of the Atom

The Discovery of the Electron

In the mid-1800s, 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.



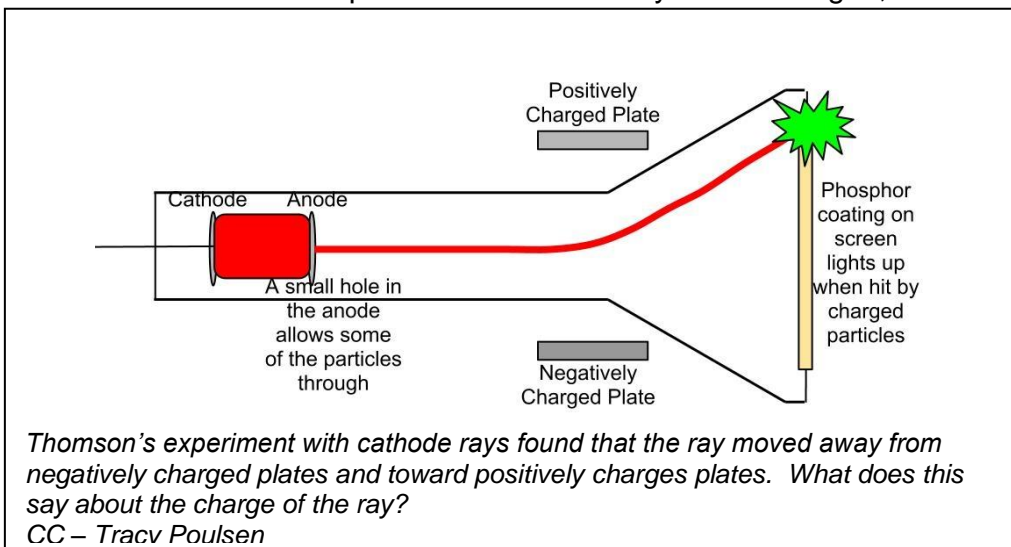
J.J. Thomson conducted experiments that suggested that Dalton's atomic theory wasn't telling the entire story.

Objectives

-Summarize evidence that led to the discovery of subatomic particles (including electrons, protons, and neutrons).

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, 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, 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.

Thomson must have suspected that cathode rays were charged, because



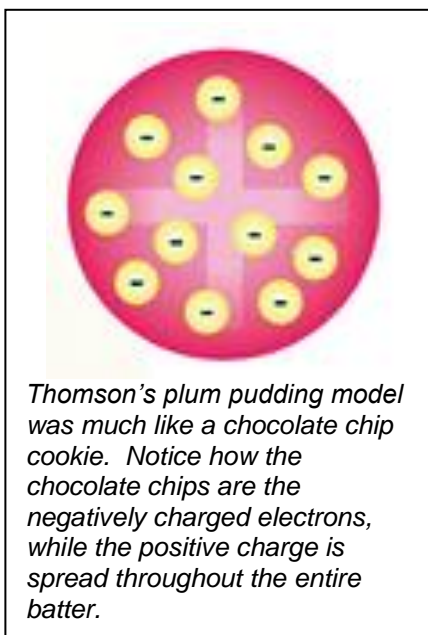
his next step was to place a positively charged metal plate on one side of the cathode ray tube and a negatively charged metal plate on the other side of the cathode ray tube. The metal plates didn't actually touch the cathode ray tube, but they were close enough that a remarkable thing happened! The flow of the cathode rays was bent upwards towards the positive metal plate and away from the negative metal plate. Using the “opposite charges attract, like charges repel” rule, Thomson argued that if the cathode rays were attracted to the positively charged metal plate and

2.2: Further Understanding of the Atom

repelled from the negatively charged metal plate, they themselves must have a negative charge!

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. But there was a problem. According to 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 could the matter making up cathode rays be smaller than an atom if atoms were indivisible? Thomson made a radical proposal: maybe atoms are divisible. 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.

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.



To account for overall neutral charge of an atom, 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. 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 model described his observations well. It would be easy to pick the chocolate chips (electrons) out of a cookie. On the other hand, it would be a lot harder to pick the batter (positive material) out, because the batter is everywhere.

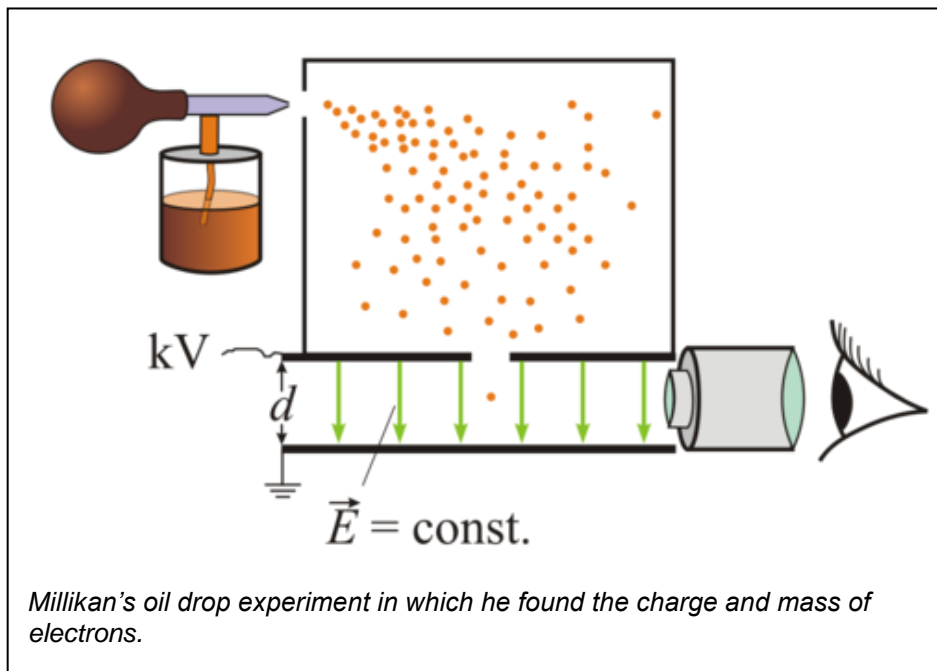
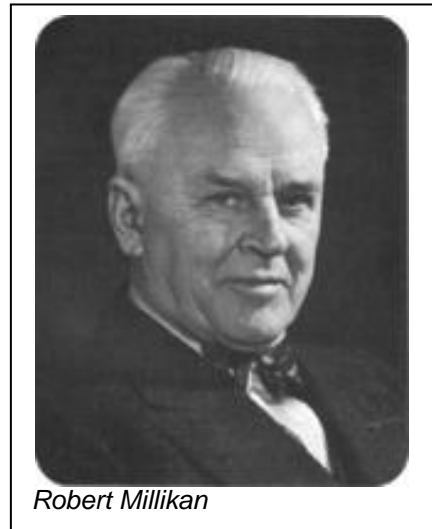
Electron A negatively charged particle found in atoms.

Plum Pudding Model: Thomson's model of the atom which had negatively charged electrons embedded in a positive mass.

2.2: Further Understanding of the Atom

The Charge of Electron and Discovering the Proton

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 $1.6 \times 10^{-19} \text{C}$, the charge on a single electron.



Based on the fact that atoms are neutral, and based on 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.

Proton: A positively charged particle found in the nucleus of atoms.

2.2: Further Understanding of the Atom

The Discovery of the Nucleus

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



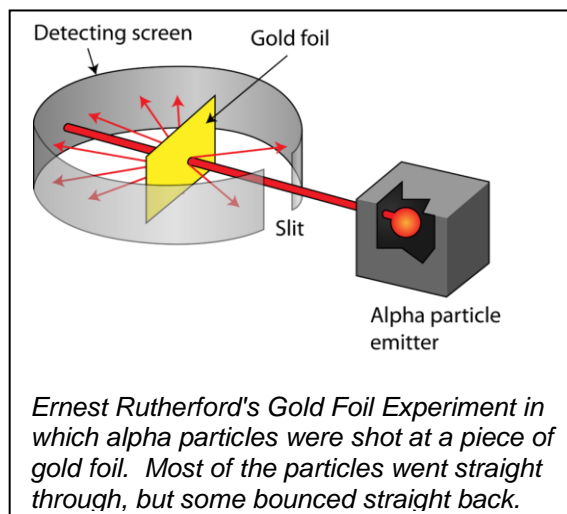
Disproving Thomson's "plum pudding" model began with the discovery that an element known as uranium emits positively charged particles called alpha particles as it undergoes radioactive decay. Radioactive decay occurs when one element decomposes into another element. It only happens with a few very unstable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

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

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

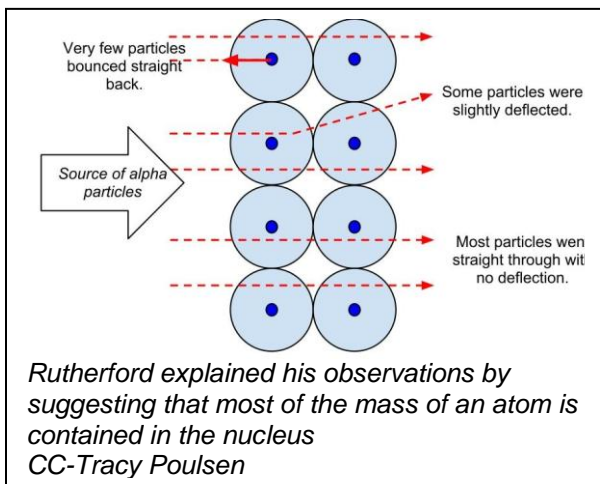
Every so often, though, one of the alpha particles would be deflected slightly as if it had bounced off of something hard. Even less often, Rutherford observed alpha particles bouncing straight back at the "gun" from which they had been fired! It was as if these alpha particles had hit a wall "head-on"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

Rutherford thought that these experimental results were rather odd. He 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



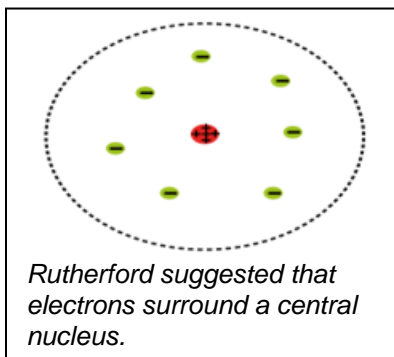
2.2: Further Understanding of the Atom

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 at the center

of the atom 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.



Because Thomson's model was unable to explain this observation, Rutherford's model became the accepted theory. According to the "plum pudding" model, electrons were like plums embedded in the positive "batter" of the atom. Rutherford's model suggested that the positive charge wasn't distributed like batter, but 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.

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.

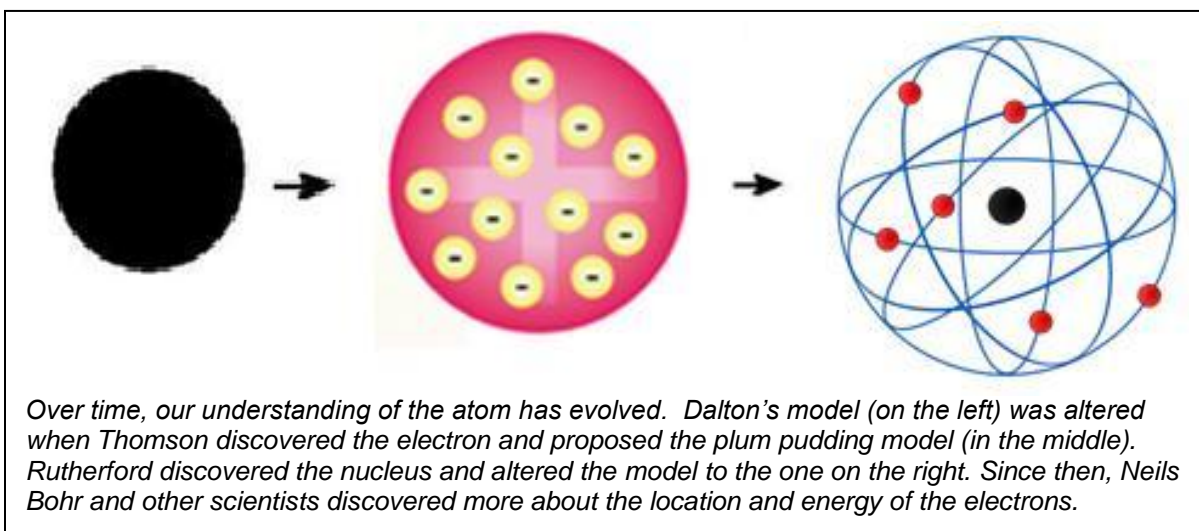
Nucleus: The small, dense center of the atom containing most of the mass of an atom.

Nuclear Model: Rutherford's model of the atom in which most of the mass is located in a small, dense, positive nucleus and the rest of the atom is mostly empty space.

2.2: Further Understanding of the Atom

Development of Atomic Models

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.



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 be replaced.

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) and atomic spectra



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

2.3: Protons, Neutrons, and Electrons in Atoms

Properties of Protons, Neutrons and Electrons

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**. Ernest Rutherford proposed the existence of a third particle, with later, James Chadwick proved that the nucleus of the atom contains this neutral particle that had been proposed by Ernest Rutherford – the **neutron**.

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 - hydrogen), and they're bound together with other neutrons and protons in the atomic nucleus.

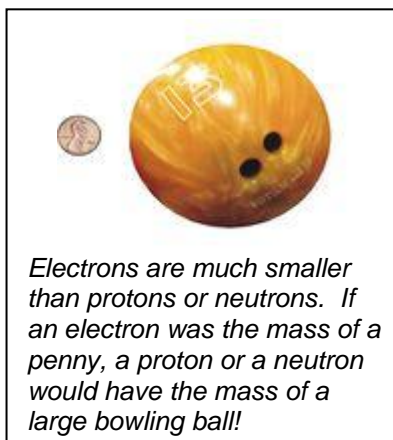
Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the 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.

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.

Objectives:

- Compare the properties of protons, neutrons, and electrons in atoms
- Identify an element based on the number of protons
- Describe the similarities and differences between isotopes of the same element
- Find the number of protons, neutrons, and electrons in an atom or ion.



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

Neutron: A neutral (no charge) particle found in the nucleus of atoms

2.3: Protons, Neutrons, and Electrons in Atoms

Atomic Number and Mass Number

John Dalton said that atoms of the same element all had the same mass. Essentially, he defined an element based on its mass. However, we have since learned that atoms of the same element can have different masses. We needed to find a different way to define elements.

Now scientists distinguish between atoms of different elements by the number of protons in an atom's nucleus. 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? How many protons do carbon atoms have?

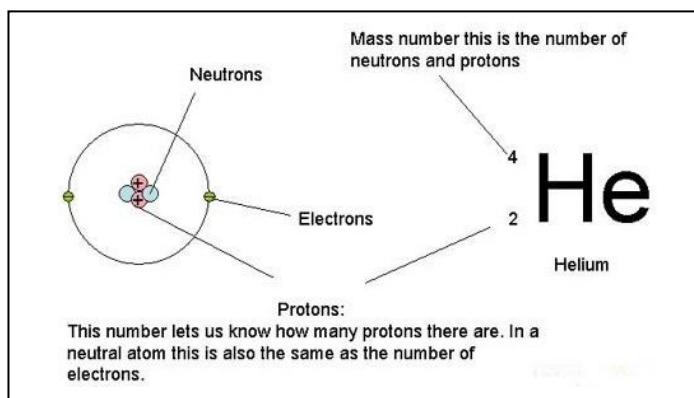
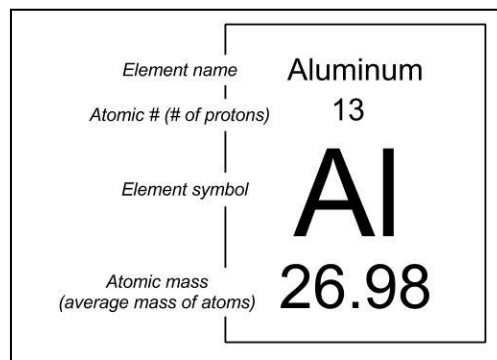
Of course, since neutral atoms have to have one electron for every proton, an element's atomic number also tells you how many electrons are in a neutral atom of that element. For example, hydrogen has an atomic number of 1. This means that an atom of hydrogen has one proton, and, if it's neutral, one 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. Counting the number of protons and neutrons tells scientists about the total mass of an atom.

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

It is important to point out that the mass number is *not* found on the periodic table. The mass number is specific to one isotope of an element.



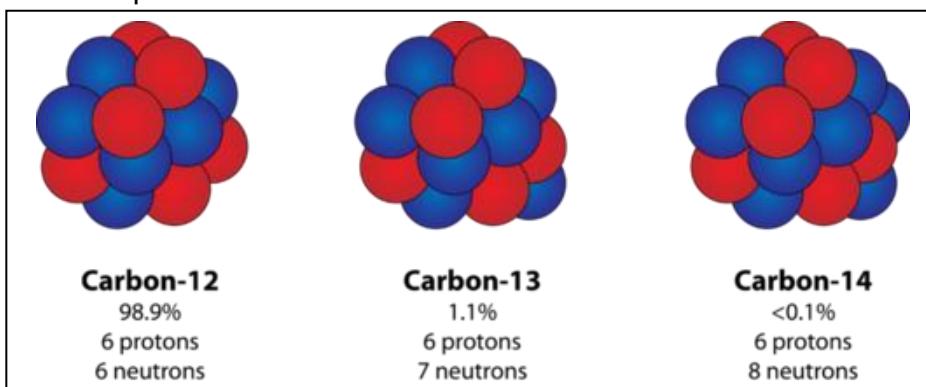
Atomic Number: The number of protons in an atom's nucleus, used to identify which element an atom is

Mass Number: The total of an atom's protons and neutrons (particles in the nucleus)

2.3: Protons, Neutrons, and Electrons in Atoms

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. However, since the isotopes of a given element contain different numbers of neutrons different isotopes have different mass numbers.



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. 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 samples of elements 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.

Atomic Mass Unit

Atoms have so little mass that common units of measuring mass aren't very useful. 1 g of iron (think of a paperclip) has over 10,000,000,000,000,000,000 atoms in it. Each of these atoms weighs about 0.000000000000000000009 grams. Units such as grams or pounds are way too big when talking about the mass of individual atoms. Chemists derived a unit that is much more useful when talking about the mass of individual atoms.

An **atomic mass unit (amu)** is defined as one-twelfth the mass of a carbon-12 atom. These units are useful, because the mass of a proton and the mass of a neutron are almost exactly 1.0 in this unit system. An atom's mass in atomic mass units is really close to its mass number.

Isotope: Atoms of the same element, but containing different numbers of neutrons

2.3: Protons, Neutrons, and Electrons in Atoms Names and Symbols of Isotopes

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.

To write a **nuclear symbol**, the mass number is placed at the upper left of the chemical symbol and the atomic number is placed at the lower left 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.



In the nickel nucleus represented above, the atomic number 28 indicates the nucleus contains 28 protons, and therefore, it must contain 31 neutrons 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.

Look for this element on your periodic table to find the number of protons.	Mass Number #p + #n # of particles in the nucleus
Element - #	

The other way of representing these nuclei would be

Nickel-59 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.

Finding Number of Protons, Neutrons, and Electrons

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 .

Example: How many protons, electrons, and neutrons are in an atom of ${}_{19}^{40}\text{K}$?

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. $40 = \#p + \#n$ or $40 = 19 + ?$ There must be **21 neutrons** to make a total of 40 protons and neutrons.

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 **30 protons.**

For all atoms with no charge, the number of electrons is equal to the number of protons. **# electrons = 30.**

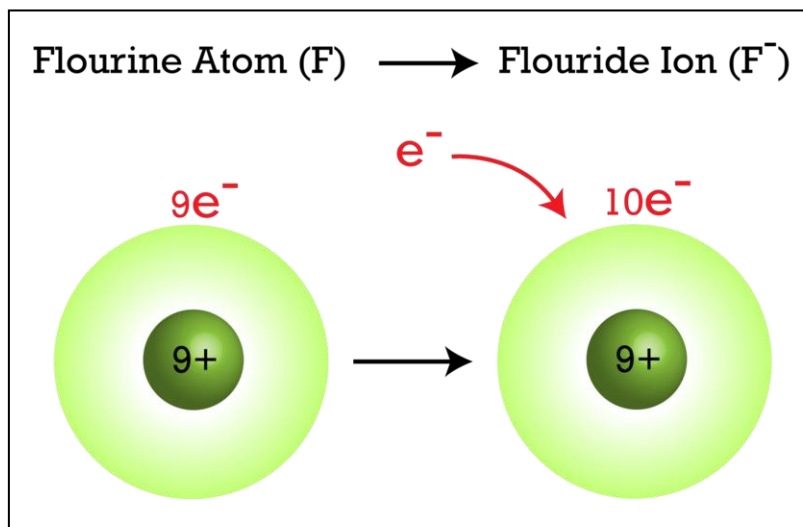
The mass number, 65, is the sum of the protons and the neutrons. $65 = \#p + \#n$ or $65 = 30 + ?$ There must be **35 neutrons** to make a total of 65 protons and neutrons.

2.3: Protons, Neutrons, and Electrons in Atoms

Ion Formation – Gaining and Losing Electrons

We have learned that atoms in their elemental state have no charge overall. This means that the number of protons will equal the number of electrons. However, when atoms combine with atoms of other elements forming compounds, they may gain or lose electrons and form **ions**. We will talk later about the formation of these compounds, but right now we have enough information to find out how many protons, neutrons, and electrons an atom has once it forms an ion.

In order for an atom to form a positively charged ion, it must lose electrons. The number of protons does not change, as this would also change the element. In order for an atom to form a negatively charged ion, it must gain electrons.



Example: How many protons, electrons, and neutrons are in an atom of $^{37}_{17}\text{Cl}$ with a -1 charge??

Solution:

Finding the number of protons is simple. The atomic number, # of protons, is listed in the bottom right corner. **# protons = 17.**

For all atoms with a negative charge, that atom gained electrons. Chlorine normally has 17 electrons with its 17 protons, but to get a -1 charge, chlorine had to gain 1 more electron. **# electrons = 18.**

The mass number, 37, is the sum of the protons and the neutrons. $37 = \#p + \#n$ or $37 = 17 + ?$ There must be **20 neutrons** to make a total of 37 protons and neutrons.

Example: How many protons, electrons, and neutrons are in an atom of $^{88}_{38}\text{Sr}^{+2}$?

Solution:

Finding the number of protons is simple. The atomic number, # of protons, is listed in the bottom right corner. **# protons = 38.**

For all atoms with a positive charge, the atom had to lose electrons. Sr atoms usually have 38 electrons and 38 protons. But this atom lost 2 electrons because it has a +2 charge. **# electrons = 36.**

The mass number, 88, is the sum of the protons and the neutrons. $88 = \#p + \#n$ or $88 = 38 + ?$ There must be **50 neutrons** to make a total of 88 protons and neutrons.

Ion: An atom or particle that has gained or lost electrons and has a negative or positive charge.

2.4: Atomic Mass

Atomic Mass – An Average

Objectives:

- Distinguish between atomic mass and mass number
- Identify which isotope of an element is more abundant
- Calculate the atomic mass of an element

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.

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:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \dots$$

Example: In a sample of neon, 90.92% of the atoms are Ne-20 with a mass of 19.99 amu. Another 0.3% of the atoms are Ne-21, with a mass of 20.99 amu. The final 8.85% of the atoms are Ne-22, with a mass of 21.99 amu. What is the atomic mass of neon?

Solution:

Neon has three isotopes. We will use the equation:

$$\text{Atomic mass} = (\%_1)(\text{mass}_1) + (\%_2)(\text{mass}_2) + \dots$$

Isotope 1: $\%_1=0.9092$ (write all percentages as decimals), $\text{mass}_1=19.99$

Isotope 2: $\%_2=0.003$, $\text{mass}_2=20.99$

Isotope 3: $\%_3=0.0885$, $\text{mass}_3=21.99$

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**

Element name	Aluminum
Atomic # (# of protons)	13
Element symbol	Al
Atomic mass (average mass of atoms)	26.98

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 neon (symbol Ne) is 20.18, which is what we calculated in the example. Not all periodic tables have the atomic number above the element's symbol and the atomic mass. The atomic number will 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.

Example: Chlorine has two isotopes, chlorine-35 and chlorine-37. Which isotope is more abundant?

Solution:

Chlorine's atomic (or average) mass is 35.45 amu as given on the periodic table. Because this average is closer to the mass of the chlorine-35 isotope, there must be more of this isotope than of chlorine-37.

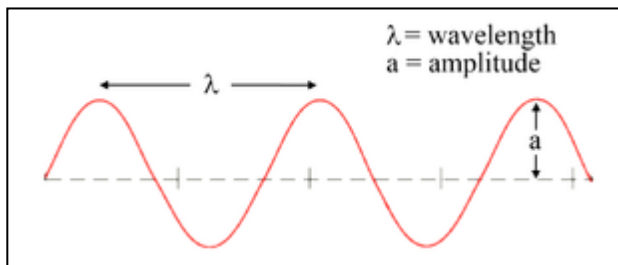
Atomic Mass: The average mass of atoms of an element; found on the periodic table.

2.5: Light and Atomic Spectra

Light Energy

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.

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 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 per 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.

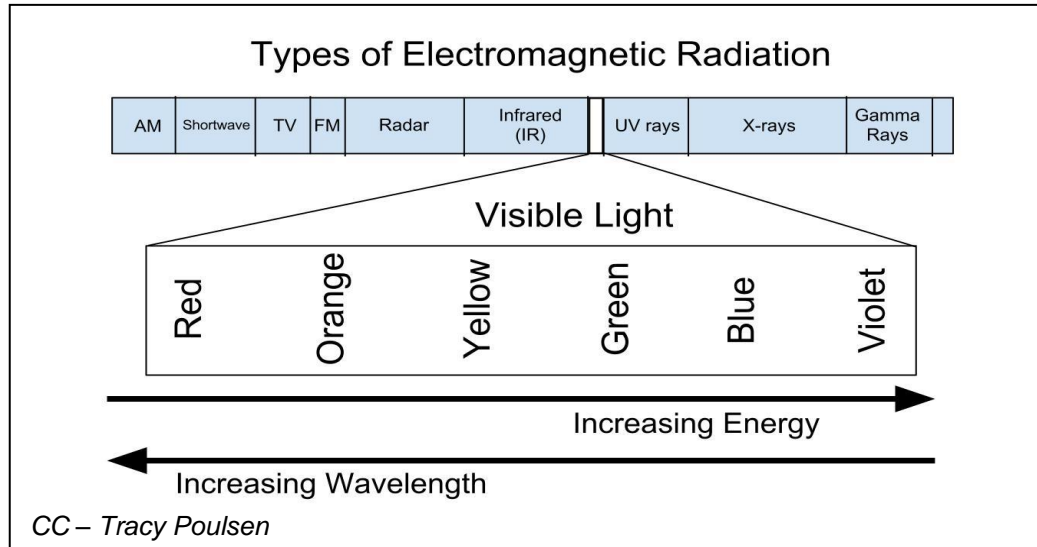
The diagram shows the electromagnetic spectrum. On the far left of the figure above are the electromagnetic waves with the highest energy. These waves are called gamma rays and can be quite dangerous in large numbers to living systems. The next lowest energy form of electromagnetic waves is called x-rays. 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 cancer. The tiny section next in the spectrum is the visible range of light. The highest form of visible light energy is violet light, with red light having

Objectives:

- Compare the wavelength and energy of a photon of light.
- Identify an element based on its unique spectrum
- Describe how electrons emit photons of light
- Give evidence supporting quantized energy levels for electrons

2.5: Light and Atomic Spectra

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.

Electromagnetic Spectrum: the range of all possible forms of “light” energy

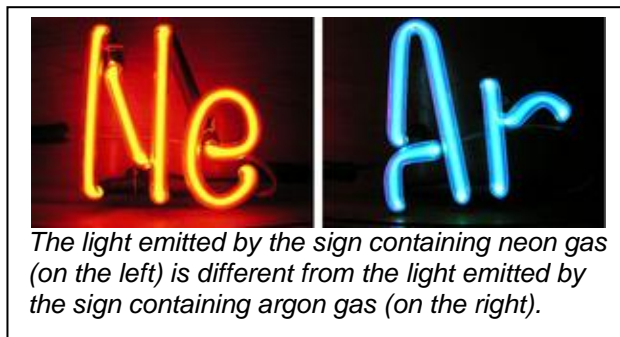
Wavelength: The distance between a point on one wave to the same point on the next wave

Frequency: the number of waves that pass a given point each second

2.5: Light and Atomic Spectra

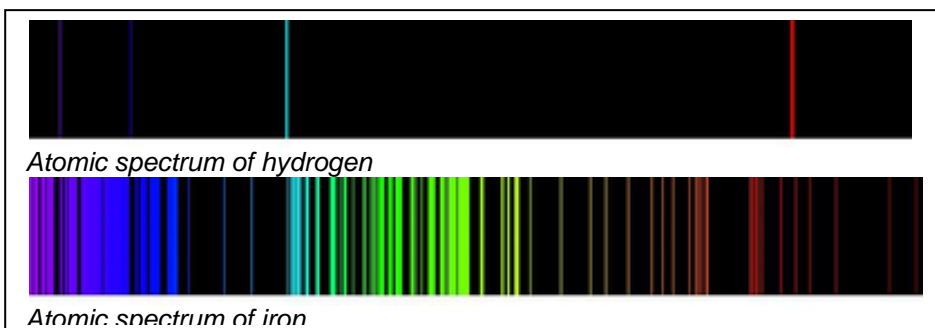
Each Element Has a Unique Spectrum

Electric light bulbs contain a very thin wire in them that emits light when heated. Any metal would work 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 other color. Every element emits light when energized by heating or passing electric current through it.



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 these energized atoms through glass prisms so 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.



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. Because each element has a unique emission spectrum, elements can be identified using them.

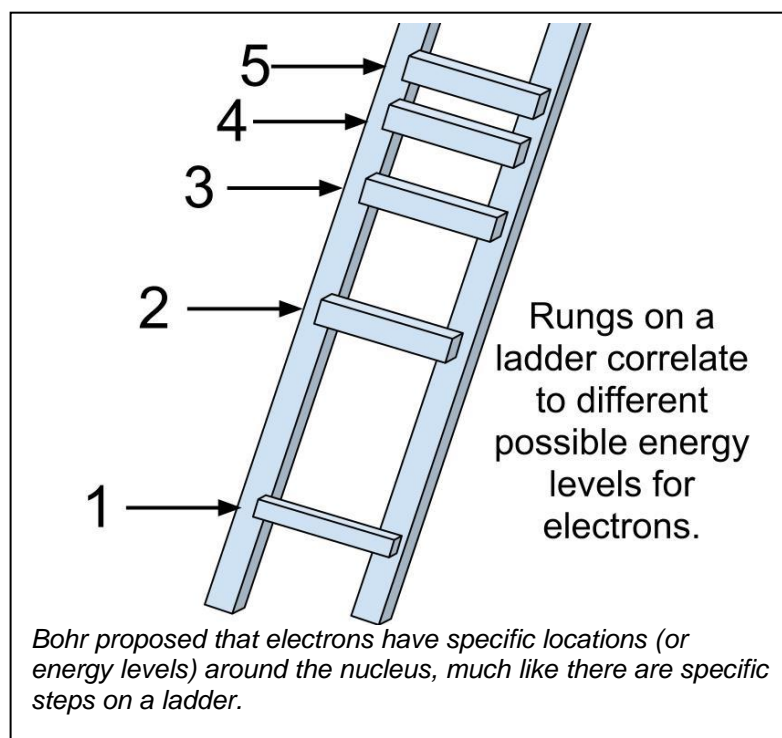
You may have heard or read about scientists discussing what elements are present in the sun or some more distant star. How could scientists know what elements were present in a place no one has ever been? Scientists determine what elements are present in distant stars by analyzing the light that comes 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 star contains hydrogen.

Atomic (Emission) Spectrum: The unique pattern of light given off when an element is heated or electric current passed through it.

2.5: Light and Atomic Spectra

Bohr's Model

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**.



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. Atoms would be like a ladder that had rungs only at certain heights. The only way an electron can be on that ladder is to be on one of the rungs. The only way an electron could move up or down would be to move to one of the other rungs. 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.

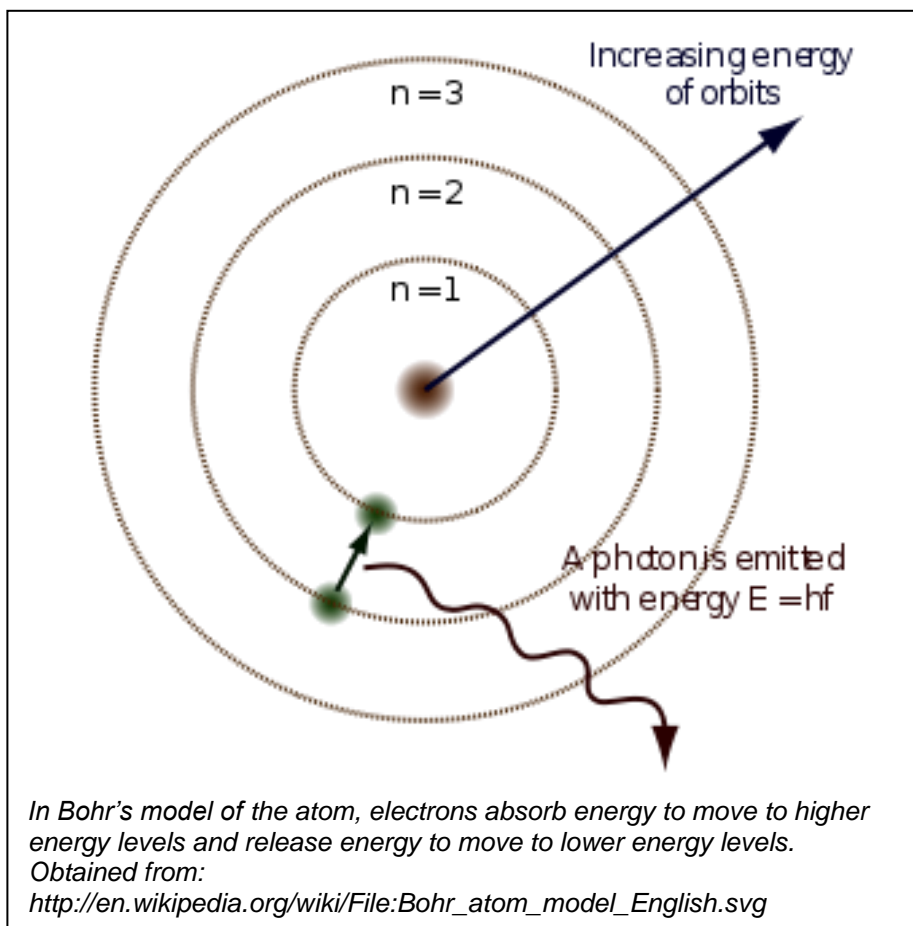
Energy Level: Acceptable distances an electron can occupy from the nucleus
Quantized: Only allowing specific amounts of energy.

2.5: Light and Atomic Spectra

Explaining 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 they had absorbed in the form of a piece of light, called a **photon**, 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.



Ground State: having the lowest energy possible
Excited State: having a higher energy level
Photon: a quantized piece of light

2.6: Electrons in Atoms

Electron Energy Levels

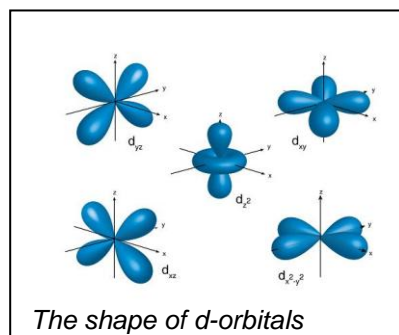
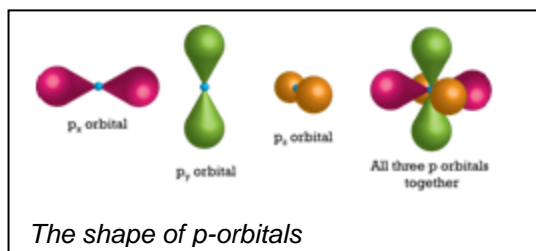
Objectives:

-Describe the energy levels, sublevels, and orbitals in which electrons are occupied

-Write an electron configuration for an element in ground state

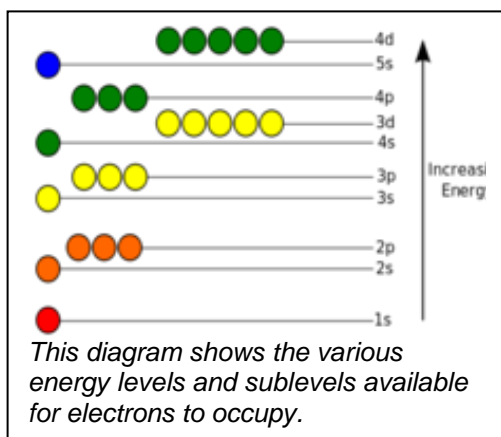
-Draw an orbital diagram for an element in ground state.

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. All put together, electrons are various distances from the nucleus (energy level), occupying various shapes (sublevels) and different specific areas and orientations (orbitals) within atoms.



The following table summarizes the possible energy levels and sublevels that compose each sublevel, how many orbitals make up 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 2p	1 3	2 6
3	3s 3p 3d	1 3 5	2 6 10
4	4s 4p 4d 4f	1 3 5 7	2 6 10 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?

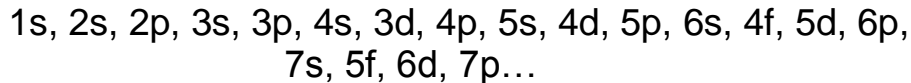
When determining where the electrons in an atom are located, a couple of rules must be followed:

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

2.6: Electrons in Atoms

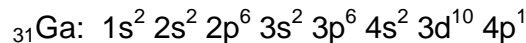
The Electron Configuration

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 the distance (energy level) and shapes which an electron occupies is taken into account, the actual filling order is:



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:



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.

The diagram shows a periodic table with color-coded blocks representing different sublevels. Arrows indicate the order in which these sublevels are filled. The blocks are:

- 1s:** Hydrogen (H) and Helium (He).
- 2s, 3s, 4s, 5s, 6s, 7s:** The s-block on the left side of the periodic table.
- 2p, 3p, 4p, 5p, 6p, 7p:** The p-block on the right side of the periodic table.
- 3d, 4d, 5d, 6d:** The d-block in the middle of the periodic table.
- 4f, 5f:** The f-block at the bottom of the periodic table.

 The filling order follows the sequence: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p.

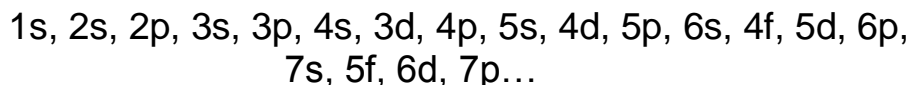
http://en.wikipedia.org/wiki/File:Electron_Configuration_Table.jpg

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

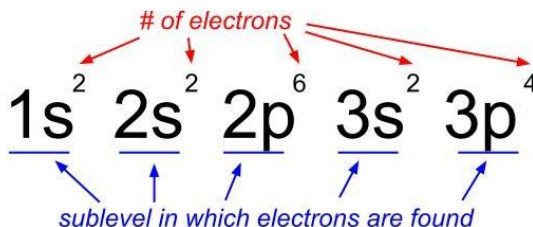
2.6: Electrons in Atoms

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 1*s*. Moving onto the second row, 2*s* is filled followed by 2*p*. Continuing with the filling order you generate the list:



An electron configuration lists the sublevels the electrons occupy and the number of electrons in each of those sublevels, written as superscripts. For example, sulfur's electron configuration is:



Example: Write the electron configurations for

(a) potassium, K

(b) arsenic, As

(c) phosphorus, P

Solution:

(a) Potassium atoms have 19 protons and, therefore, 19 electrons. Using our periodic table, we see that the first sublevel to fill is 1*s*, which can hold 2 of those 19 electrons. Next to fill is 2*s*, which also holds 2 electrons. Then comes 2*p* which holds 6. We keep filling up the sublevels until all 19 of the electrons have been placed in the lowest energy level possible. 4*s* 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) ${}_{33}\text{As}$: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^3$

(c) ${}_{15}\text{P}$: $1s^2 2s^2 2p^6 3s^2 3p^3$

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.

Electron Configuration: Lists the number of electrons in each sublevel in an atom.

2.6: Electrons in Atoms

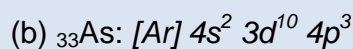
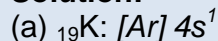
Abbreviated Electron Configurations

As the electron configurations become longer and longer, it becomes tedious to write them out. A shortcut has been devised. The electron configuration for potassium is $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$, which is the same as argon's except that it has one more electron. 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 $[Ar]4s^2 3d^1$. Abbreviated electron configurations always put the previous noble gas symbol on the periodic table in brackets and then add any additional electrons to get to the desired element.

Example: Write the abbreviated electron configurations for

- (a) potassium, K
- (b) arsenic, As
- (c) phosphorus, P

Solution:



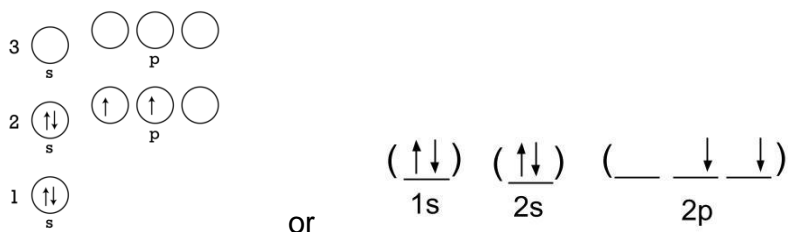
Orbital Diagrams

An **orbital diagram** is similar to writing an electron configuration except that each orbital is drawn with a line or circle and the electrons are drawn within their orbitals. "s" sublevels are shown with one line or circle because they have 1 orbital. "p" sublevels are shown with 3 circles because they have 3 orbitals. "d" sublevels and "f" sublevels are shown with five or seven circles respectively for the same reason.

Rules for Drawing Orbital Diagrams:

- 1) Electrons enter the lowest energy orbital available. This order can be obtained by looking at an electron configuration.
- 2) No more than two electrons can be placed in any orbital and they must have opposite spins (the arrows showing electrons must point in opposite directions if they are sharing an orbital – line or circle)
- 3) Before a second electron pairs up in a single orbital, all the orbitals of that sub-level must contain at least one electron. This is known as **Hund's Rule**.

Carbon has the electron configuration of $1s^2 2s^2 2p^2$. Its orbital diagram would be:



2.6: Electrons in Atoms

A carbon atom has six protons and six electrons. When placing the electrons in the appropriate orbitals, the first two electrons go into the 1s orbital, the second two go into the 2s orbital, and the last two go into 2p orbitals. Since electrons repel each other, the electrons in the 2p orbitals go into separate orbitals before pairing up. Hund's rule is a statement of this principle, stating that electrons will spread out into equal-energy orbitals before pairing up. Note that it doesn't matter which two of the three orbitals these unpaired electrons are placed, as long as they are in separate orbitals. Also, note that it doesn't matter the direction of the electron's spin (arrow) as long as the unpaired electrons match each other.

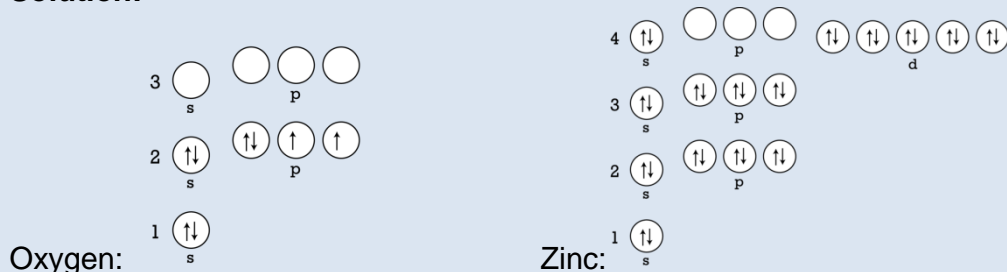
Element Symbol	Electron Configuration Code	Orbital Representation
H	$1s^1$	1 \uparrow s
He	$1s^2$	1 $\uparrow\downarrow$ s
Li	$1s^2 2s^1$	2 \uparrow ○ ○ ○ s p 1 $\uparrow\downarrow$ s
Be	$1s^2 2s^2$	2 $\uparrow\downarrow$ ○ ○ ○ s p 1 $\uparrow\downarrow$ s
B	$1s^2 2s^2 2p^1$	2 $\uparrow\downarrow$ \uparrow ○ ○ s p 1 $\uparrow\downarrow$ s
C	$1s^2 2s^2 2p^2$	2 $\uparrow\downarrow$ \uparrow \uparrow ○ s p 1 $\uparrow\downarrow$ s
N	$1s^2 2s^2 2p^3$	2 $\uparrow\downarrow$ \uparrow \uparrow \uparrow s p 1 $\uparrow\downarrow$ s

Orbital diagrams of elements 1-7

Example: Draw the orbital diagram for (a) oxygen,)

(b) zinc, Zn

Solution:



Abbreviated Electron Configuration: Gives the previous noble gas element listed in brackets and then lists the number of electrons in each sublevel after that point.

Orbital Diagram: A diagram showing the orbitals and spins of electrons in an atom

Chapter 2 Summary

2.1: Early Ideas of Atoms

- 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. Atoms are indivisible.
 2. All atoms of a given element are identical in mass and other properties. The atoms of different elements differ in mass and other properties.
 3. Atoms of one element can combine with atoms of another element to form “compounds”. During a chemical reaction, atoms are rearranged, but they do not break apart, nor are they created or destroyed.

2.2: Further Understanding of the Atom

- Dalton’s Atomic Theory wasn’t entirely correct. It turns out that atoms can be divided into smaller subatomic particles.
- Through his cathode-ray tube experiment, Thomson found that all atoms have negatively charged particles called electrons
- 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.
- Robert Millikan was able to determine the charge and mass of the electron.
- Positively charged particles, called protons, were also found in atoms.
- 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.

2.3: Protons, Neutrons, and Electrons in Atoms

- 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).
- 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).
- 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.

Chapter 2 Summary

- 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.
- The total of an atom's protons and neutrons is its mass number.
- Isotopes are atoms of the same element with different masses due to different numbers of neutrons.
- The nuclear symbol shows the mass number, atomic number, and chemical symbol for a specific isotope.
- In a neutral atom, the number of protons equals the number of electrons. However, when an ion is formed, electrons are gained (resulting in an overall negative charge) or lost (resulting in an overall positive charge).
- Various atomic models have been developed to describe observations made by science. Rutherford's nuclear model was incomplete as it could not describe atomic spectrum of elements.

2.4: Atomic Mass

- 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.
- In addition to the element's symbol, most periodic tables will also contain the element's atomic number, and element's atomic mass.

2.5: Light and Atomic Spectra

- 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) and energy
- 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.
- 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.

2.6: Electrons in Atoms

- Bohr's model was only successful in calculating energy levels for the hydrogen atom. Other scientists were able to further develop the atomic model to account for all elements.
- Electrons are located in orbitals, in various sublevels and energy levels of atoms
- Electrons will occupy the lowest energy level possible.

Chapter 2 Summary

- It is possible to write the electron configuration, which gives the number of electrons in each sublevel of an element, using a periodic table
- Abbreviated electron configurations are shortened versions of electron configurations in which the previous noble gas symbol is given in brackets followed by all remaining electrons and sublevels.
- An orbital diagram is a pictorial representation of electrons in various sublevels and orbitals.

2.1: Early Ideas of Atoms Review Questions

- 1) 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'?
- 2) 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.
- 3) When will Dalton's Atomic Theory become a law?
 - a) When it has enough evidence to support it
 - b) When scientists everywhere agree with it.
 - c) It cannot become a law.
 - d) When the idea of atoms has been proven beyond all doubt.
- 4) What does the law of conservation of mass state? How does the idea that everything is made of atoms explain this law?
- 5) What does the law of definite proportions state? How does the idea that everything is made of atoms explain this law?
- 6) What does the law of multiple proportions state? How does the idea that everything is made of atoms explain this law?
- 7) 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?

Chapter 2 Summary

2.2: Further Understanding of the Atom Review

Questions

- 8) Dalton's atomic theory originally stated: "Matter is made of tiny particles called atoms. Atoms are indivisible." How should this statement be rewritten to reflect our current understanding of atoms?

#9-13: Match each conclusion regarding subatomic particles and atoms with the observation/data that supports it.

Conclusion	Observations
9) All atoms have electrons	a. Most alpha particles shot at gold foil go straight through, without any change in their direction.
10) Atoms are mostly empty space.	b. A few alpha particles shot at gold foil bounce in the opposite direction.
11) Electrons have a negative charge	c. Some alpha particles (with positive charges) when shot through gold foil bend away from the gold.
12) 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.
13) 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.

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.

- 14) What is the main dispute between the two scientists' theories?
- 15) 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.
- 16) If a positively charged particle was shot at a thin sheet of gold foil, what would the first scientist predict to happen? What would the second scientist predict?

Chapter 2 Summary

2.3: Protons, Neutrons, and Electrons in Atoms Review Questions

#17-20: Label each of the following statements as true or false.

- 17) The nucleus of an atom contains all of the protons in the atom.
- 18) The nucleus of an atom contains all of the electrons in the atom.
- 19) Neutral atoms must contain the same number of neutrons as protons.
- 20) Neutral atoms must contain the same number of electrons as protons.

#21-24: For each of the following statements, indicate whether the statement is describing protons, neutrons, or electrons.

- 21) Has a charge of +1
- 22) Has the least mass
- 23) Used to determine the element
- 24) Is neither attracted to nor repelled by charged objects

#25-29: Indicate whether each statement is true or false.

- 25) An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- 26) A neutral atom with 4 protons must have 4 electrons.
- 27) An atom with 7 protons and 7 neutrons will have a mass number of 14.
- 28) An atom with 7 protons and 7 neutrons will have an atomic number of 14.
- 29) A neutral atom with 7 electrons and 7 neutrons will have an atomic number of 14.

#30-34: 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
30) an atom with 2 protons and 1 neutron	a. a C (carbon) atom with 6 neutrons
31) a Be (beryllium) atom with 5 neutrons	b. an atom with 2 protons and 2 neutrons
32) 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
33) an atom with 1 proton and a mass number of 1	d. an atom with an atomic number of 1 and 1 neutron
34) an atom with an atomic number of 7 and 7 neutrons	e. an atom with an atomic number of 4 and 6 neutrons

Chapter 2 Summary

#35-36: Write the nuclear symbol for each element described:

35) 32 neutrons in an atom with mass number of 58

36) An atom with 10 neutrons and 9 protons.

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

37) ${}^4_2\text{He}$

38) Sodium-23

39) ${}^1_1\text{H}$

40) Iron-55

41) ${}^{37}_{17}\text{Cl}$

42) Boron-11

43) ${}^{238}_{92}\text{U}$

44) Uranium-235

45) ${}^{27}_{13}\text{Al}^{+3}$

46) ${}^{19}_9\text{F}^-$

47) ${}^{66}_{30}\text{Zn}^{2+}$

48) ${}^{34}_{16}\text{S}^{2-}$

2.4: Atomic Mass Review Questions

49) Boron has two naturally occurring isotopes B-10 and B-11. Which isotope is more abundant? Justify your answer.

50) 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-65 and have a mass of 64.93amu. What is the atomic mass of copper?

51) 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: Light and Atomic Spectra Review Questions

52) Which color of visible light has the longer wavelength, red or blue?

53) What is the relationship between the energy of electromagnetic radiation and the wavelength of that radiation? As wavelength increases, the energy _____.

Chapter 2 Summary

54) Of the two waves drawn below, which one has the most energy? How do you know?



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

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

57) Bohr's model of the atom is frequently referred to as the "quantum model". What does it mean to be quantized? How are electrons in atoms quantized?

58) 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?

59) According to Bohr's theory, how can an electron gain or lose energy?

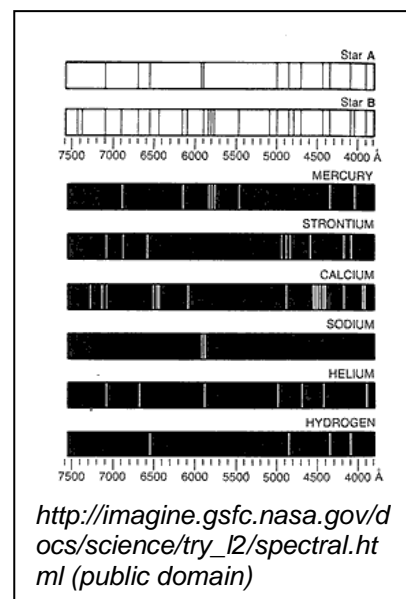
60) What happens when an electron in an excited atom returns to its ground level?

61) Why do electrons of an element release only a specific pattern of light? Why don't they produce all colors of light?

#62-63: Use the given spectra to answer the following questions.

62) Which elements make up star A?

63) Which elements make up star B?



Chapter 2 Summary

2.6: Electrons in Atoms Review Questions

- 64) Which principal energy level holds a maximum of eight electrons?
- 65) Which sub-energy level holds a maximum of six electrons?
- 66) Which sub-energy level holds a maximum of ten electrons?
- 67) In which energy level and sub-level of the carbon atom is the outermost electron located?
- 68) How many electrons are in the 2p sub-energy level of a neutral nitrogen atom?
- 69) Which element's neutral atoms will have the electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^1$?
- 70) What energy level and sub-level immediately follow 5s in the filling order?
- 71) What is the outermost energy level and sub-level used in the electron configuration of potassium?

#72-78: Write electron configurations for each of the following neutral atoms:

72) Magnesium

73) Nitrogen

74) Yttrium

75) Tin

76) Krypton

77) Cesium

78) Uranium

#79-84: Write the abbreviated electron configuration for each of the following neutral atoms:

79) Fluorine

80) Aluminum

81) Titanium

82) Arsenic

83) Rubidium

84) Carbon

Chapter 2 Summary

*#85-88: Draw the orbital diagram for each of the following elements.
(*Hint: You wrote the electron configuration for many of these elements in
the previous problems.)*

85) Magnesium

86) Nitrogen

87) Tin

88) Arsenic

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Chapter 3: Organizing Elements

Reihen	Gruppe I. — R ⁰	Gruppe II. — R ⁰	Gruppe III. — R ⁰ ³	Gruppe IV. RH ⁴ R ⁰ ⁴	Gruppe V. RH ⁵ R ⁰ ⁵	Gruppe VI. RH ⁶ R ⁰ ⁶	Gruppe VII. RH R ⁰ ⁷	Gruppe VIII. — R ⁰ ⁸
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	Su=118	Sb=122	Tc=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	—	— — — —

What is this? This is actually an early form of something all of us would recognize – the Periodic Table of Elements.

- How and why was the periodic table made?
- What other ways have elements been classified and arranged?
- What patterns and information can we get from this chart?

This chapter will address these questions and others.

3.1: The Development of the Periodic Table

Early Attempts to Organize Elements

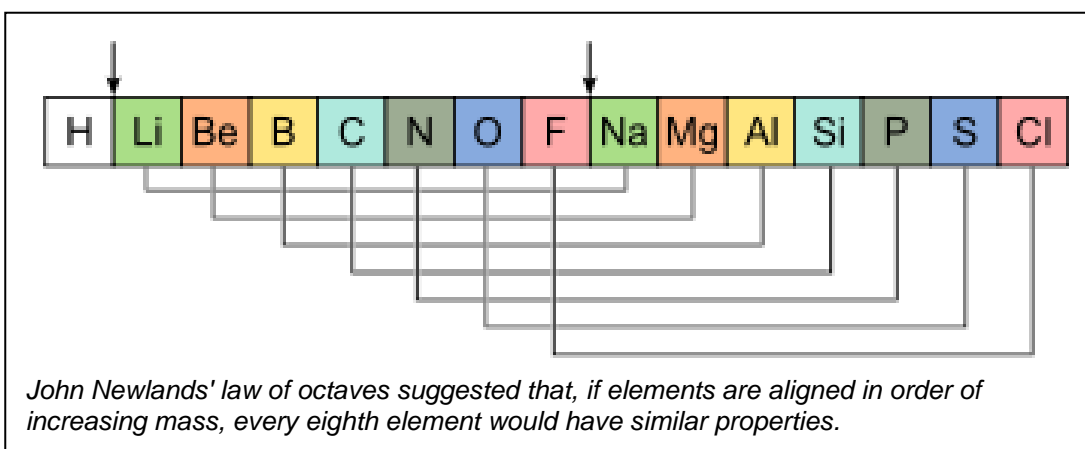
By the year 1700, only a handful of elements had been identified and isolated. Several of these, such as copper and lead, had been known since ancient times. As scientific methods improved, the rate of discovery dramatically increased. With the ever-increasing number of elements, chemists recognized that there may be some kind of systematic way to organize the elements. The question was: how?

A logical way to begin to group elements together was by their chemical properties. In other words, putting elements in separate groups based on how they reacted with other elements. In 1829, a German chemist, Johann Dobereiner (1780-1849), placed various groups of three elements into groups called triads. One such triad was lithium, sodium, and potassium. Triads were based on both physical as well as chemical properties. Dobereiner found that the atomic masses of these three elements, as well as other triads, formed a pattern. When the atomic masses of lithium and potassium were averaged together

$\frac{6.94+39.10}{2} = 23.02$, it was approximately equal to the atomic mass of sodium (22.99). These three elements also displayed similar chemical reactions, such as vigorously reacting with the members of another triad: chlorine, bromine, and iodine. While Dobereiner's system would pave the way for future ideas, a limitation of the triad system was that not all of the known elements could be classified in this way.

Objectives

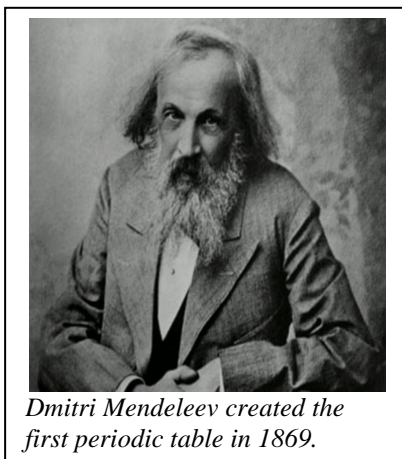
- Give advantages Mendeleev's periodic table has over other methods of organizing elements
- Identify ways in which the current periodic table differs from its original form
- Describe the similarities between elements within the same group or family



English chemist John Newlands (1838-1898) ordered the elements in increasing order of atomic mass and noticed that every eighth element exhibited similar properties. He called this relationship the "Law of Octaves." Unfortunately, there were some elements that were missing and the law did not seem to hold for elements that were heavier than calcium. Newlands's work was largely ignored and even ridiculed by the scientific community in his day. It was not until years later that another more extensive periodic table effort would gain much greater acceptance and the pioneering work of John Newlands would be appreciated.

3.1: The Development of the Periodic Table

Mendeleev's Table of Elements



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

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

with water. When it reacts with water, it produces hydrogen gas and lithium hydroxide. As we proceed through the elements with increasing mass, we will come to the element sodium. Sodium is a shiny metal, soft

Mendeleev first organized elements in order of increasing mass and then put them in groups of similar properties.

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

Reihen	Gruppe I. — R ⁰	Gruppe II. — R ⁰	Gruppe III. — R ⁰	Gruppe IV. RH ⁴ R ⁰	Gruppe V. RH ⁵ R ⁰	Gruppe VI. RH ⁶ R ⁰	Gruppe VII. RH ⁷ R ⁰	Gruppe VIII. — R ⁰
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=65.
5	(Ca=63)	Zn=65	—=68	—=72	As=75	Se=78	Br=80	
6	Rb=85	Sr=87	?Y=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

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.

3.1: The Development of the 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).

Property	Mendeleev's prediction for Eka-aluminium	Actual properties of Gallium
atomic mass	68	69.72
density (g/cm ³)	6.0	5.904
melting point (°C)	Low	29.78
oxide's formula	Ea ₂ O ₃ (density - 5.5 g cm ⁻³) (soluble in both alkalis and acids)	Ga ₂ O ₃ (density - 5.88 g cm ⁻³) (soluble in both alkalis and acids)
chloride's formula	Ea ₂ Cl ₆ (volatile)	Ga ₂ Cl ₆ (volatile)

Property	Mendeleev's predictions for Eka-silicon	Actual properties of Germanium
atomic mass	72	72.61
density (g/cm ³)	5.5	5.35
melting point (°C)	high	947
color	grey	grey
oxide type	refractory dioxide	refractory dioxide
oxide density (g/cm ³)	4.7	4.7
oxide activity	feebly basic	feebly basic
chloride boiling point	under 100°C	86°C (GeCl ₄)
chloride density (g/cm ³)	1.9	1.9

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 what 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.

3.1: The Development of the Periodic Table

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.

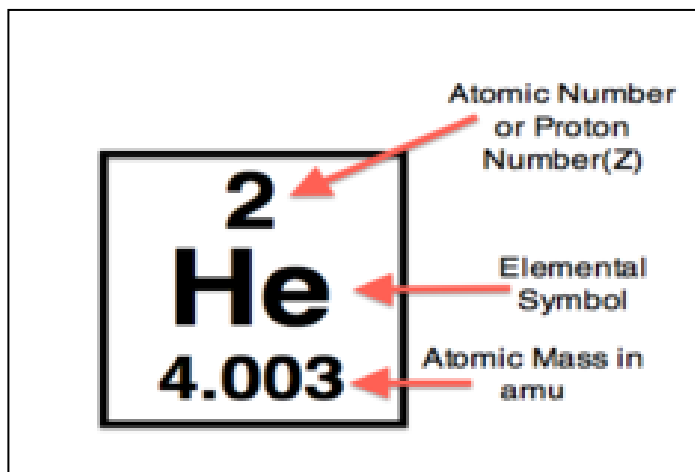
As previously noted, Mendeleev organized elements in order of increasing atomic mass, with some problems in the order of masses. In 1914 Henry Moseley found a relationship between an element's X-ray wavelength and its atomic number, and therefore organized the table by nuclear charge (or atomic number) rather than atomic weight. The **Periodic Law** states that properties of elements are a function of their atomic number. 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.



Periodic Law: Properties of elements are a function of their atomic number
Group: vertical columns of elements, in which the elements have similar chemical and physical properties.

3.1: The Development of the Periodic Table

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. The elements in 1A are all very reactive and form compounds in the same ratios with similar properties with other elements. 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.

1A 1	2A 2																	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18		
Alkali Metals	Alkaline Earth Metals																						Halogens	Noble Gases	
		Transition Metals																							

It is important to recognize a couple of other important groups on the periodic table by their group name. Group 7A (or 17) elements are also called **halogens**. This group contains very reactive nonmetals elements.

The **noble gases** are in group 8A. These elements also have similar properties to each other, the most significant property being that they are extremely unreactive rarely forming compounds. We will learn the reason for this later, when we discuss how compounds form. The elements in this group are also gases at room temperature.

An alternate numbering system numbers all of the *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. 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.

3.2: Valence Electrons

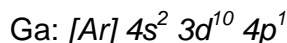
Valence Electrons

Objectives:

- Determine the number of valence electrons in an atom.
- Describe the significance of valence electrons in determining chemical properties of elements

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.



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 chemical reactions, but not the electrons in the filled *3d* subshell.

Gallium therefore has three valence electrons: 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.

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.

Counting Valence Electrons																																																																																								
1	2		skip a few												3	4	5	6	7	8																																																																				
1 H	3 Li	4 Be	11 Na	12 Mg	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr	53 I	54 Xe	85 At	86 Rn	118 Uuo																																																													
2 He	10 Ne	18 Ar	36 Kr	54 Xe	86 Rn	118 Uuo	119 Uut	120 Uuq	121 Uuq	122 Uuq	123 Uuq	124 Uuq	125 Uuq	126 Uuq	127 Uuq	128 Uuq	129 Uuq	130 Uuq	131 Uuq	132 Uuq	133 Uuq	134 Uuq	135 Uuq	136 Uuq	137 Uuq	138 Uuq	139 Uuq	140 Uuq	141 Uuq	142 Uuq	143 Uuq	144 Uuq	145 Uuq	146 Uuq	147 Uuq	148 Uuq	149 Uuq	150 Uuq	151 Uuq	152 Uuq	153 Uuq	154 Uuq	155 Uuq	156 Uuq	157 Uuq	158 Uuq	159 Uuq	160 Uuq	161 Uuq	162 Uuq	163 Uuq	164 Uuq	165 Uuq	166 Uuq	167 Uuq	168 Uuq	169 Uuq	170 Uuq	171 Uuq	172 Uuq	173 Uuq	174 Uuq	175 Uuq	176 Uuq	177 Uuq	178 Uuq	179 Uuq	180 Uuq	181 Uuq	182 Uuq	183 Uuq	184 Uuq	185 Uuq	186 Uuq	187 Uuq	188 Uuq	189 Uuq	190 Uuq	191 Uuq	192 Uuq	193 Uuq	194 Uuq	195 Uuq	196 Uuq	197 Uuq	198 Uuq	199 Uuq	200 Uuq

The number of valence electrons in an atom can be easily found by counting the *s* and *p* columns in the periodic table.

Valence Electrons: The electrons in the outermost energy level in an atom.

3.3: Metals, Nonmetals, and Metalloids

Metals, Non-metals, and Metalloids

Our periodic table places elements with similar properties in the same vertical column or the same family. Another way to group elements is into three more basic larger classes – metals, nonmetals, and metalloids..

1 H																	2 He
3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu	86 Rn
87 Fr	88 Ra	89 Ac	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr	

Objectives:
-Indicate whether a given element is a metal, nonmetal, or metalloid
-Compare and contrast properties of metals, nonmetals, and metalloids.

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

Most of the chemical elements are metals. Most metals are shiny, very dense, and have 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 heat and electricity. In chemical reactions, they tend 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 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.

Malleable: bendable; shapeable
Ductile: can be stretched into a wire
Conductor: allows heat and electricity to travel through
Insulator: does not allow heat and electricity to travel through.

3.4: Periodic Trends

Periods of the Periodic Table

Objectives:

- Explain how the periodic table is periodic
- Summarize the patterns in atomic radius, ionization energy, and electronegativity of elements on the periodic table.
- Account for these patterns in terms of atomic structure.

If you can locate an element on the Periodic Table, you can use the element's position to figure out the energy level of the element's valence electrons. A **period** is a horizontal row of elements on the periodic table. 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.

Periods of the Periodic Table (energy level of valence electrons)																			
1 →	1																2		
	H																He		
2 →	3	4												5	6	7	8	9	10
	Li	Be												B	C	N	O	F	Ne
3 →	11	12												13	14	15	16	17	18
	Na	Mg												Al	Si	P	S	Cl	Ar
4 →	19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36	
	K	Ca	Sc	Ti	V	Cr	Mn	Fe	Co	Ni	Cu	Zn	Ga	Ge	As	Se	Br	Kr	
5 →	37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54	
	Rb	Sr	Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	
6 →	55	56	57	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86	
	Cs	Ba	La	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	
7 →	87	88	89	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118	
	Fr	Ra	Ac	Rf	Db	Sg	Bh	Hs	Mt	Ds	Rg	Cn	Uut	Fl	Uup	Lv	Uuo		

CC – Tracy Poulsen

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 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 two main factors. These factors are:

- The number of protons in the nucleus (called the **nuclear charge**).
- The energy levels of the valence electrons (and the number of electrons in the outer energy level). The energy level of the valence electrons is equal to the period of the element on the periodic table.

Period: A horizontal row on the periodic table; equal to the energy level of the valence electrons of an element

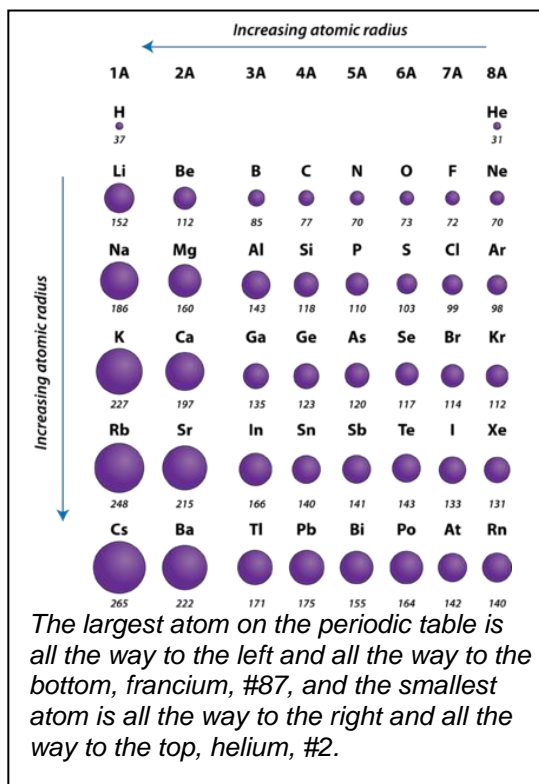
Nuclear Charge: The total charge of the nucleus of an atom; equal to the number of protons

3.4: Periodic Trends

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 elements in the same group. Each atom in group 1 (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). When we examine the energy levels for period 2, we find that the outermost energy level does not change – all of them have valence electrons in the second energy level. Looking at the elements in period 2, the number of protons increases from lithium with three protons, to fluorine with nine protons. 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*.

Example: Which of the following has a greater radius? Why?

- (a) As or Sb
- (b) Ca or K
- (c) Polonium or Sulfur

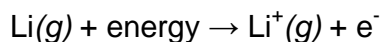
Solution:

- (a) Sb. It is below As, so its valence electrons are in a higher energy level.
- (b) K because it has valence electrons in the same energy level, but fewer protons to attract the electrons.
- (c) Polonium because it has valence electrons in a higher energy level.

3.4: Periodic Trends

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:



With the addition of energy, a lithium ion can be formed from the lithium atom by losing one electron. 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 increases. 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 greater ionization energy? Why?

- (a) As or Sb
- (b) Ca or K
- (c) Polonium or Sulfur

Solution:

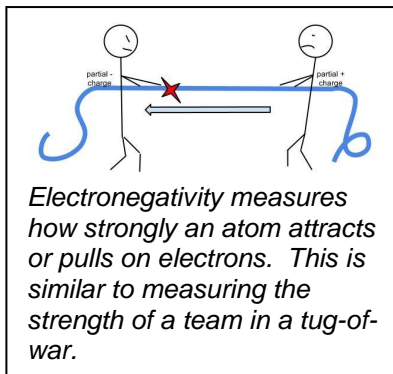
- (a) As, because its valence electrons are in a lower energy level.
- (b) Ca because it has more protons to attract its valence electrons.
- (c) S because its valence electrons are in a lower energy level.

Ionization Energy: The energy required to remove a valence electron in an atom.

3.4: Periodic Trends

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.

Pauling Electronegativity Values																																		
1 H 2.20																	5 B 2.04	6 C 2.55	7 N 3.04	8 O 3.44	9 F 3.98													
3 Li 0.98	4 Be 1.57																	13 Al 1.61	14 Si 1.90	15 P 2.19	16 S 2.58	17 Cl 3.16												
11 Na 0.93	12 Mg 1.31																	19 K 0.82	20 Ca 1.00	21 Sc 1.36	22 Ti 1.54	23 V 1.63	24 Cr 1.66	25 Mn 1.55	26 Fe 1.83	27 Co 1.88	28 Ni 1.91	29 Cu 1.90	30 Zn 1.65	31 Ga 1.81	32 Ge 2.01	33 As 2.18	34 Se 2.55	35 Br 2.96
37 Rb 0.82	38 Sr 0.95	39 Y 1.22	40 Zr 1.33	41 Nb 1.6	42 Mo 2.16	43 Tc 1.9	44 Ru 2.2	45 Rh 2.28	46 Pd 2.20	47 Ag 1.93	48 Cd 1.69	49 In 1.78	50 Sn 1.96	51 Sb 2.05	52 Te 2.1	53 I 2.68																		
55 Cs 0.79	56 Ba 0.89	57 La 1.1	72 Hf 1.3	73 Ta 1.5	74 W 2.36	75 Re 1.9	76 Os 2.2	77 Ir 2.20	78 Pt 2.28	79 Au 2.54	80 Hg 2.00	81 Tl 1.82	82 Pb 2.33	83 Bi 2.02	84 Po 2.0	85 At 2.2																		
87 Fr 0.7	88 Ra 0.9																																	

For this class, you do not need to memorize the electronegativity values. You do, however, need to understand the overall pattern that as you go from left to right across a period, the elements are more electronegative. As you move down a vertical column, the elements are less electronegative.

Electronegativity: The relative attraction an atom has for electrons in a bond.

Chapter 3 Summary

3.1: The Development of the Periodic Table

- 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.
- 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

3.2: Valence Electrons

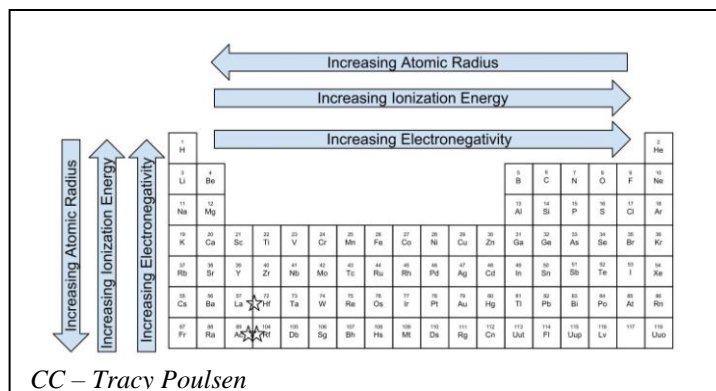
- Valence electrons are the electrons in the outermost principal quantum level of an atom.
- The number of valence electrons is important, because the chemical behavior of an element depends primarily by the arrangement of the electrons in the valence shell.

3.3: Metals, Nonmetals, and Metalloids

- 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.

3.4: Periodic Trends

- The horizontal rows on the periodic table are called periods.
- 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.

Chapter 3 Summary

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: The Development of the Periodic Table Review

Questions

- 1) What general organization did Mendeleev use when he constructed his table?
- 2) List two advantages Mendeleev’s organization had over other methods of organizing elements, such as Newlands’.
- 3) How would the discovery of new elements have affected Newland’s Law of Octaves? What about Mendeleev’s arrangement of the elements?
- 4) What discovery did Henry Moseley make that changed how we currently recognize the order of the elements on the periodic table?
- 5) List three ways in which our current periodic table differs from the one originally made by Mendeleev.

Multiple Choice

- 6) Which of the following elements is in the same family as fluorine?
 - a) silicon
 - b) antimony
 - c) iodine
 - d) arsenic
 - e) None of these.
- 7) Elements in a _____ have similar chemical properties.
 - a) period
 - b) family
 - c) both A and B
 - d) neither A nor B

Chapter 3 Summary

- 8) Which of the following elements would you expect to be most similar to carbon?
- Nitrogen
 - Boron
 - Silicon

#9-12: Give the name of the family in which each of the following elements is located:

9) astatine

10) krypton

11) barium

12) francium

#13-18: Which family is characterized by each of the following descriptions?

13) A very reactive family of nonmetals

14) A nonreactive family of nonmetals

15) Forms colorful compounds

16) A very reactive family of metals

3.2: Valence Electrons Review Questions

17) How many valence electrons are present in the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^3$? Circle the valence electrons in the electron configuration.

18) How many valence electrons are present in the following electron configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^1$? Circle the valence electrons in the electron configuration.

#21-28: For each of the following atoms, indicate the total number of valence electrons in each atom:

19) Fluorine

23) Aluminum

20) Bromine

24) Gallium

21) Sodium

25) Argon

22) Cesium

Chapter 3 Summary

3.3: Metals, Nonmetals, and Metalloids Review

Questions

#28-33: Label each of the following elements as a metal, nonmetal, or metalloid.

26) Carbon

29) Plutonium

27) Bromine

30) Potassium

28) Oxygen

31) Helium

#34-40: Given each of the following properties, label the property of as that of a metal, nonmetal, or metalloid.

32) Lustrous

36) Insulators

33) Semiconductors

37) Conductors

34) Brittle

38) Along the staircase

35) Malleable

39) 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? Site specific properties in your explanation.

3.4: Periodic Trends Review Questions

Multiple choice

40) 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.

41) Which of the following would have the largest ionization energy?

- a) Na
- b) Al
- c) H
- d) He

42) Which of the following would have the smallest ionization energy?

- a) K
- b) P
- c) S
- d) Ca

Chapter 3 Summary

- 43) Arrange the following in order of increasing atomic radius: Tl, B, Ga, Al, In.
- 44) Arrange the following in order of increasing atomic radius; Ga, Sn, C.
- 45) Which of the following would have a smaller radius: indium or gallium? Why?
- 46) Which of the following would have a smaller radius: potassium or cesium? Why?
- 47) Which of the following would have a smaller radius: titanium or polonium? Why?
- 48) Define ionization energy.
- 49) Place the following elements in order of increasing ionization energy: Na, S, Mg, Ar
- 50) Explain why bromine atoms have higher ionization energy than iodine atoms.
- 51) Which of the following would have a lower ionization energy – potassium or sodium? Why?
- 52) Define electronegativity.
- For each pair of elements, choose the element that has the lower electronegativity.*
- 53) Li or N
- 54) Cl or Na
- 55) Na or K
- 56) Mg or F

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Chapter 4

Describing Compounds



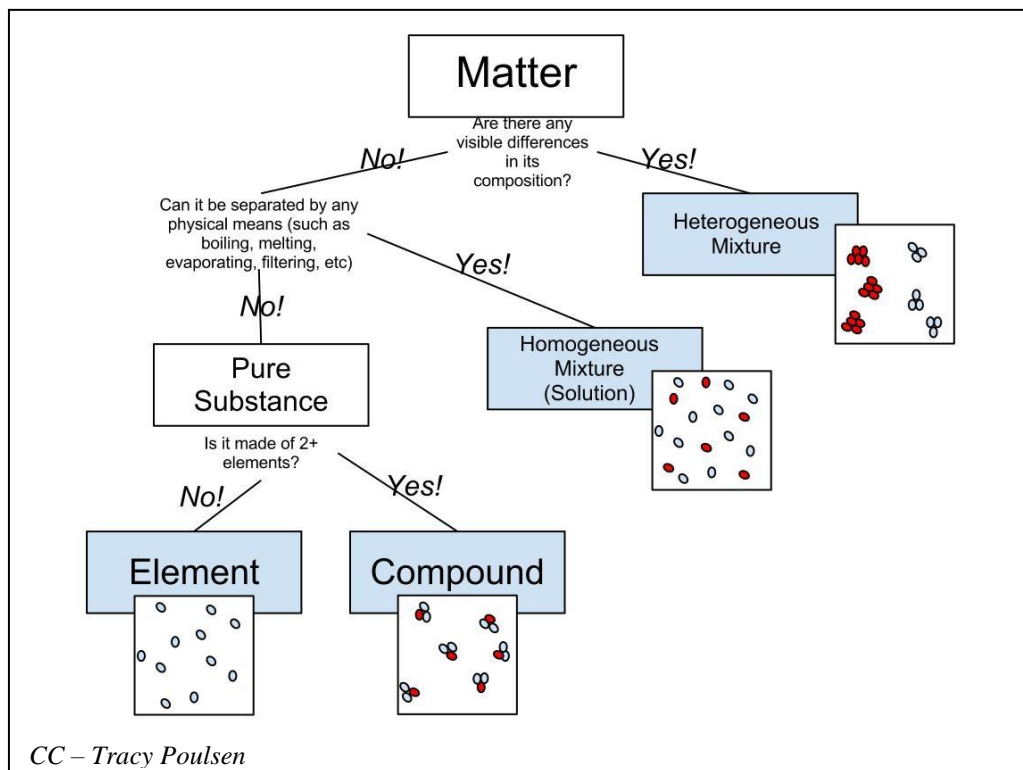
Did you ever play the card game called go fish? Players try to form groups of cards of the same value, such as four sevens, with the cards they are dealt or by getting cards from other players or the deck. This give and take of cards is a simple analogy for the way atoms give and take valence electrons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

4.1: Classifying Matter Substances and Mixtures

Objectives:

- Distinguish between elements, compounds, and mixtures.
- Interpret a chemical formula
- Describe how the properties of a compound compare to properties of elements and other compounds.

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.



Elements contain only one type of atom – all of the atoms have the same number of protons. Compounds, on the other hand, have two or more different types of elements bonded together chemically, forming a substance with different properties.

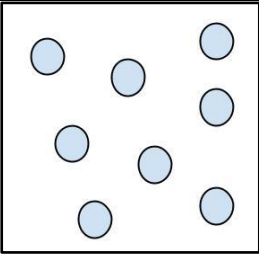
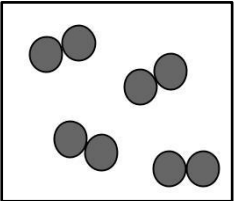
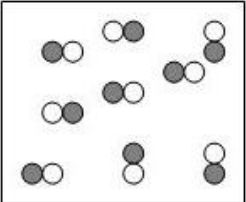
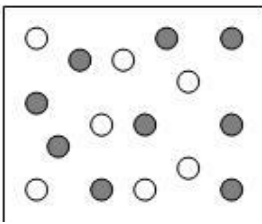
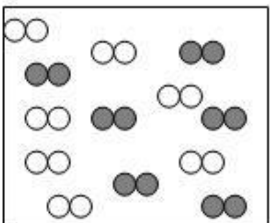
Mixtures are physical combinations of two or more elements and/or compounds. The term “physical combination” refers to mixing two different substances together where the substances do not chemically react. The physical appearance of the substances may change but the atoms and/or molecules in the substances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

4.1: Classifying Matter

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)		<i>Matter with only one type of atom is called an element.</i>
Pure substance (element)		<i>Although the chlorine atoms are bonded in pairs, since there is only one type of atom, this is an element.</i>
Pure substance (compound)		<i>When two or more elements are bonded together, a compound is produced.</i>
Mixture		<i>When two or more pure substances (in this case, two elements) are combined, but not bonded together, a mixture is produced.</i>
Mixture		<i>When two or more pure substances are combined, but not bonded together, a mixture is produced.</i>

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, we will discuss mixtures in more detail.

Pure Substance: a form of matter with one set of properties

Element: matter that contains only one type of atom

Compound: the pure substance formed when 2 or more elements combine chemically

Mixture: combinations formed when 2 or more pure substances are combined physically, with each substance retaining its own properties

4.1: Classifying Matter

Chemical Formulas

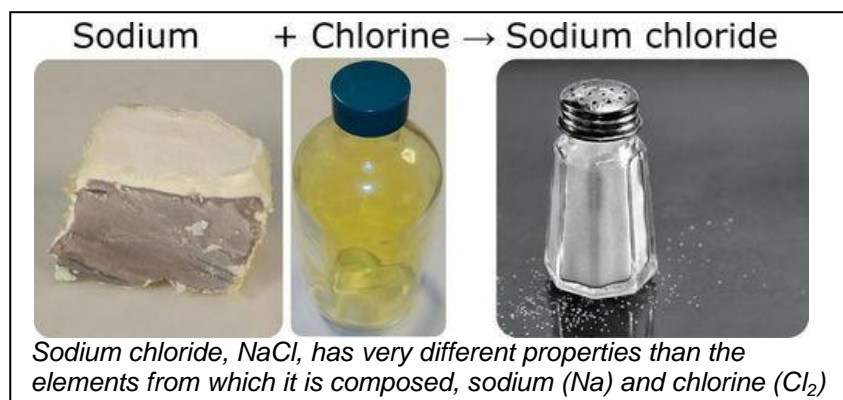
The **formula** for a compound uses the symbols to indicate the type of atoms involved and uses subscripts to indicate the number of each atom in the formula. For example, aluminum combines with oxygen to form the compound aluminum oxide. To form aluminum oxide requires two atoms of aluminum and three 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 $\text{Ca}(\text{OH})_2$, 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.

Each Compound has Unique Properties

It is important to point out what when elements combine to form 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: $\text{C}_{12}\text{H}_{22}\text{O}_{11}$. 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



4.1: Classifying Matter

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.

Carbon Dioxide (CO_2)

Every time you exhale, you release carbon dioxide into the air. It's an odorless, colorless gas. Carbon dioxide contributes to global climate change, but it isn't directly harmful to human health.



Carbon Monoxide (CO)

Carbon monoxide is produced when matter burns. It's a colorless, odorless gas that is very harmful to human health. In fact, it can kill people in minutes. Because you can't see or smell carbon monoxide, it must be detected with an alarm like the one below.



Water (H_2O)

Water is odorless and colorless. We drink it, bathe in it, and use it to wash our clothes. In fact, we can't live without it.



Hydrogen Peroxide (H_2O_2)

Hydrogen peroxide is also odorless and colorless. It's used as an antiseptic to kill germs on cuts. It's also used as bleach to remove color from hair.



Formula: Shows the symbols and number of each element in a compound.

4.2: Types of Compounds

Why Compounds Form

Objectives:

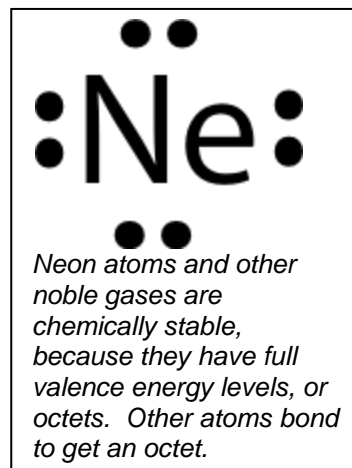
-Describe ways in which atoms can bond to form octets
-Compare the properties and electron bonding structure in ionic, covalent, and metallic compounds.

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 than 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.

Octet Rule: Atoms are most stable with eight valence electrons.

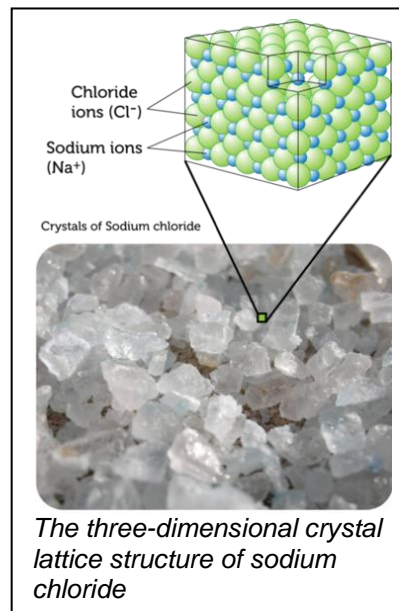
4.2: Types of Compounds

Properties of Ionic Compounds

Ionic compounds are formed when a metal transfers electrons to a nonmetal, forming positive and negatively charged ions which are attracted to each other. 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 form crystals.

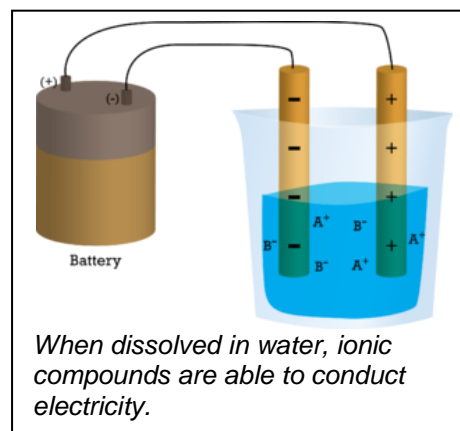
The image shows the solid structure of sodium chloride. Each sodium ion is touching six chloride ions – the four surrounding ones and one above and one below. Each chloride ion is touching six sodium ions in the same way. In ionic compounds negative ions are surrounded by positive ions and vice versa. 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.

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 nonmetallic 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, behaving as separate ions.



The 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.

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. Ionic compounds are **electrolytes**, meaning that they conduct electricity when dissolved in water.

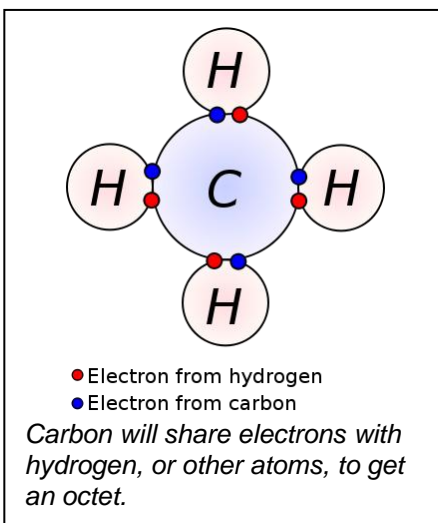


Ionic Bond: The attraction between positive and negative ions; formed when electrons are transferred from a metal to a nonmetal.

Electrolyte: a substance that conducts electricity when dissolved in water.

4.2: Types of Compounds

Properties of Covalent Compounds



In ionic compounds, we learned that atoms are able to achieve an 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 of forming a covalent bond. 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.

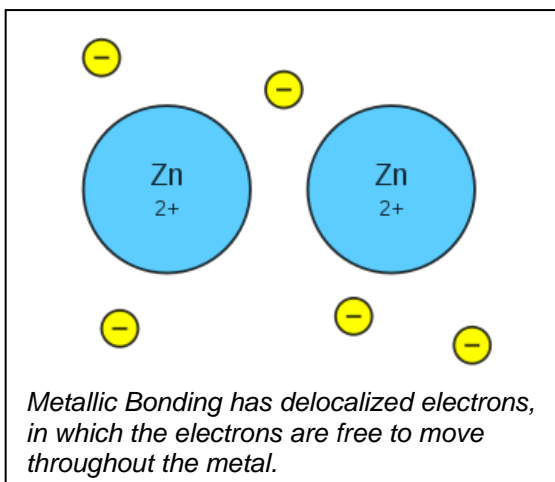
The term *covalent bond* dates from 1939. The prefix *co-* means *jointly* (as in, coworker, cooperate, etc), etc.; “*valent*” is referring to an atom’s valence electrons. Thus, a “**co-valent**

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; they are nonelectrolytes.

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 atom 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.

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).

Covalent Bond: Formed when two atoms share electrons, with both atoms’ nuclei attracted to the electrons in the bond.

Metallic Bond: The type of bond formed when electrons are loosely held on to by the nuclei of nearby atoms.

4.3: Ions

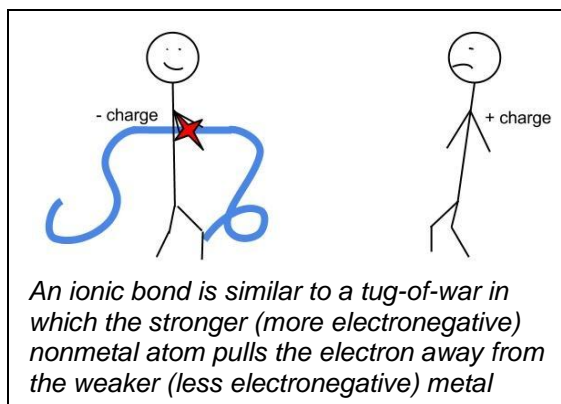
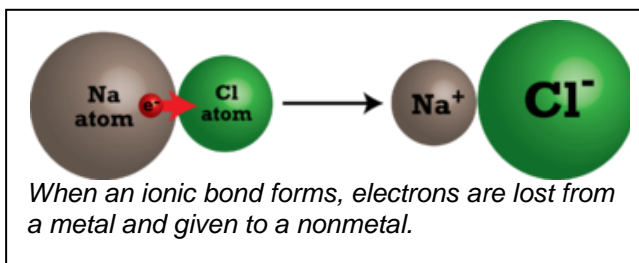
Forming Ions

The process of transferring an electron from a sodium atom to a chlorine atom as shown in 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.

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**. 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**.



Objectives:

- Describe how atoms gain or lose electrons to form ions.
- Predict the charge of main-body ions
- Name main-body, polyatomic, and transition metal ions.

Cation: A positively charged particle formed when electrons are lost
Anion: A negatively charged particle formed when electrons are gained
Electrostatic Attraction: The attraction between particles of opposite charge.

4.3: 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 +1 ions: Li^+ , Na^+ , K^+ , etc. We need to note that while hydrogen is in this same column, it is not considered to be a metal. There are times that hydrogen acts as if it is a metal and forms +1 ions; 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 +1 charge.

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 lose 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 nonmetals will gain two electrons to obtain an 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.

Charges of Ions																													
1 H	+2																5 B	6 C	7 N	8 O	9 F	10 Ne							
3 Li	4 Be															13 Al	14 Si	15 P	16 S	17 Cl	18 Ar								
11 Na	12 Mg	+? use Roman numerals to indicate charge (top #)														31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr								
19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
55 Cs	56 Ba	57 La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn												
87 Fr	88 Ra	89 Ac	104 Rf	105 Db	106 Sg	107 Bh	108 Hs	109 Mt	110 Ds	111 Rg	112 Cn	113 Uut	114 Fl	115 Uup	116 Lv	117	118 Uuo												

change name to end in "-ide"

CC – Tracy Poulsen

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.

4.3: Ions

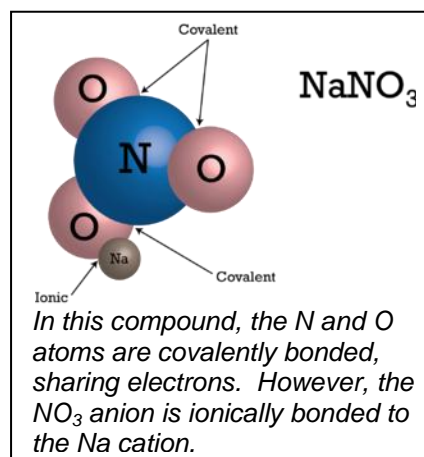
Polyatomic Ions

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.

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.



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 ₂ H ₃ O ₂ ⁻	Oxalate, C ₂ O ₄ ²⁻ Silicate, SiO ₃ ²⁻ Thiosulfate, S ₂ O ₃ ²⁻	
Permanganate, MnO ₄ ⁻	Chromate, CrO ₄ ²⁻ Dichromate, Cr ₂ O ₇ ²⁻	
Cyanide, CN ⁻ Thiocyanate, SCN ⁻		

Polyatomic Ion: Two or more atoms bonded together sharing a charge; polyatomic ions act as one particle

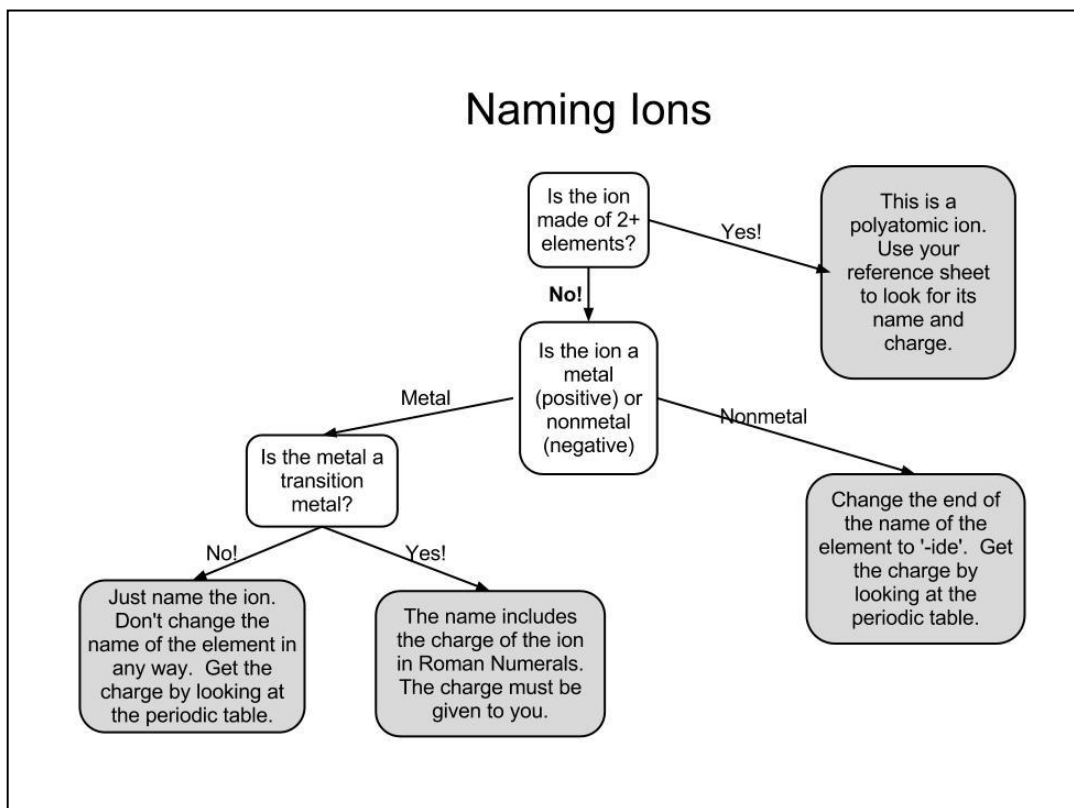
4.3: Ions

Transition Metals Ions

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.

It is important to note that the Roman numerals give the *charge* of the iron ion, not the number of ions in the compound. Ionic compounds in many ways act like separate ions and we write their names and formulas to reflect that.



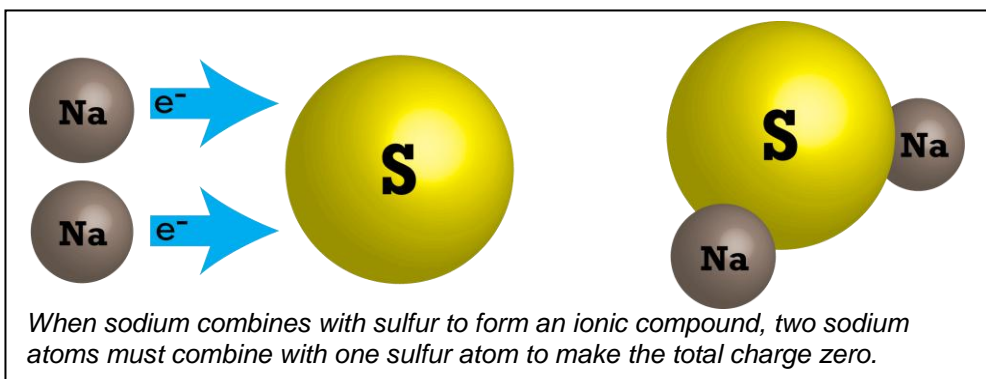
4.4: Ionic Names and Formulas

Writing 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 sodium, which gives off one electron, is to combine with sulfur, which takes two electrons, then two sodium atoms must combine with one sulfur atom. The formula would be Na_2S .

Objectives:

-Write the names and formulas of ionic compounds



This is the main rule to remember when writing formulas of ionic compounds: the total charge will always be zero. The positive ions must cancel out the charge of the negative ions.

To write the formula for an ionic compound:

1) Write the symbol and charge of the cation (first word)

- If the element is in group 1, 2, Al with a consistent charge, you can get the charge using your periodic table.
- If the metal is a transition metal with a variable charge, the charge will be given to you in Roman numerals.

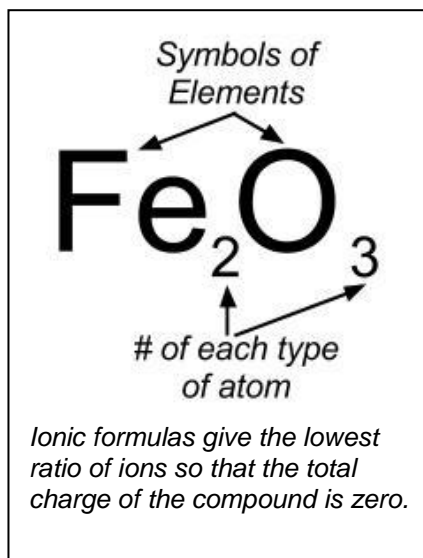
2) Write the symbol and charge of the anion (second word).

- Look at your polyatomic ion chart first. If your anion is a polyatomic ion, write the ion in parentheses.
- 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.



4.4: Ionic Names and Formulas

Examples of Writing Ionic Formulas

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 $1 \cdot (+3)$ and $3 \cdot (-1)$. Giving: $\text{Al}_1^{3+} \text{Cl}_3^{-1}$

The final formula is: AlCl_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 $1 \cdot (+4)$ and $2 \cdot (-2)$. Giving: $\text{Pb}_1^{4+} \text{O}_2^{-2}$

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: $(\text{NO}_3)^-$ (this is a polyatomic ion)

To balance the charges you need $1 \cdot (+2)$ and $2 \cdot (-1)$. Giving: $\text{Ca}_1^{2+} (\text{NO}_3)_2^{-1}$

The final formula is: $\text{Ca}(\text{NO}_3)_2$

In this case you need to keep the parentheses. There are two of the group $(\text{NO}_3)^-$.

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: $(\text{SO}_4)^{2-}$ (this is a polyatomic ion)

To balance the charges you need $1 \cdot (+2)$ and $1 \cdot (-2)$. Giving: $\text{Mg}_1^{2+} (\text{SO}_4)_1^{-2}$

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: $(\text{C}_2\text{H}_3\text{O}_2)^-$ (this is a polyatomic ion)

To balance the charges you need $1 \cdot (+2)$ and $2 \cdot (-1)$. Giving: $\text{Cu}_1^{2+} (\text{C}_2\text{H}_3\text{O}_2)_2^{-1}$

The final formula is: $\text{Cu}(\text{C}_2\text{H}_3\text{O}_2)_2$

In this case you need to keep the parentheses. There are two of the group $(\text{C}_2\text{H}_3\text{O}_2)^-$. Without the parentheses, you are merely changing the number of oxygen atoms.

4.4: Ionic Names and Formulas

Naming Ionic Compounds

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.

Iron(III) Oxide

Name of the cation

Name of the anion

Names of ionic compounds are two words: the name of the cation and the name of the anion. No information about the number of each ion is included.

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: $Cu | Cl_2$

4.4: Ionic Names and Formulas

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: $1 \cdot (+2)$ to cancel out $2 \cdot (-1)$. 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: $1 \cdot (+4)$ to cancel out $2 \cdot (-2)$. 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: $1 \cdot (+3)$ to cancel out $3 \cdot (-1)$. 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.

4.5: Covalent Compound Formation

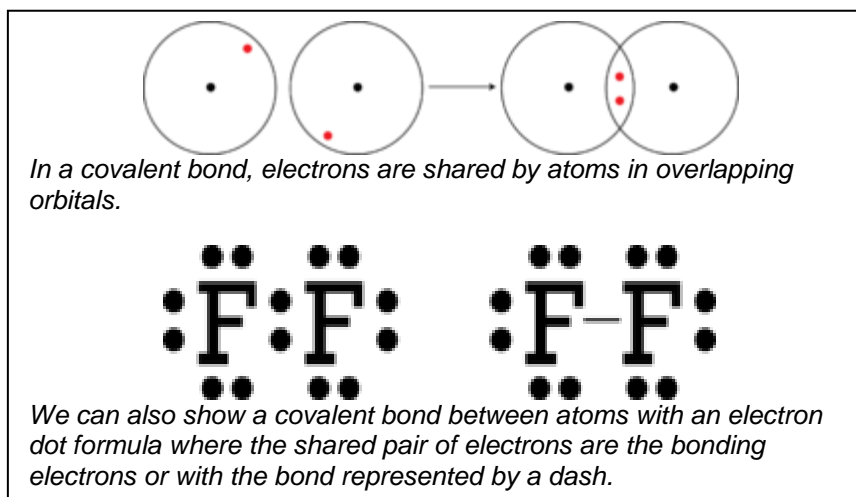
Forming Covalent Bonds

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**.

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.

Objectives:

- Describe how atoms share electrons to form octets
- Indicate which elements form covalent bonds in their elemental state (i.e.: diatomic elements)
- Write the names of covalent compounds



Diatomic Elements

The diagram above shows an example of covalent bonds between two atoms of the same element, in this case two atoms of fluorine. The diagram represents a fluorine molecule. Because there is only one type of element present, it is still called an element – not a compound. Fluorine normally occurs in diatomic (“two-atom”) molecules. Several other elements also occur as diatomic molecules: hydrogen, nitrogen, and all but one of the halogens (fluorine, chlorine, bromine, and iodine).

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

The diatomic hydrogen molecule, H₂, 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.

Diatomic Element: Elements that exist in nature as two-atom molecules.

4.5: Covalent Compound Formation

Naming Covalent Compounds

In naming ionic compounds, there is no need to indicate the number of atoms of each element in a formula because there is only one possible compound that can form from the ions present. The ions combine to cancel the charge of the ions. When aluminum combines with sulfur, the only possible compound is aluminum sulfide, Al_2S_3 .

With covalent compounds, however, we have a very different situation. Nonmetals are combining together by sharing their valence electrons and there is usually more than one way in which they may share electrons to obtain octets. There are six different covalent compounds that can form between nitrogen and oxygen, and for two of them, nitrogen has the same oxidation number. Chemists devised a nomenclature system for covalent compounds that would indicate how many atoms of each element is present in a molecule of the compound.

Greek Prefixes	
Prefix	Number of Atoms
Mono-	1
Di-	2
Tri-	3
Tetra-	4
Penta-	5
Hexa-	6
Hepta-	7
Octa-	8
Nona-	9
Deca-	10

Greek Prefixes

In naming binary covalent compounds, four rules apply:

- The first element in the formula is named first using the normal name of the element.
- The second element is named as if it were an anion. There are no ions in these compounds, but we use the “-ide” ending on the second element.
- Greek prefixes are used for each element to indicate the number of atoms of that element present in the compound. (Remember, prefixes come at the beginning of words)
- The prefix “mono-” is never used for naming the first element. For example, CO is called carbon monoxide, *not* monocarbon monoxide.

Example: Write the correct formula or name for each of the given covalent compounds.

- N_2O
- N_2O_3
- SF_6
- carbon dioxide
- tetraphosphorus decaoxide

Solution:

- dinitrogen monoxide: there are two nitrogen atoms (di) and 1 oxygen atom (mono). Also, remember to change the end of the name to “-ide”
- dinitrogen trioxide: there are two nitrogen atoms (di) and 3 oxygen atom (tri). Also, remember to change the end of the name to “-ide”
- sulfur hexafluoride: there is 1 sulfur atom (would be mono, but we don't put that prefix on the first word) and 6 fluorine atom (hexa). Also, remember to change the end of the name to “-ide”
- CO_2
- P_4O_{10}

4.6: Modeling Covalent Compounds

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.

Objectives:

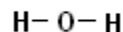
-Draw Lewis Structures for Covalent Compounds
-Indicate the 3d shape of simple covalent compounds.

Example: Draw a Lewis structure for water, H₂O.

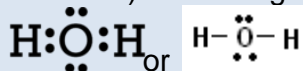
Solution:

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:



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:



4.6: Modeling Covalent Compounds

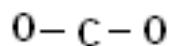
Example: Draw a Lewis structure for CO_2

Solution:

1) *add up all available valence electrons:* $1(4) + 2(6) = 16$

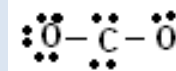
2) *Draw a skeleton.*

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:

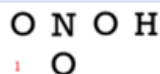


Check:

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:



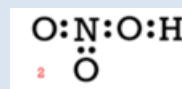
Solution:

1) *add up all available valence electrons:*

$1(1) + 1(5) + 3(6) = 24$

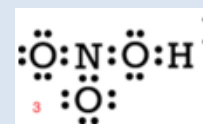
2) *Draw a skeleton.*

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 remaining 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

4.6: Modeling Covalent Compounds

Lewis Structures Do Not Show Shape

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.

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**. 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° .

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**.

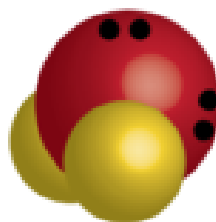
We have a similar problem in the case of a molecule such as water, H_2O . In water, the oxygen atom in the middle is bonded to the two hydrogen atoms with two lone pairs. Once again, we only consider the location of atoms when we discuss shape. When a molecule has a central atom bonded to two other atoms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



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



A central atoms bonded to three other atoms with one lone pairs has a trigonal pyramidal shape.



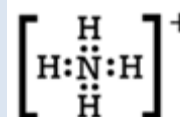
A central atom bonded to two other atoms with two lone of electrons has a bent

4.6: Modeling Covalent Compounds

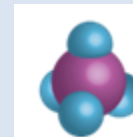
As you can probably imagine, there are different combinations of bonds making different shapes of molecules. Some of the possible shapes are listed in the table. However, it is important to note that some molecules obtain geometries that are not included here.

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

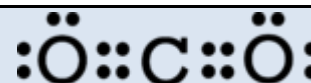
Example: Determine the shape of ammonium, NH_4^+ , given by the following Lewis structure:



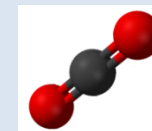
Solution: To answer this question, you need to count the number of atoms around the central atom and the number of unshared pairs. In this example, there are four atoms bonded to the N with zero unshared pairs of electrons. The shape must be tetrahedral.



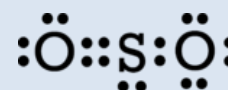
Example: Determine the shape of carbon dioxide, CO_2 , given by the following Lewis structure:



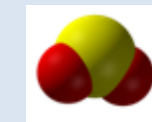
Solution: To answer this question, you need to count the number of atoms around the central atom and the number of unshared pairs. In this example, there are two atoms bonded to the C with zero unshared pairs of electrons. The shape must be linear, according to the table.



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.



VSEPR Theory: Valence Shell Electron Pair Repulsion Theory; The theory used to predict the 3-day shape of molecules.

4.7: Polar and Nonpolar Compounds

Electronegativity and Ion Formation

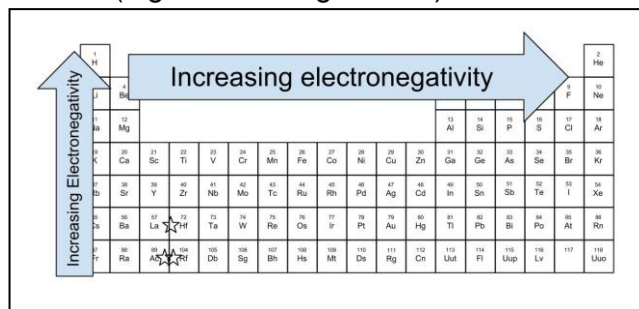
The ability of an atom in a molecule to attract shared electrons is called electronegativity. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither 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.

Objectives:

- Indicate whether a compound is polar or nonpolar from its Lewis Structure
- Identify molecules capable of forming hydrogen bonds

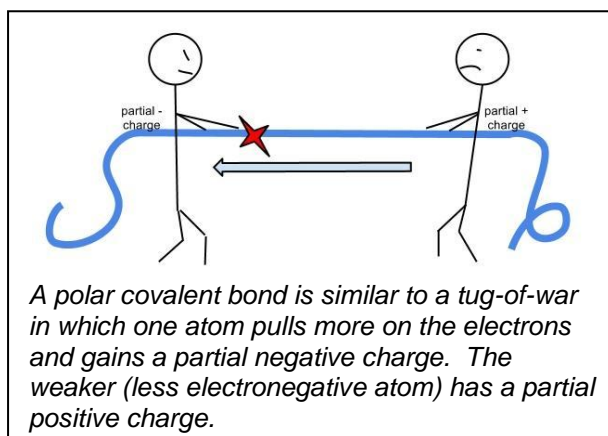
Electronegativity and Covalent Bonding

What happens in covalent compounds? In covalent compounds, there isn't as big difference between the



electronegativities of the atoms that are bonding. If 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. This type of bond is known as a **nonpolar covalent bond**, because the electrons are evenly shared.

What about the molecules whose electronegativities are not the same but the difference is not big enough to form an ionic bond? For these molecules, the electrons remain shared by the two



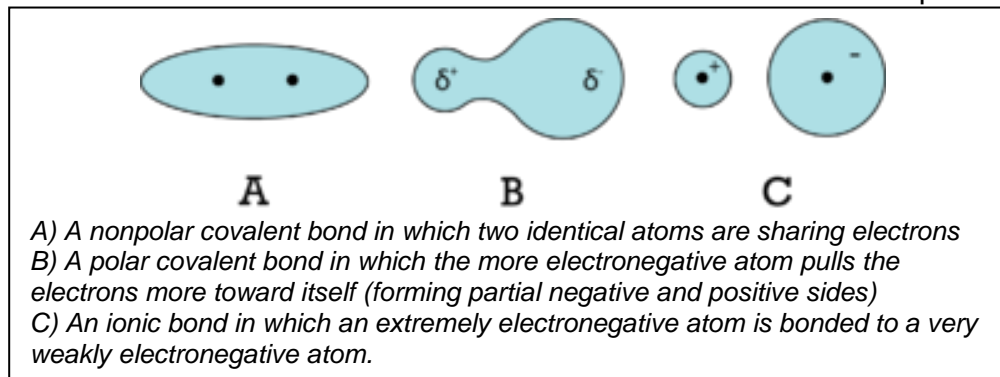
atoms but they are not shared equally. The shared electrons are pulled closer to the more electronegative atom. This results in an uneven distribution of electrons over the molecule and causes slight charges on opposite ends of the molecule. The negative electrons are around the more electronegative atom more of the time creating a partial negative side. The other side has a resulting partial positive charge. These charges are not full +1 and -1 charges, they are fractions of charges. For small fractions of charges, we use the symbols δ^+ and δ^- . These molecules have slight opposite charges on opposite ends of the molecule and said to have a dipole or are called **polar molecules**.

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

4.7: Polar and Nonpolar Compounds

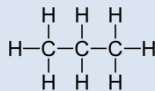
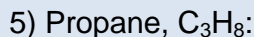
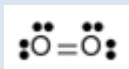
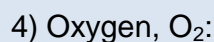
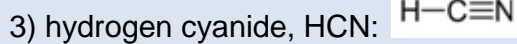
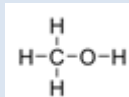
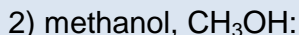
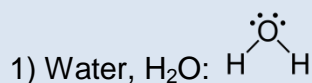
shared with no partial charges forming. This molecule is **nonpolar**. Molecule B is a polar covalent bond formed between atoms whose electronegativities are not the same but whose electronegativity difference is not very big, making this molecule **polar**. Molecule C is an ionic bond formed between atoms whose electronegativity difference is very large.

So far we have talked about individual bonds being polar or nonpolar. What about entire molecules? In order to determine if a molecule is polar



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

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



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 $-\text{OH}$ side is different from the other 3 $-\text{H}$ 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.

Polar Molecule: A molecule with a partial positive and partial negative side caused by the uneven distribution of electrons.

Nonpolar Molecule: A molecule in which electrons are evenly distributed.

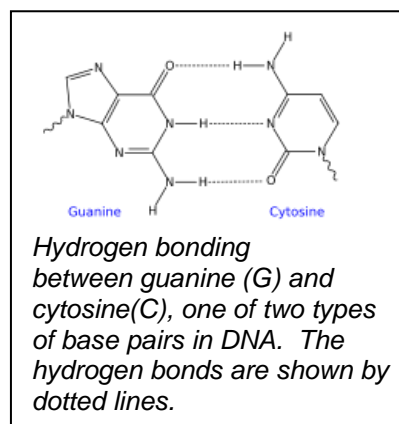
4.7: Polar and Nonpolar Compounds

Hydrogen Bonding

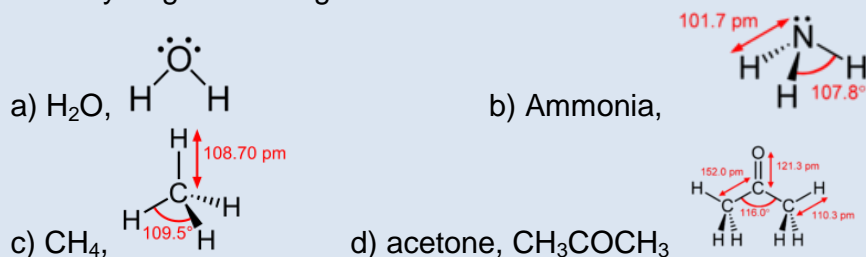
When a hydrogen atom is bonded to a very electronegative atom, including fluorine (F), oxygen (O), or nitrogen (N), a very polar bond is formed. The electronegative atom obtains a negative partial charge and the hydrogen obtains a positive partial charge. These partial charges are similar to what happens 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 partial charges are 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 DNA and proteins

Hydrogen bonding 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.



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



Solution:

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

Hydrogen Bond: A strong attraction between very polar molecules in which hydrogen is bonded to a very electronegative atom, such as fluorine, oxygen, or nitrogen

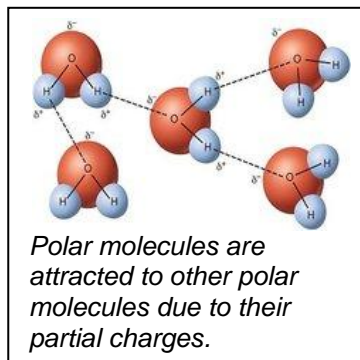
4.8: Properties of Covalent Compounds

Opposite Charges Attract

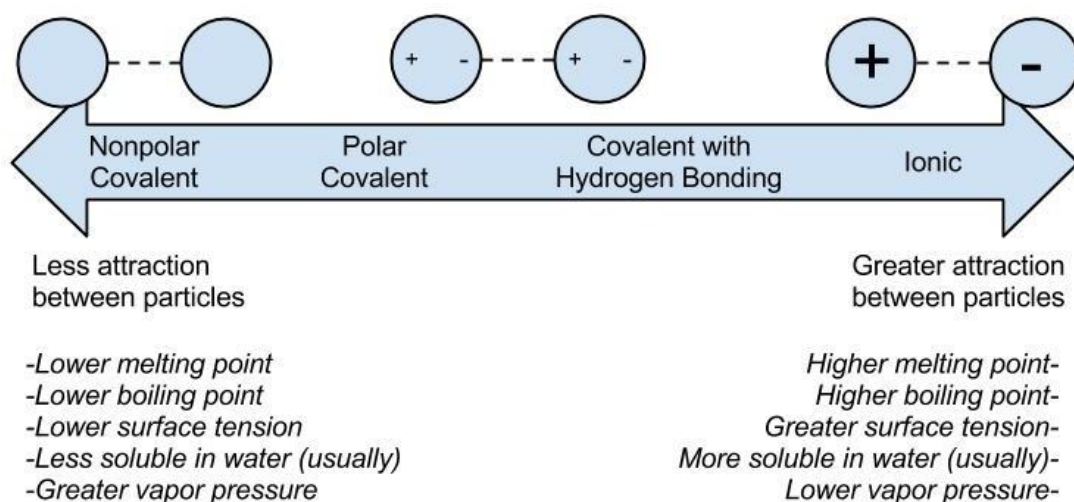
Objectives:

- Compare properties of covalent compounds based on the attractive forces between molecules
- Give examples of water's unique properties due to its hydrogen bonding.

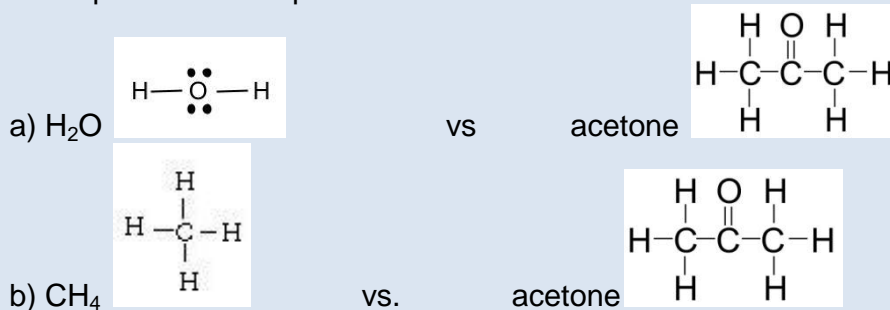
Polar molecules can be attracted to each other due attraction between opposite charges. Polarity and hydrogen bonding underlie a number of physical properties including surface tension, solubility, and melting- and boiling-points. The more attracted molecules are to other molecules, the higher the melting point, boiling point, and surface tension.



The following diagram summarizes many of the properties that are a result of the attraction between molecules or ions.



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.



Solution:

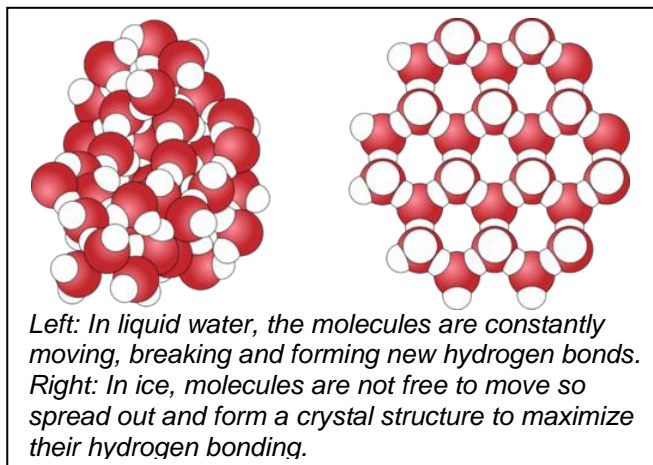
- a) H_2O (polar, hydrogen bonding) vs. acetone (polar, no hydrogen bonding). H_2O will have a higher melting point because compounds with hydrogen bonding tend to have higher melting points than polar compounds.
- b) CH_4 (nonpolar, no hydrogen bonding) vs. acetone (polar, no hydrogen bonding). Acetone will have a higher melting point because polar molecules tend to have higher melting points than nonpolar molecules.

4.8: Properties of Covalent Compounds

Water - a Unique Molecule

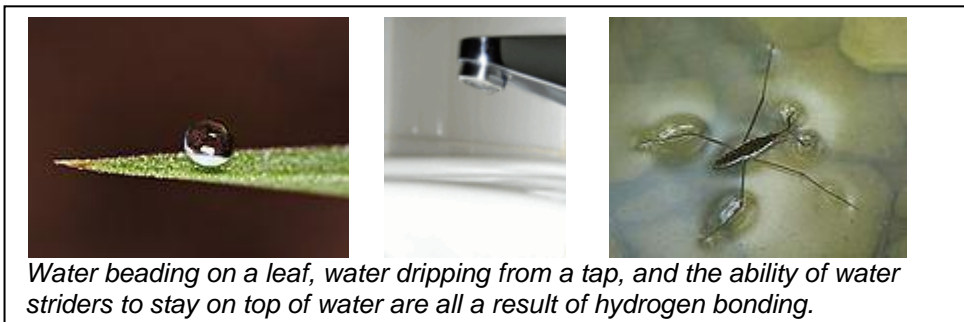
The most important, most common, and perhaps simplest example of a hydrogen bond is found between water molecules. This interaction between neighboring water molecules is responsible for many of the important properties of water.

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. In other words ice floats. This is 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).



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.

Chapter 4 Summary

4.1: Classifying Matter

- A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Elements and compounds are both example of pure substances.
- Mixtures are physical combinations of two or more elements and/or compounds. When substances combine to form a mixture, they retain their original properties.
- 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.
- When elements combine to form compounds, the compound has different properties than the original elements.
- Each compound has unique properties even if they are composed of the same elements.

4.2: Types of Compounds

- The octet rule states that atoms will lose, gain, or share valence electrons to obtain eight valence electrons.
- Ionic compounds form when metals lose electrons to nonmetals.
- 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 form when two nonmetals share valence electrons.
- 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.

4.3: Ions

- 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 charges and some have variable charges. When naming these charge variable ions, their charges are included in Roman numerals.

4.4: Ionic Names and Formulas

- Ionic bonds are formed by transferring electrons from metals to non-metals after which the oppositely charged ions are attracted to each other.
- The total charge of any ionic compound is zero.

Chapter 4 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.
- Ionic compounds are named by naming the cation, then naming the anion.
- Ionic formulas include the number or ratio of each ion in a compound. The name does not.

4.5: Covalent Compound Formation

- Covalent bonds are formed by electrons being shared between two nonmetal atoms.
- Half-filled orbitals of two atoms are overlapped and the valence electrons shared by the atoms.
- Some elements, called diatomic elements, are found in nature bonded in pairs to complete their octets.

4.6: Modeling Covalent Compounds

- Lewis structures show which electrons are bonded and which are unshared in a covalent compound.
- Lewis structures do not show the shape of molecules.
- VSEPR theory is used to predict the 3d shape of molecules and states that valence pairs of electrons repel each other.

4.7: Polar and Nonpolar Compounds

- 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.
- Hydrogen bonding plays roles in many compounds including DNA, proteins, and polymers

4.8: Properties of Covalent Compounds

- Compounds with hydrogen bonding tend to have higher melting points, higher boiling points, and greater surface tension.
- The unique properties of water are a result of hydrogen bonding

Further Reading / Supplemental Links

- Tutorial on bonding:
http://visionlearning.org/library/module_viewer.php?mid=55&l=
- http://www.up.ac.za/academic/chem/mol_geom/mol_geometry.htm
- Vision Learning: Water Properties & Behaviors
http://visionlearning.com/library/module_viewer.php?mid=57&l=&c3=

Chapter 4 Summary

4.1: Classifying Matter Review Questions

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

- 1) Salt, NaCl
 - 2) Gold
 - 3) Sugar, C₆H₁₂O₆
 - 4) Salad dressing
 - 5) Salt water
 - 6) Water
 - 7) Copper
 - 8) Air
 - 9) Milk
- 10) Hemoglobin, the compound in red blood cells that carries oxygen throughout y our body. It is composed of iron, carbon, oxygen, hydrogen and small amounts of other elements. Would you expect hemoglobin to be magnetic, like elemental iron is? Why or why not?
- 11) Table sugar is composed of the elements carbon, hydrogen, and oxygen, and has the chemical formula C₁₂H₂₂O₁₁. Would you expect sugar to have the same properties as elemental carbon? Why or why not?
- 12) Isopropyl alcohol (found in rubbing alcohol) has the formula C₃H₇OH. Acetone (in nail-polish remover) has the formula C₃H₆O. These two compounds have different melting points, boiling points, and have different uses and chemical properties. Why do these two compounds, with similar chemical formulas and composed of the same three element, have different properties?

4.2: Types of Compounds Review Questions

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

#13-21: Given the following chemical formulas, label each compound as ionic, metallic, or covalent.

- | | |
|----------------------|-------------------------------------|
| 15) H ₂ O | 17) NO |
| 16) MgO | 18) Li ₃ PO ₄ |

Chapter 4 Summary

19) Rb_3N

22) CH_4

20) CCl_4

23) NH_3

21) $\text{Ni}_3(\text{PO}_4)_2$

#22-29: Label each of the following properties as a property of an ionic, covalent, or metallic compound:

24) Low melting point

25) Conducts electricity in the solid state

26) Conducts electricity when dissolved in water

27) Forms brittle crystal structures

28) Formed when an element with low electronegativity bonds with an atom with high electronegativity

29) Formed when electrons are transferred between atoms

30) Formed when electrons are shared between atoms

31) Formed when electrons are free to move

32) Steve is given two white substances in the lab to identify. He measures and records their properties in the given data table. Help Steve match the properties of the unknowns (from the table below) to the substance names. Give at least two pieces of evidence to support your claim.

List of Unknown names:

Sodium chloride, NaCl

Sucrose (table sugar, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$)

Unknown Substance	Conducts when dissolved	Malleable	Conducts as a solid	Dissolves in water	Melting Point ($^{\circ}\text{C}$)
1	No	No	No	Yes	164°C
2	Yes	No	No	Yes	$\sim 800^{\circ}\text{C}$

4.3: Ions Review Questions

33) Define the term cation.

34) Define the term anion.

Chapter 4 Summary

#32-43: Predict the charge of each ion. Then give the name each ion would have.

35) Cl

36) Br

37) N

38) O

39) Ca

40) F

41) Mg

42) Li

43) I

44) Na

45) K

46) Al

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

#45-53: Name the following ions.

48) Cu^{2+}

49) Co^{2+}

50) Co^{3+}

51) Cu^{+}

52) Ni^{2+}

53) Cr^{3+}

54) Fe^{2+}

55) Fe^{3+}

56) Mn^{2+}

57) What are polyatomic ions?

4.4: Ionic Names and Formulas Review Questions

#55-72: Fill in the chart by writing formulas for the compounds that might form between the ions in the columns and rows.

	Na^{+}	Ca^{2+}	Fe^{3+}
NO_3^{-}	58)	59)	60)
SO_4^{2-}	61)	62)	63)
Cl^{-}	64)	65)	66)
PO_4^{3-}	67)	68)	69)
OH^{-}	70)	71)	72)

Chapter 4 Summary

CO_3^{2-}	73)	74)	75)
--------------------	-----	-----	-----

#73-90: Write the formulas from the names of the following compounds.

76) Magnesium sulfide

84) Magnesium oxide

77) Lead(II) nitrate

85) Calcium oxide

78) Calcium hydroxide

86) Copper(I) bromide

79) Potassium carbonate

87) Aluminum sulfide

80) Aluminum bromide

88) Hydrogen carbonate

81) Iron (III) nitrate

89) Potassium permanganate

82) Iron (II) chloride

90) Copper(I) dichromate

83) Copper(II) nitrate

4.5: Covalent Compound Formation Review Questions

Name the following compounds.

91) KCl

97) $\text{NaC}_2\text{H}_3\text{O}_2$

92) MgO

98) CuO

93) CuSO_4

99) FeCl_2

94) CoBr_2

100) LiCl

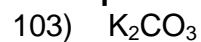
95) MgF_2

101) MgBr_2

96) Ni(OH)_2

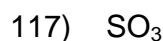
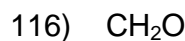
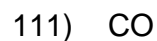
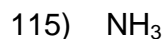
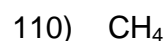
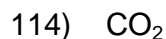
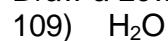
102) Cu_2O

Chapter 4 Summary



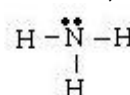
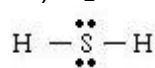
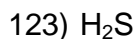
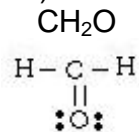
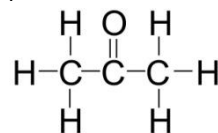
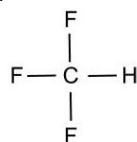
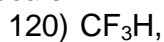
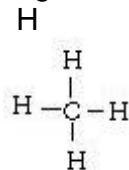
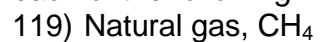
4.6: Modeling Covalent Compounds Review Questions

Draw a Lewis structure for each of the following compounds.



118) For each Lewis Structure you drew for #112-120, indicate the 3d shape around the central atom.

Given the Lewis structure, predict the 3d shape around the central atom in each of the following molecule:



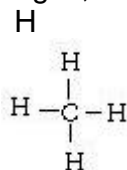
Chapter 4 Summary

4.7: Polar and Nonpolar Compounds

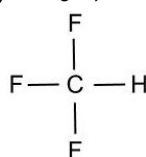
- 125) Predict which of the following bonds will be more polar and explain why; P-Cl or S-Cl.
- 126) What does it mean for a molecule to be “polar”?
- 127) Which three elements, when bonded with hydrogen, are capable of forming hydrogen bonds?
- 128) Molecules that are polar exhibit dipole-dipole interaction. What’s the difference between dipole-dipole interactions and hydrogen bonding? Which interaction is stronger?
- 129) 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.

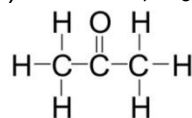
130) Natural gas, CH₄



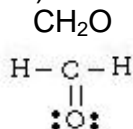
131) CF₃H,



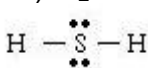
132) Acetone, C₃H₆O



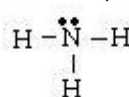
133) Formaldehyde,



134) H₂S



135) Ammonia, NH₃



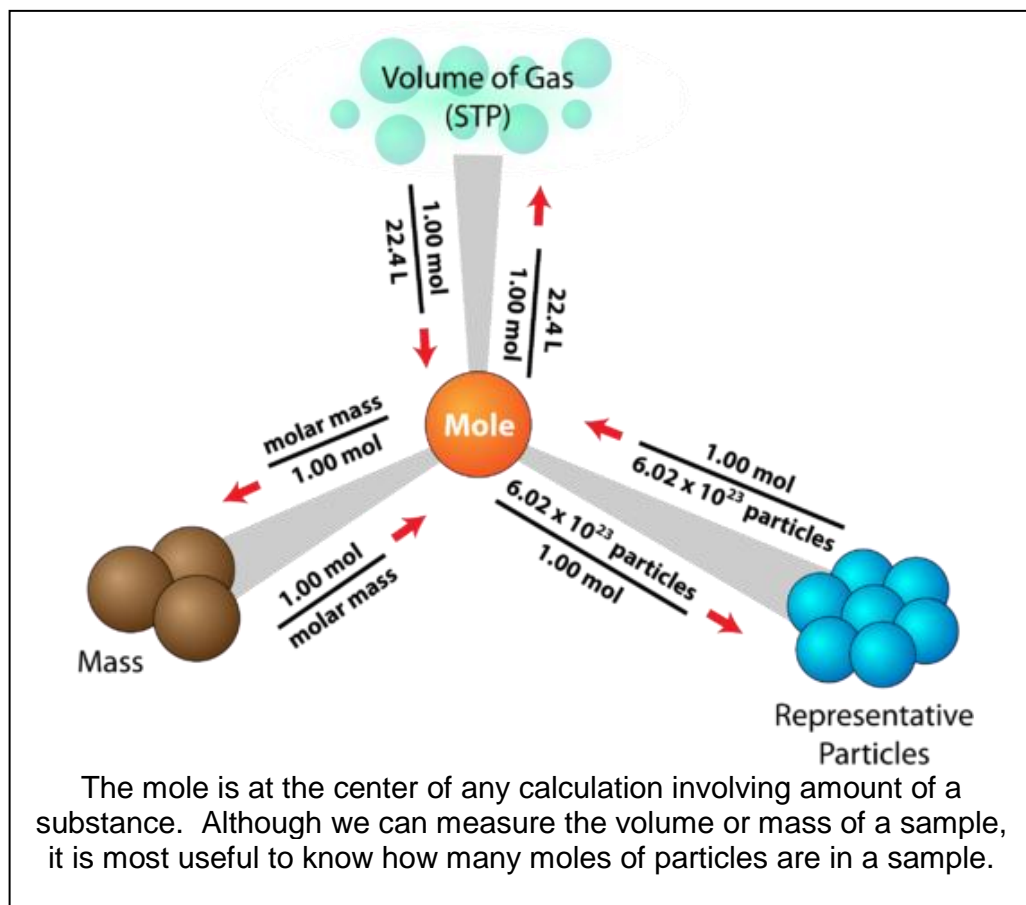
Use the Lewis structure given in the previous problem set to indicate which of the compounds in each pair has the given property.

- 136) higher melting point: ammonia or methane
- 137) higher boiling point: water or CH₃Cl
- 138) more soluble in water: ammonia or CHCl₃
- 139) higher melting point: SiF₄ or ammonia

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Chapter 5: Math in Chemistry

Problem Solving and the Mole



5.1: Measurement Systems

Units of Measurement

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. The requirements of science in the 1600s, 1700s, and 1800s necessitated even more accurate, reproducible measurements.

Why Scientists Use 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 to explain why the metric system is so much easier to manipulate. If you wanted to calculate how many inches were in two miles, you need to know how many inches are in a foot and many feet are in a mile. What happens if you never memorized these facts? 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.

The Metric System is Based on 10's

The metric system uses a number of prefixes along with the base units. Each base unit can be combined with different prefixes to define smaller and larger quantities. For example, there are 100 centigrams in 1 gram and 100 centiliters in 1 liter. It doesn't matter the unit, the prefix always indicates the relationship between them. Each unit uses the same prefixes, so we don't need to memorize relationships. Some common metric equivalencies are found in the table below.

Prefix	Base unit
100 centi(<u>base</u>)	1 (<u>base</u>)
1 kilo(<u>base</u>)	1000 (<u>base</u>)
1000 milli(<u>base</u>)	1 (<u>base</u>)

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

Objectives:

- Give advantages of using the SI (metric) system of units
- Indicate the most common units used for mass, volume, temperature, and quantity in chemistry

5.1: Measurement Systems

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 official SI unit for mass is the kilogram, but for the quantities we use in chemistry, the unit g is usually better.

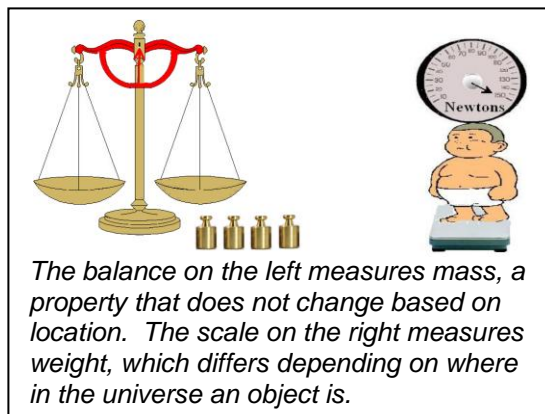
The **weight** of an object is the force of attraction between the object and the earth (or whatever large body it is resting on). We call this force of attraction the force of gravity. Since the force of gravity is not the same at every point on the earth's surface, the weight of an object depends it is located. For example, a man who weighs 180 pounds on Earth 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.

Length

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 SI unit of length is the **meter**.

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



5.1: Measurement Systems

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) or milliliter (mL)**.

$$1\text{ L} = 1000\text{ mL} = 1000\text{ cm}^3$$

Temperature

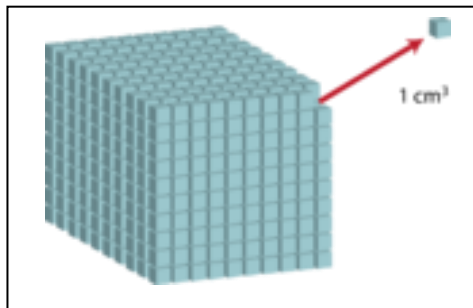
When used in a scientific context, the words heat and temperature do NOT mean the same thing. Temperature represents the average kinetic energy of the particles that make up a material. If you heat up a material, its temperature increases as the particles move faster.

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 relationship between the Celsius scale and the Kelvin scale is $K = ^\circ\text{C} + 273$. 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.



SI Units: the internationally accepted system of measurements; also called the metric system

5.2: Scientific Notation Using Scientific Notation?

Objectives:

-Convert numbers between standard and scientific notation

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 kg; and the mass of an electron is 0.000000000000000000000000000009 kg. It is very inconvenient to write such numbers and even more inconvenient to attempt to carry out mathematical operations with them. 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 9×10^6 . In a similar manner, 0.00000004 can be written as $4 \times \frac{1}{10^8}$ or 4×10^{-8} .

This is called the "coefficient." → **6.5** x 10^{**4**} ← This is called the "exponent."

As you can see from the examples, to convert a number from decimal form to scientific notation, you count the spaces that you need to move the decimal. If the number is large, the exponent will be positive. If the number is less than 1, the exponent will be negative.

Examples of Scientific Notation	
Decimal Notation	Scientific Notation
95,672	9.5672×10^4
8,340	8.34×10^3
100	1×10^2
7.21	7.21×10^0
0.014	1.4×10^{-2}
0.000000008	8.0×10^{-9}

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:

$$1.20EE24 \div 6.02EE23 \quad \text{OR} \quad 1.20EXP24 \div 6.02EXP23$$

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.

5.3: Significant Figures

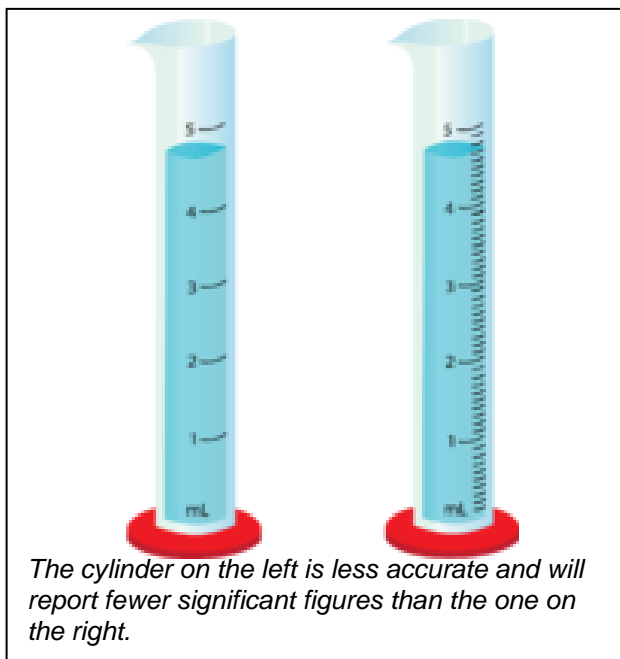
Significant Figures Show Accuracy

The numbers you use in math class are considered to be exact numbers. When you are given the number 2 in a math problem, it does not mean 1.999 rounded up to 2, nor does it mean 2.00001 rounded down to 2. In math class, the number 2 means exactly 2.000000... with an infinite number of zeros – a perfect 2! Such numbers are produced only by definition, *not* by measurement. We can define 1 foot to contain exactly 12 inches with both numbers being perfect numbers, but we cannot measure an object to be exactly 12 inches long. In the case of measurements, we can only read our measuring instruments to a limited number of subdivisions. Measurements do not produce perfect numbers; the only perfect numbers in science are defined numbers, such as conversion factors.

It is very important to recognize and report the limitations of a measurement along with the magnitude and unit of the measurement. Many times, the measurements made in an experiment are analyzed for regularities. If the numbers reported show the limits of the measurements, the regularity, or lack thereof, becomes visible.

Equipment Determines Significant Figures

The choice of measuring instrument determines the unit of measure and the number of significant figures in the measurement. Consider the two graduated cylinders. Both cylinders are marked to measure milliliters, but the cylinder on the left only shows graduations for whole milliliters. In comparison, the cylinder on the right has calibrations for tenths of milliliters. The measurer reads the volume from the calibrations and estimates one place beyond the calibrations. For the cylinder on the left, a reasonable reading is 4.5 mL. For the cylinder on the right, the measurer estimates one place beyond the graduations and obtains a reasonable reading of 4.65 mL. The choice of the measuring instrument determines both the units and the number of significant figures. If you were mixing up some hot chocolate at home, the cylinder on the left would be adequate. If you were measuring out a chemical solution for a very delicate reaction in the lab, however, you would need the cylinder on the right.



Objectives:
-Describe the need for significant figures in science
-Report a measurement to the correct number of significant figures
-When performing calculations, round the answer to the correct number of significant figures.

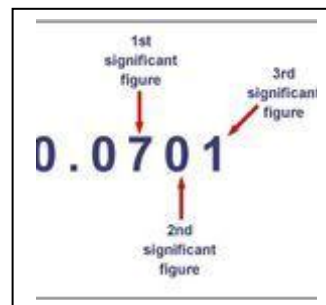
Significant Figures: Digits of a number that are used to express the accuracy of a measurement or calculation

5.3: Significant Figures

Rules for Counting Number of Significant Figures

In our system of writing measurements to show significant figures, we must distinguish between measured zeros and place-holding zeros. Here are the rules for determining the number of significant figures reported in a measurement.

- All non-zero digits (1-9) are significant.
- All zeros between non-zero digits are significant.
- All beginning zeros (leading zeros) are *not* significant.
- Ending zeros are significant if the decimal point is actually written but *not* significant if the decimal point is an understood decimal (the decimal point is not written in).



A useful tool to keep in mind is using scientific notation to report numbers. If a number is written in scientific notation, all of the digits written are significant. This is one more reason why scientific notation is frequently used among scientists. This is especially useful when trying to designate which zeros are significant in a measurement and which are not.

Example: Determine the number of significant figures in each of the following:

- a) 22.437 b) 7004 c) 10.032 d) 0.002
e) 0.003003 f) 4700 g) 450. h) 23.500

Solution: For each solution, the digits that count as significant are underlined. The total number of significant figures is then reported.

- a) 22.437 (5) All non-zero digits are significant.
b) 7004 (4) All zeros between non-zero digits are significant.
c) 10.032 (5) All zeros between non-zero digits are significant.
d) 0.002 (1) All beginning zeros are *not* significant.
e) 0.003003 (4) All zeros between non-zero digits are significant. All beginning zeros are *not* significant.
f) 4700 (2) Ending zeros are significant if the decimal point is actually written in but *not* significant if the decimal point is an understood decimal.
g) 450. (3) Ending zeros are significant if the decimal point is actually written in but *not* significant if the decimal point is an understood decimal.
h) 23.500 (5) Ending zeros are significant if the decimal point is actually written in but *not* significant if the decimal point is an understood decimal.

Example: Round each of the following numbers to 3 significant figures.

- a) 30890 b) 30
c) 0.025399 d) 0.025

Solution:

- a) 30900 b) 30.0
c) 0.0254 d) 0.0250

5.3: Significant Figures

Significant Figures in Calculations

In addition to using significant figures to report measurements, we also use them to report the results of calculations made with measurements. The results of mathematical operations on measurements must indicate the number of significant figures in the original measurements. Most of the calculations we will be doing will be addition/subtraction or multiplication/division. These two basic functions have different rules for determining the number of significant figures after performing a mathematical operation. Most of the errors that occur in this area result from using the wrong rule, so always double check that you are using the correct rule for the mathematical operation involved.

Addition and Subtraction

The answer to an addition or subtraction operation must not have any digits further to the right than the shortest addend. In other words, the answer should have as many decimal places as the addend with the smallest number of decimal places. The following steps will help you keep track of significant figures when adding and subtracting:

- Write the numbers you are adding/subtracted over each other with the decimal point lined up
- Your final answer can have a significant figure in the last column which all of the addends have significant figures

$$\begin{array}{r} 24.300 \text{ mL} \\ - 10.24 \text{ mL} \\ \hline 14.06 \text{ mL} \end{array}$$

When adding and subtracting numbers, round your answer according to place value

Example: What is $13.3843 + 1.012 + 3.22$ rounded to the correct number of significant figures?

Solution: You must first line up the decimal points and look for the last column in which every number has a digit. In this case, that is the hundredths column. We must round to the hundredths column.

$$\begin{array}{r} 13.3843 \text{ cm} \\ 1.012 \text{ cm} \\ + 3.22 \text{ cm} \\ \hline 17.6163 \text{ cm} \end{array}$$

Final Answer: 17.62 cm

Sometimes you have a column with no number in that column. In math classes, you were taught that these blank spaces can be filled in with zeros and the answer would be 17.6163cm. In the sciences, however, these blank spaces are unknown numbers, *not* zeros. Since they are unknown numbers, you cannot substitute any numbers into the blank spaces. As a result, you cannot know the sum of adding (or subtracting) any column of numbers that contain an unknown number. When you add the columns of numbers in the example above, you can only be certain of the sums for the columns with known numbers in each space in the column. In science, the process is to add the numbers in the normal mathematical process and then round off all columns that contain an unknown number (a blank space). Therefore, the correct answer for the example above is 17.62 cm and has only four significant figures.

5.3: Significant Figures

Example: What is the sum of 12m and 0.00045 m?

Solution: You must first line up the decimal points and look for the last column in which every number has a digit. In this case, that is the ones column. We must round to the ones place.

$$\begin{array}{r} 12 \text{ m} \\ + 0.00045 \text{ m} \\ \hline 12.00045 \text{ m} \\ 12 \text{ m} \end{array}$$

In this case, the addend 12 has no digits beyond the decimal. Therefore, all columns past the decimal point must be rounded off in the final answer. We get the seemingly odd result that the answer is still 12, even after adding a number to 12. This is a common occurrence in science and is absolutely correct.

Multiplication and Division

The answer for a multiplication or division operation must have the same number of significant figures as the factor with the least number of significant figures. Follow the given steps:

- Count the number of significant figures reported in each of your multiplicands.
- Perform your calculation. You must round your answer to the smallest number of digits you counted in the previous step.

$$\begin{array}{ccccccc} \text{3 sigfigs} & \text{4 sigfigs} & \infty \text{ sigfigs} & \text{4 sigfigs} & & \text{3 sigfigs} & \\ 1.50\text{g Ag} & \frac{1 \text{ mol Ag}}{107.9\text{g Ag}} & \frac{1 \text{ mol Cu}}{2 \text{ mol Ag}} & \frac{63.55\text{g Cu}}{1 \text{ mol Cu}} & = & 1.03069\dots\text{g Ag} & \\ & & & & & \mathbf{1.03\text{g Ag}} & \end{array}$$

When multiplying and dividing numbers, round the final answer to the same number of significant figures as the multiplicand with the fewest significant figures.

Example: What is the product of 3.556 cm and 2.4 cm?

Solution: For multiplication and division, we must count the number of significant figures in each multiplicand or dividend. 3.556 cm has 4 significant figures. 2.4 only has two significant figures. The answer will be rounded to two significant figures:

$$3.556 \text{ cm} \cdot 2.4 \text{ cm} = 8.5344 \text{ cm}^2$$

Rounded to two significant figures: 8.5 cm^2

Example: What is the reported value of the calculation: $20.0 \text{ cm} \cdot 5.0000 \text{ cm}$?

Solution: For multiplication and division, we must count the number of significant figures in each multiplicand or dividend. 20.0 cm has 3 significant figures. 5.0000 has five significant figures. The answer will be rounded to three significant figures:

$$20.0 \text{ cm} \cdot 5.0000 \text{ cm} = 100 \text{ cm}^2$$

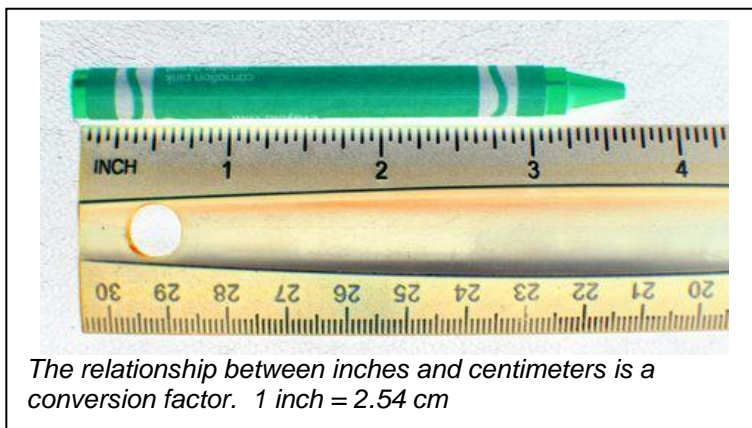
Rounded to three significant figures: $100. \text{ cm}^2$ or $1.00 \times 10^2 \text{ cm}^2$

*The decimal must be written in to show that the two ending zeros are significant. If the decimal is omitted (left as an understood decimal), the two zeros will not be significant and the answer will be wrong.

5.4: Factor-Label Method

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



The relationship between inches and centimeters is a conversion factor. 1 inch = 2.54 cm

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.

Objectives:

- Use a table of conversion factors to solve problems using the factor-label method
- Calculate and use density in factor-label problems

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 within the metric system, as covered earlier in this chapter.
- Common knowledge ratios (such as 60 seconds = 1 minute)

Conversion Factor: A ratio used to convert one unit of measurement into a different unit of measurement.

5.4: Factor-Label Method

Density is a Ratio

Density is an important physical property of matter. It reflects how closely packed the particles of matter are. When particles are packed together more tightly, matter has greater density. Differences in density of matter explain many phenomena, not just why helium balloons rise. For example, differences in density of cool and warm ocean water explain why currents such as the Gulf Stream flow through the oceans.

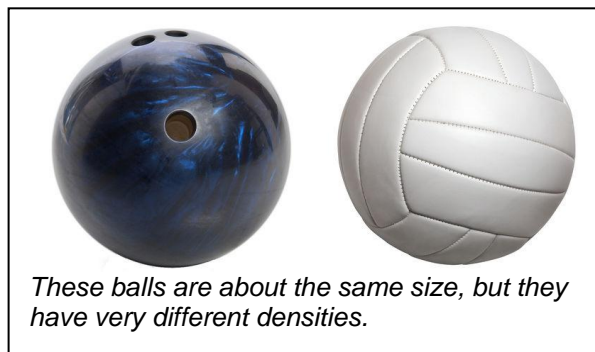
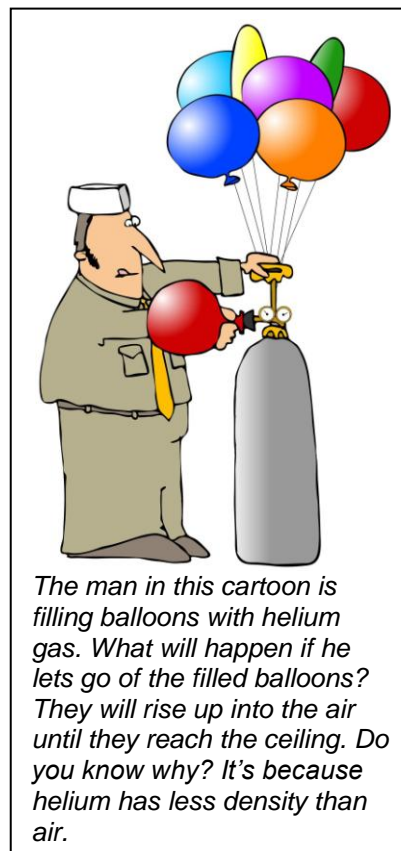
To better understand density, think about a bowling ball and volleyball. Imagine lifting each ball. The two balls are about the same size, but the bowling ball feels much heavier than the volleyball. That's because the bowling ball is made of solid plastic, which contains a lot of tightly packed particles of matter. The volleyball, in contrast, is full of air, which contains fewer, more widely spaced particles of matter. In other words, the matter inside the bowling ball is denser than the matter inside the volleyball.

Calculating Density

The density of matter is a ratio showing the amount of mass in a given volume of space. Density of matter can be calculated with this formula:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{m}{V}$$

Density typically has units of g/mL or g/cm³. We can use this equation to calculate the density or we can use it as a ratio in many problems using the factor-label method.



Example: What is the density of a liquid that has a volume of 30.mL and a mass of 300.g?

Solution:

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$
$$D = \frac{300. \text{g}}{30. \text{mL}} = 10 \text{ g/mL}$$

5.4: Factor-Label Method

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 that will cancel units you don't want and produce units you do want. The conversion factors are created from the equivalency relationships between the units 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 relationship $1000 \text{ millimeters (mm)} = 1 \text{ meter (m)}$.

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

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 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:

Given: 1.53 g

Find: ? cg

Ratios: $100\text{cg}=1\text{g}$ (given in the second table),

The problem is set-up to cancel out the unit g and give us units of cg. Once the units cancel, multiply the top numbers and divide by the bottom numbers. We get:

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

5.4: Factor-Label Method

Example: Convert 1000 inches to feet.

Solution:

Given: 1000 in

Find: ? ft

Ratios: 12 inches = 1 ft. Set the problem up to cancel inches and leave feet. Once the units cancel, multiply the top numbers and divide by the bottom numbers. We get:

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

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 = 1 m, and 1000 m = 1 km. We will set up a series of conversion factors so that each conversion factor produces the next unit in the sequence.

Example: Convert 12 miles to kilometers.

Solution:

Given: 12 miles

Find: ? km

Ratios: 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. Once the units cancel, multiply the top numbers and divide by the bottom numbers. We get:

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

Example: Aluminum has a density of 2.70 g/mL. What is the mass of a cylinder of aluminum with a volume of 7.5 mL?

Solution:

Given: 7.5 mL

Find: ? g (mass)

Ratios: This problem gave us a ratio within the problem that we can use as a conversion factor: 2.70 g = 1 mL of aluminum.

$$7.5 \text{ mL} \cdot \frac{2.70 \text{ g}}{1 \text{ mL}} = 20. \text{ g}$$

Conversion factors units that are squared or cubed can also be produced by this method.

Example: Convert 1500 cm² to m².

Solution:

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

OR

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

Factor-Label Method: A problem-solving method involving multiplying by ratios and cancelling units

5.5: The Mole

Units of Counting

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.” 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.

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 had no hand in determining this number, rather it was named in honor of Avogadro.

Objectives:

- Give the SI unit for counting
- Convert between molecules (or atoms) and moles of a substance
- Find the molar mass of a substance
- Convert between mass and moles of a substance



A mole is not only a small rodent, but is also the SI unit of counting



Avogadro's Number is named after the Italian chemist Amedeo Avogadro

Mole: The SI unit of counting, equal to 6.02×10^{23} objects or particles

5.5: The Mole

The Relationship between Molecules and Moles

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:

Given: 1×10^9 molecules H_2O

Find: mol H_2O

Ratios: Avogadro's number converts between moles and number of particles. Set it up to cancel out the units.

$$1 \times 10^9 \text{ molecules } H_2O \cdot \frac{1 \text{ mol } H_2O}{6.02 \times 10^{23} \text{ molecules } H_2O} = 1.7 \times 10^{-15} \text{ mol } H_2O$$

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 $C_6H_{12}O_6$?

Solution:

Given: 0.00100 mol $C_6H_{12}O_6$

Find: molecules $C_6H_{12}O_6$

Ratios: Avogadro's number converts between moles and number of particles. Set it up to cancel out the units.

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

The Relationship between Mass and Moles

Scientists chose the number 6.02×10^{23} to represent the number of items in a mole because of the relationship between this number and the mass (grams) of an element. 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 $H_2O = 2(1.0084 \text{ g/mol}) + 15.9994 \text{ g/mol} = 18.0162 \text{ g/mol}$. This means that 1 mole of water has a mass of 18.0162 grams.

Element name	Aluminum
Atomic # (# of protons)	13
Element symbol	Al
Atomic mass (average mass of atoms)	26.98

The atomic mass of an element is the decimal number found on the periodic table

Example: Find the molar mass of each of the following:

a) S

b) CO_2

c) F_2

d) H_2SO_4

e) $Al_2(SO_4)_3$

Solution: You will need a periodic table to solve these problems. Look for each element's mass.

a) Look for sulfur on the periodic table. Its molar mass is **32.065 g/mol**. That means that one mole of sulfur has a mass of 32.065 grams.

b) This compound contains one carbon atom and two oxygen atoms. To find the

5.5: The Mole

molar mass of CO_2 , We get: $(12.01) + 2(16.00) = \mathbf{44.01 \text{ g/mol}}$. 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: $2(19.00) = \mathbf{38.00 \text{ g/mol}}$
- 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: $2(1.008) + 32.065 + 4(16.00) = \mathbf{100.097 \text{ g/mol}}$
- e) This compound contains two aluminum atoms, three sulfur atoms, and twelve oxygen atoms. To find the molar mass of $\text{Al}_2(\text{SO}_4)_3$, we need to add the mass of all of these atoms. We get: $2(26.98) + 3(32.065) + 12(16.00) = \mathbf{342.155 \text{ g/mol}}$

Converting Between Mass and Moles

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.

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:

Given: 108 g H_2O

Find: mol H_2O

Ratios: The relationship between moles and mass is the molar mass of water, H_2O .

$$108 \text{ g } \text{H}_2\text{O} \cdot \frac{1 \text{ mol } \text{H}_2\text{O}}{18.02 \text{ g } \text{H}_2\text{O}} = \mathbf{5.99 \text{ mol } \text{H}_2\text{O}}$$

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:

Given: 7.50 mol CaO

Find: g CaO

Ratios: The relationship between moles and mass is the molar mass of calcium oxide, CaO .

$$7.50 \text{ mol } \text{CaO} \cdot \frac{56.0 \text{ g } \text{CaO}}{1 \text{ mol } \text{CaO}} = \mathbf{420. \text{ g } \text{CaO}}$$

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 later. Being able to use these ratios is a very important skill that we will use frequently, so make sure you master it.

Molar mass: the mass, in grams, of one mole of a substance

Chapter 5 Summary

5.1: Measurement Systems

- 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 basic unit of mass in the International System of Units is the kilogram.
- 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 $\text{K} = ^{\circ}\text{C} + 273$.

5.2: Scientific Notation

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

5.3: Significant Figures

- All measurements have some amount of error. Significant figures are used in science to represent the accuracy of measurements.
- Some ratios used in science are exact, and have infinite significant figures. These include defined quantities, such as $1000 \text{ mL} = 1\text{L}$.
- When adding and subtracting numbers in science, the answer is rounded to the same place value as the least significant addend.
- When multiplying and dividing numbers in science, the answer is rounded to the same number of significant figures as the multiplicand with the fewer number of significant figures.

5.4 The Factor-Label Method

- 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 that will cancel units you don't want and produce units you do want.

5.5: The Mole

- There are 6.02×10^{23} particles in 1 mole. This number is called Avogadro's number.
- The molar mass is the mass in grams one mole of a substance.
- 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.

Chapter 5 Summary

Further Reading / Supplemental Links

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

5.1: Measurement Systems 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) Would it be comfortable to swim in a swimming pool whose water temperature was 275 K? Why or why not?

5.2: Scientific Notation Review Questions

- 10) When is it useful to use scientific notation?

Write the following numbers in scientific notation.

- 11) 0.0000479
- 12) 4260
- 13) 251,000,000
- 14) 0.00206

Chapter 5 Summary

Write each of the following numbers in standard notation.

15) 2.3×10^4

16) 9.156×10^{-4}

17) 7.2×10^{-3}

18) 8.255×10^6

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

19) $2.0 \times 10^3 \cdot 3.0 \times 10^4$

20) $2.0 \times 10^3 / 3.0 \times 10^4$

21) $4.2 \times 10^{-4} / 3.0 \times 10^{-2}$

22) $7.3 \times 10^{-7} \cdot 8.0 \times 10^{-3}$

5.3: Significant Figures Review Questions

Indicate the number of significant figures in each of the following reported measurements.

23) 65.2 g

26) 5.70500 mg

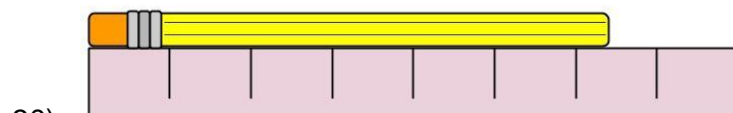
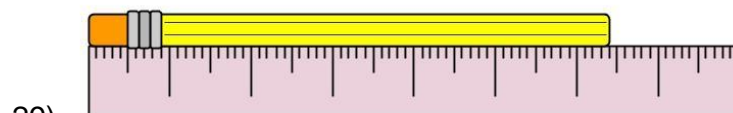
24) 14,000 people

27) 300. K

25) 0.0053 ns

28) 102.0

Use the image to report the length of the pencil to the correct number of significant figures. The big dashes mark centimeters in each image.



Write each of the following showing three significant figures.

31) 70238g

33) 12 ft

32) 0.10392 m

34) 0.00237 s

Perform each calculation to the correct number of significant figures.

35) $1.0234 + 1.1 + 0.0056$

36) $584.65 + 1100$

37) $900 + 500$

38) $3.84 \times 21.69 / 2.8 / 1.62$

Chapter 5 Summary

39) 1.20 m x 0.87 m

5.4: The Factor-Label Method 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.

40) What is the diameter of a 9" cake pan in centimeters?

41) 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?

42) 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?

43) 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)?

44) What is your mass in grams? (Start with your weight in pounds)

45) 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.)

46) If I drive 75 miles/hr, how long in minutes will it take me to drive 500 km?

47) A male elephant seal weighs about 4 tons. What is the mass of the seal in grams? (There are 2000 lbs in one ton.)

Chapter 5 Summary

48) My car gets about 37 miles per gallon. How many km/liter is this?
(There are 4 quarts in a gallon)

49) 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.

5.5: The Mole Review Questions

How many molecules are present in the following quantities?

50) 0.250 mol H₂O

51) 0.0045 mol Al₂(CO₃)₃

How many moles are present in the following quantities?

52) 1×10^{20} molecules H₂O

53) 5 billion atoms of carbon

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

54) H₂O

55) NaOH

56) NH₄Cl

57) H₂SO₄

58) Al₂(CO₃)₃

59) PbO₂

Convert the following to moles.

60) 60.0 g NaOH

61) 5.70 g H₂SO₄

62) 2.73 g NH₄Cl

63) 10.0 g PbO₂

Chapter 5 Summary

Convert the following to grams.

64) 0.100 mol CO₂

65) 0.500 mol (NH₄)₂CO₃

66) 0.437 mol NaOH

67) 3.00 mol H₂O

How many molecules or formula units are present in the following masses?

68) 1.00 g Na₂CO₃

69) 950 g H₂O

What is the mass of each of the following samples?

70) 2.0×10^{23} molecules H₂

75×10^{24} molecules NaCl

77) 8.6×10^{22} molecules NaOH

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Chapter 6

Solutions



Salt water, such as water from the ocean, is a solution, or even mixture, of salt and water. In this chapter, we will explore what solutions are, how they form, and how properties of solutions differ from the original substances that were mixed.

6.1: Solution Formation

Types of Mixtures

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 unbound 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).

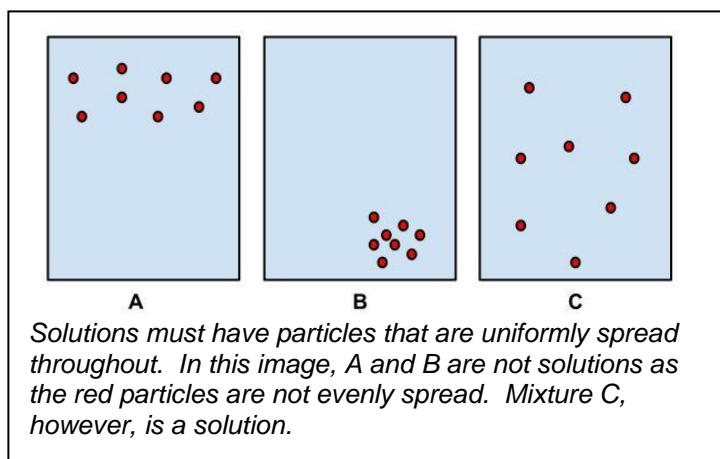
Homogeneous Mixtures

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.

A **Solution** is an even (or homogeneous) mixture of substances. When you consider that the prefix "homo" means "same", this definition makes perfect 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.

Objectives:

- Indicate whether a mixture is a homogeneous mixture (solution) or not
- Identify the solute and solvent in a mixture
- Describe the process through which a solute dissolves in water
- Predict whether two substances will mix to form a solution
- Identify factors that will increase or decrease the rate of dissolving



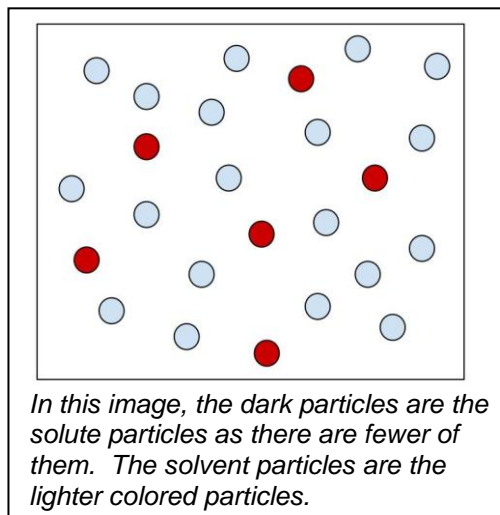
A point should be made here that when a solution is said to have uniform properties throughout, the definition is referring to properties at the particle level. Well, what does this mean? Let's consider brass as an example. The brass is an alloy made from copper and zinc. To the naked eye a brass coin 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

6.1: Solution Formation

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 $\frac{1}{4}$ 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 $\frac{1}{2}$ 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.

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

Solution: A mixture in which the substances are uniformly distributed throughout

Solvent: In a mixture, the substance present in the greater amount

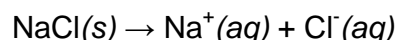
Solute: In a mixture, the substance present in the lesser amount

6.1: Solution Formation

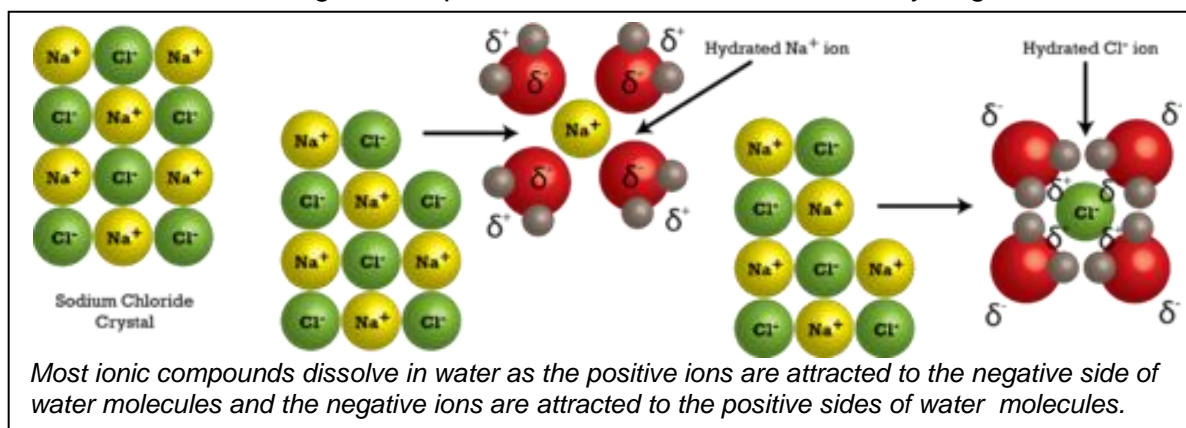
Ionic Compounds in Solution

We all know that salt dissolves in water. Why does it dissolve in water but not vegetable oil? How does salt dissolve in water?

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 charged 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:

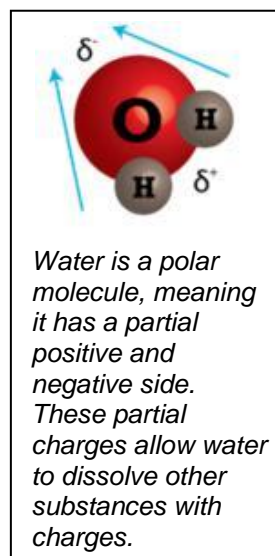


To understand why salt will dissolve in water, we first must remember what it means for water to be polar. The more electronegative oxygen atom pulls the shared electrons away from the hydrogen atoms in a water molecule causing an unequal distribution of electrons. The hydrogen end



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. 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

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).



6.1: Solution Formation

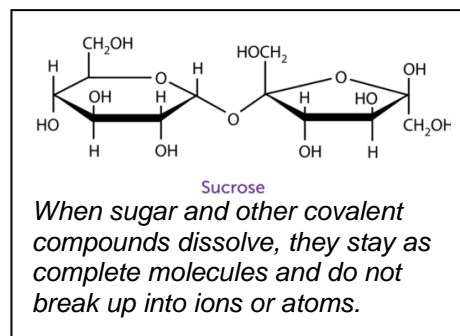
Covalent Compounds in Solution

Some covalent compounds, with dissolve in water (such as rubbing alcohol) but other covalent compounds don't (such as vegetable oil). Why will some dissolve and some not?

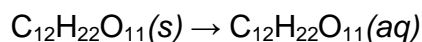
Some other covalent compounds, aside from water, are also polar. Having these partial positive and negative charges within the molecule gives polar compounds the ability to be attracted to water as well. Because of these partial charges, polar molecules are able to dissolve in other polar compounds.

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. You can write the following equation for the dissolution of sugar in water.



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) $\text{C}_2\text{H}_5\text{OH}$

Solution:

LiF will separate into ions when dissolved in solution, because it is an ionic compound.

P_2F_5 and $\text{C}_2\text{H}_5\text{OH}$ are both covalent and will stay as molecules in a solution.

6.1: Solution Formation

Predicting if a Solution will form

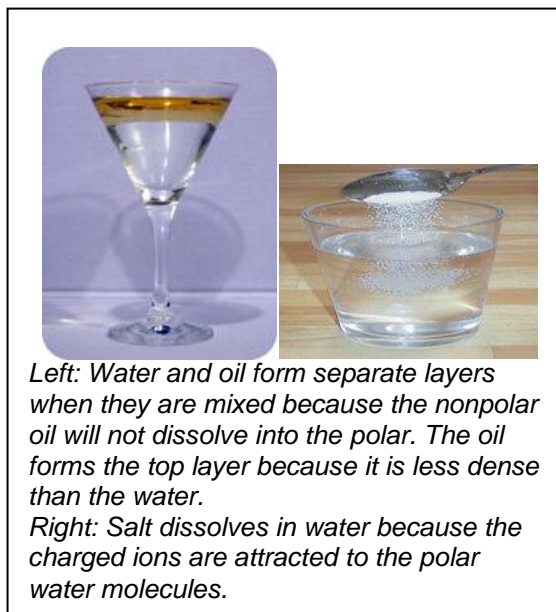
A simple way to predict which compounds will dissolve in other compounds is the phrase “like dissolves like”. What this means is that polar compounds dissolve polar compounds, nonpolar compounds dissolve nonpolar compounds, but polar and nonpolar do not dissolve in each other.

Even some nonpolar substances dissolve in 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.



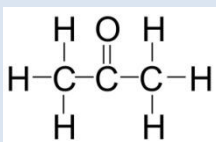
Left: Water and oil form separate layers when they are mixed because the nonpolar oil will not dissolve into the polar. The oil forms the top layer because it is less dense than the water.

Right: Salt dissolves in water because the charged ions are attracted to the polar water molecules.

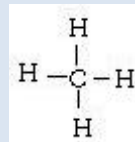
Example: Which of the following would you predict would dissolve in water, a polar solvent??

(a) LiF

(b) Acetone,



(c) Methane,



Solution:

a) This compound is ionic (metal bonded to nonmetal). Most ionic compounds dissolve in water

b) Acetone is polar. “Like dissolves like” means polar acetone will dissolve in water.

c) Methane is nonpolar. Nonpolar compounds will not dissolve in polar water.

6.1: Solution Formation

Rate of Dissolving

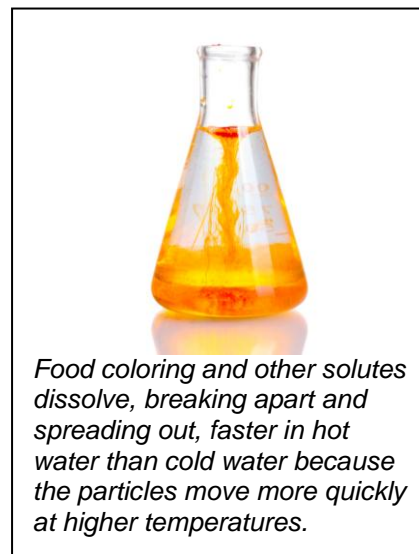
When you make hot chocolate, you must get the sugar and other particles to dissolve and spread throughout evenly. Did you ever get impatient and start drinking the cocoa before all the sugar has dissolved? The first few sips would probably taste watery and there would be cocoa sitting on the bottom of the mug as you finished your drink. What could you do to dissolve the sugar faster?

Stirring

Stirring a solute into a solvent speeds up the rate of dissolving because it helps distribute the solute particles throughout the solvent. For example, when you add sugar to iced tea and then stir the tea, the sugar will dissolve faster. If you don't stir the iced tea, the sugar may eventually dissolve, but it will take much longer.

Temperature

The temperature of the solvent is another factor that affects how fast a solute dissolves. Generally, a solute dissolves faster in a warmer solvent than it does in a cooler solvent because particles have more energy of movement. For example, if you add the same amount of sugar to a cup of hot tea and a cup of iced tea, the sugar will dissolve faster in the hot tea.



Particle Size

A third factor that affects the rate of dissolving is the size of solute particles. For a given amount of solute, smaller particles have greater surface area. With greater surface area, there can be more contact between particles of solute and solvent. For example, if you put granulated sugar in a glass of iced tea, it will dissolve more quickly than the same amount of sugar in a cube. That's because all those tiny particles of granulated sugar have greater total surface area than a single sugar cube. This means that there are more particles on the outside that are able to be attracted to and mix with the solvent particles.



Granulated sugar will dissolve faster than sugar cubes because there are more molecules on the surface able to be attracted by the water.

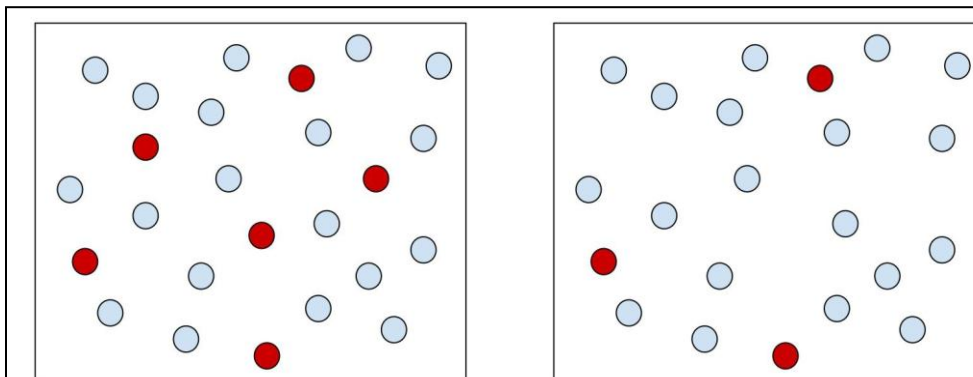
6.2: Concentration

Concentration is a Ratio

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.

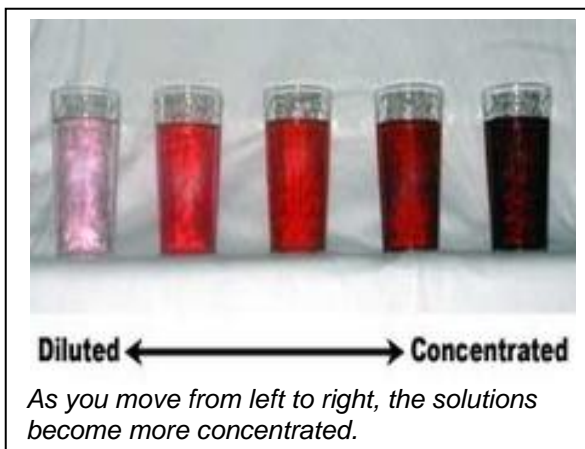
Objectives:

- Describe what is meant by the terms “concentrated” and “dilute”
- Calculate the concentration of a solution in molarity, and molality,
- Indicate when units of parts per million would be a useful unit of concentration.



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)

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 versus 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.

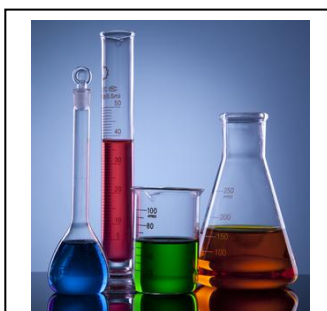
6.2: Concentration

Solving Concentration Problems

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.

Molarity



Molarity is ratio of moles of solute over 1 liter of solution.

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.

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

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

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

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

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

6.2: Concentration

Molality

Molality is another way to measure concentration of a solution. **Molality** is calculated by dividing the number of moles of solute by the number of kilograms of solvent. Molality has the symbol, *m*.

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

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 H₂O

Find: $\text{molality } (m) = \frac{\text{mol HCl}}{\text{kg H}_2\text{O}}$

$$\frac{12.5 \text{ g HCl}}{115 \text{ g H}_2\text{O}} \cdot \frac{1 \text{ mol HCl}}{36.46 \text{ g HCl}} \cdot \frac{1000 \text{ g H}_2\text{O}}{1 \text{ kg H}_2\text{O}} = \frac{2.98 \text{ mol HCl}}{1 \text{ kg H}_2\text{O}} = 2.98 \text{ m HCl}$$

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.

Parts Per Million

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 know it if you see it.

Concentration: A ratio showing the amount of solute compared to the solvent or total amount of solution

Molarity: A ratio giving the number of moles of solute per 1 L of total solution

Molality: A ratio giving the number of moles of solute per 1 kg of solvent

Parts per Millions: A unit of concentration used for very low concentrations, shows the grams of solute per million grams of the solution

6.3: Colligative Properties

What are Colligative Properties?

Objectives:

- Indicate how the boiling or freezing point of a solution differs from that of the original solvent
- Given a pair of solutions, indicate which will have a higher boiling point or lower freezing point.

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?

This is an example of a colligative property. **Colligative properties**

are properties that differ based on

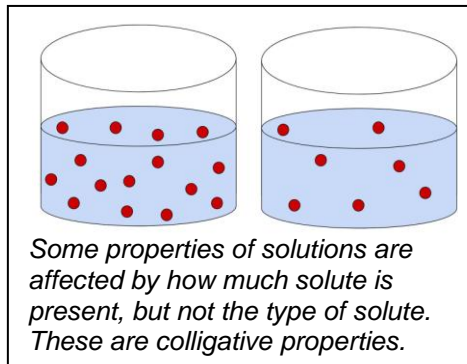
the concentration of solute in a solvent, but not on the type of solute. What this means for the

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.



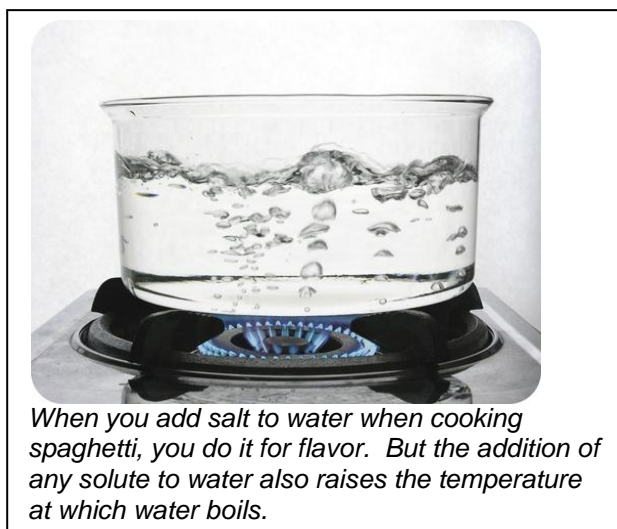
Boiling Point Elevation

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. Furthermore, the more solute that is present, the more the boiling point increases This is called **Boiling Point Elevation**.

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.



Colligative Properties: Properties of a solution that depend on the concentration, not type, of solute present.

6.3: Colligative Properties

Freezing Point Depression

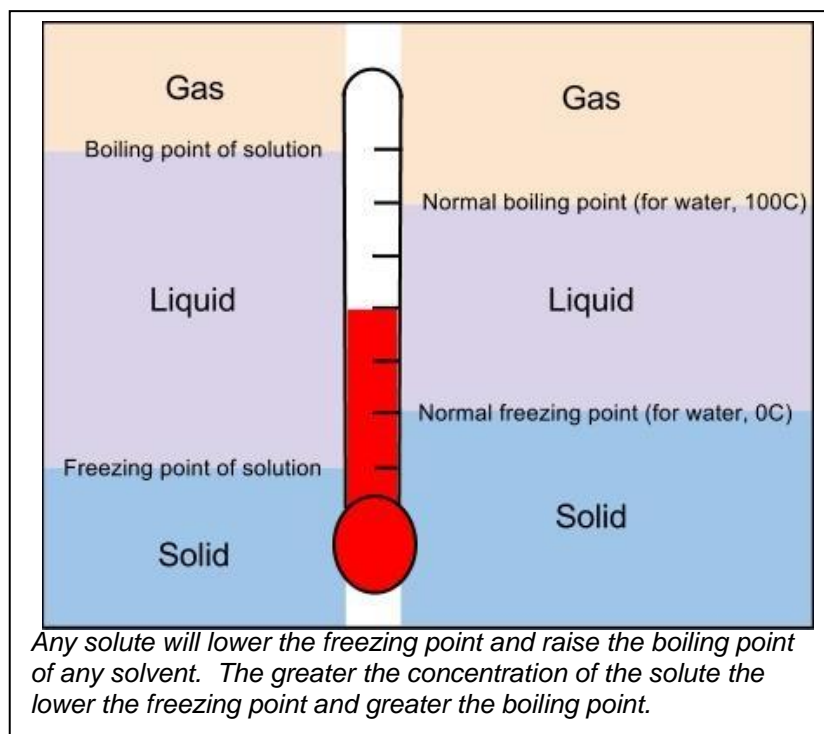
The effect of adding a solute to a solvent has the opposite effect on the freezing point of a solution as it does on the boiling point. A solution will have a lower freezing point than a pure solvent. If a substance is added to a solvent (such as water), the solute-solvent interactions prevent the solvent from going into the solid phase. The solute-solvent interactions require the temperature to decrease further in order to solidify the solution. Any solute added to a solvent will decrease that temperature that the solvent will freeze. Furthermore, the more solute that is present, the more the freezing point decreases. This is called **Freezing Point Depression**.



Salt is frequently put on the roads in the winter because it lowers the freezing point of water making it so that it must be colder before water will freeze to form ice.

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.

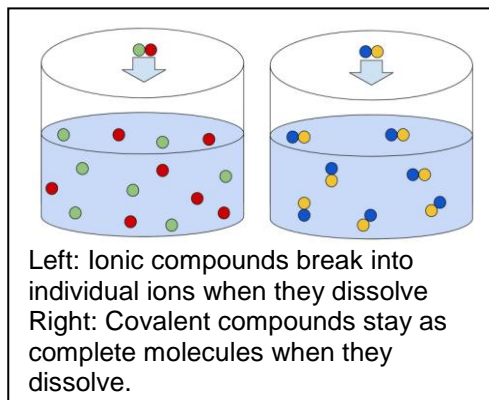
Both freezing point depression and boiling point elevation due to the presence of a solute are colligative properties. That is, the amount of change in the freezing point and boiling point is related to number of particles of solute in a solution and is not related to chemical composition of the solute.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



Any solute will lower the freezing point and raise the boiling point of any solvent. The greater the concentration of the solute the lower the freezing point and greater the boiling point.

6.3: Colligative Properties

Comparing the Boiling and Freezing Points 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.

We will need to be able to count the number of particles into which a formula dissolves. For example, sodium chloride, NaCl, will dissociate into two ions. Lithium nitrate, LiNO_3 , will also dissolve into two parts – Li^+ and NO_3^- ions. Calcium chloride, CaCl_2 , will dissolve into 3 ions. Covalent compounds always dissolve into particles of 1 complete formula.

To compare properties of solutions, follow these general steps:

1. Label each solute as ionic or covalent. If the solute is ionic, determine the number of *ions* in the formula. Be careful to look for polyatomic ions.
2. 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.
3. 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:

$0.1\ m\ \text{NaCl}$ $0.1\ m\ \text{C}_6\text{H}_{12}\text{O}_6$ $0.1\ m\ \text{CaCl}_2$

Solution:

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

- $0.1\ m\ \text{NaCl}$: this compound is ionic (metal with nonmetal), and will dissolve into 2 parts. The total final concentration is: $(0.1\ m)(2) = 0.2\ m$
- $0.1\ m\ \text{C}_6\text{H}_{12}\text{O}_6$: this compound is covalent (nonmetal with nonmetal), and will stay as 1 part. The total final concentration is: $(0.1\ m)(1) = 0.1\ m$
- $0.1\ m\ \text{CaCl}_2$: this compound is ionic (metal with nonmetal), and will dissolve into 3 parts. The total final concentration is: $(0.1\ m)(3) = 0.3\ m$

Remember, the greater the concentration of particles, the lower the freezing point will be. $0.1\ m\ \text{CaCl}_2$ will have the lowest freezing point, followed by $0.1\ m\ \text{NaCl}$, and the highest of the three solutions $0.1\ m\ \text{C}_6\text{H}_{12}\text{O}_6$. All of them will have a lower freezing point than water.

Chapter 6 Summary

6.1: Solution Formation

- A solution is a mixture that has the same properties throughout.
- 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).
- 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).
- When ionic compounds dissolve, they separate into individual ions. When covalent compounds dissolve they separate into complete molecules.
- 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.
- 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).
- The rate at which a solute dissolves depends on the temperature, particle size (surface area) and whether or not the mixture is stirred.

6.2: Concentration

- Concentration is the measure of how much of a given substance is mixed with another substance.
- Molarity is the number of moles of solute per liter of solution.
- Molality is calculated by dividing the number of moles of solute by the kilograms of solvent. It is less common than molarity but more accurate because of its lack of dependence on temperature.
- Parts per million is a useful unit for very low concentrations.

6.3: Colligative Properties of Solutions

- 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.
- Ionic compounds split into ions when they dissolve, forming more particles. Covalent compounds stay as complete molecules when they dissolve.
- Boiling points of solutions are higher than the boiling points of the pure solvents.
- Freezing points of solutions are lower than the freezing points of the pure solvents.

6.1: Solution Formation Review Questions

- 1) What does the phrase “like dissolves like” mean? Give an example.

- 2) Why will LiCl not dissolve in CCl₄ but it will dissolve in water?

Chapter 6 Summary

- 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.
- 5) When making a chocolate-flavored drink, you must dissolve the chocolate syrup in water or milk. List 3 ways you can make it dissolve faster.

6.2: Concentration Review Questions

- 6) 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?
- 7) 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?
- 8) 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.
- 9) 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?
- 10) 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?
- 11) Find the mass of aluminum nitrate, $Al(NO_3)_3$, required to mix with 750g of water to make a 1.5*m* solution.

Chapter 6 Summary

- 12) 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?
- 13) What is the molarity of a solution prepared by mixing 12.5 grams FeCl₃ in enough water to make 300 mL of solution?
- 14) If 5 grams of NaCl are mixed in enough water to make .5L of solution. What is the molarity of the solution?
- 15) What is the molality of a solution made by mixing 15 grams of Ba(OH)₂ in 250 grams of water?
- 16) A solution is made by mixing 10.2 grams of CaCl₂ in 250 grams of water. What is the molality of the solution?

6.3: Colligative Properties Review Questions

- 17) 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.
- 18) 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?)
- 19) We use sodium chloride, NaCl, on the roads in the winter. Would a different compound, such as CaCl₂ or table sugar have the same effect on the ice? Why or why not?

Chapter 6 Summary

20) Besides adding flavor, what effect does adding salt to water that you cook spaghetti in?

21) 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.

22) Salt, NaCl

23) Acetone, C₂H₂O

24) Benzene, C₆H₆

25) Copper(II) nitrate, Cu(NO₃)₂

26) AlCl₃

27) Potassium hydroxide, KOH

For each pair of solutions, indicate which would have a lower freezing point. Which would have a higher boiling point?

28) 0.20 m KI or pure water

29) 0.2 m KI or 0.2 m CaCl₂

30) 0.1 m KI or 0.1 m C₆H₁₂O₆

31) 0.2 m NaCl or 0.3 m C₆H₁₂O₆

Chapter 7

Reactions



Communities often use fireworks to celebrate important occasions. Fireworks certainly create awesome sights and sounds! Do you know what causes the brilliant lights and loud booms of a fireworks display? The answer is chemical changes.

7.1: Physical and Chemical Change

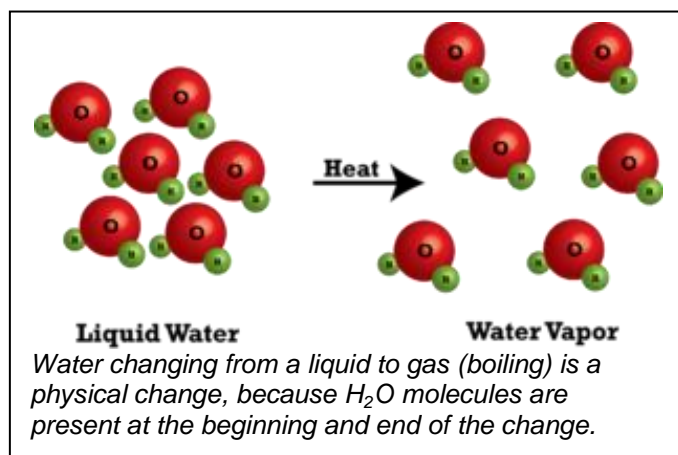
Objectives:

-Distinguish between physical and chemical changes.

Physical Change

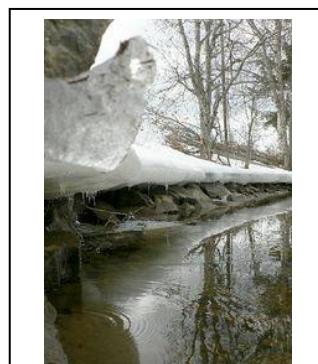
Change is happening all around us all of the time. Just as chemists have classified elements and compounds, they have also classified types of changes. Changes are either classified as physical or chemical changes.

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:

- 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.

7.1: Physical and Chemical Change

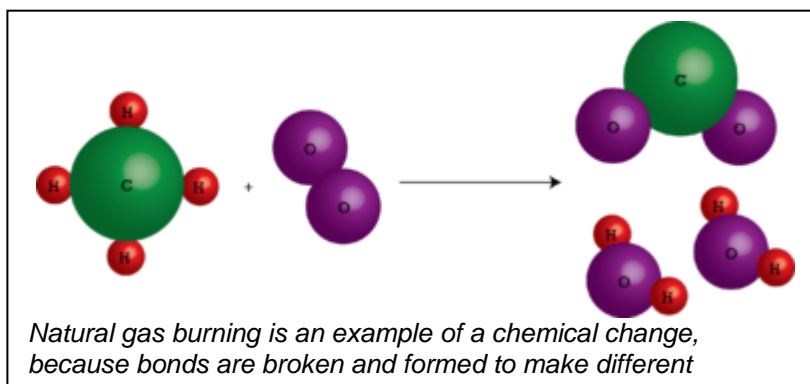
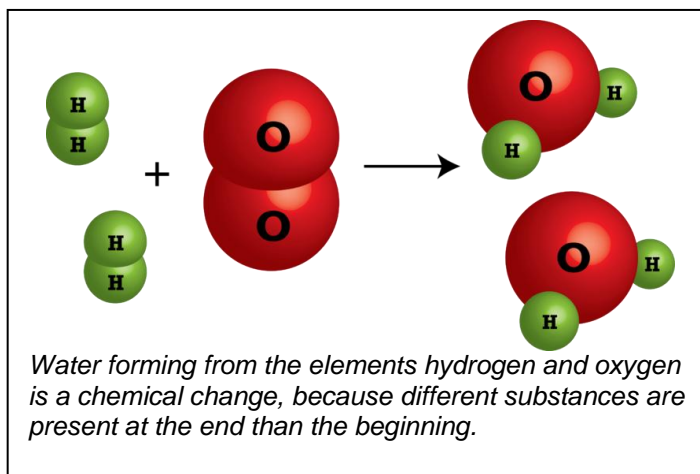
Chemical Change

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 different properties.

Chemical changes are frequently harder to reverse than physical changes.

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.



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, CH₄, and two molecules of oxygen, O₂, while on the right we have two molecules of water, H₂O, and one molecule of carbon dioxide, CO₂. 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.



Firework displays are an example of a chemical change.

7.1: Physical and 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:

Release of Bubbles	Change of Color	Production of an Odor	Release of Heat and Light	Production of Loud Sounds
				
Bubbles are released when a chemical change produces a gas. The bubbles in this test tube were released when vinegar was added to baking soda. When the two substances combine, they change to water and the gas carbon dioxide.	These rusty pipes were once silver-colored. What happened to them? Iron in the pipes combined with oxygen in the air to produce a new substance—iron oxide—which is reddish brown. Iron oxide is commonly called rust.	You can tell that the food in this can has a stinky odor! When food spoils, it undergoes chemical changes that release unpleasant odors.	Burning is a chemical change that releases both heat and light. When a substance such as candle wax burns, it combines with oxygen and changes to other substances, including carbon dioxide and water vapor.	Gunshots are very loud sounds. They occur because explosive chemical changes take place inside the gun when the shooter pulls the trigger. The changes also propel a bullet out of the end of the gun.

Other observations may include the formation of a precipitate. A precipitate is a solid that is formed when two liquids are mixed together. The solid will cause the solution to look cloudy.

Example: Label each of the following changes as a physical or chemical change. Given evidence to support your answer.

- boiling water
- a nail rusts
- a green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
- two colorless solutions are mixed. The resulting mixture has a yellow precipitate

Solution:

- physical: boiling and melting are physical changes. When water boils no bonds are broken or formed. The change could be written: $\text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_2\text{O}(\text{g})$
- 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: $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
- 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.
- chemical: the formation of a precipitate and the color change from colorless to yellow indicates a chemical change.

Physical Change: A change in which particles may move, but the same substances are present at the end of the change.

Chemical Change: A change in which bonds are broken and/or formed, resulting in the formation of different substances with different properties.

7.2: Reaction Rate

Collision Theory and Reaction Rate

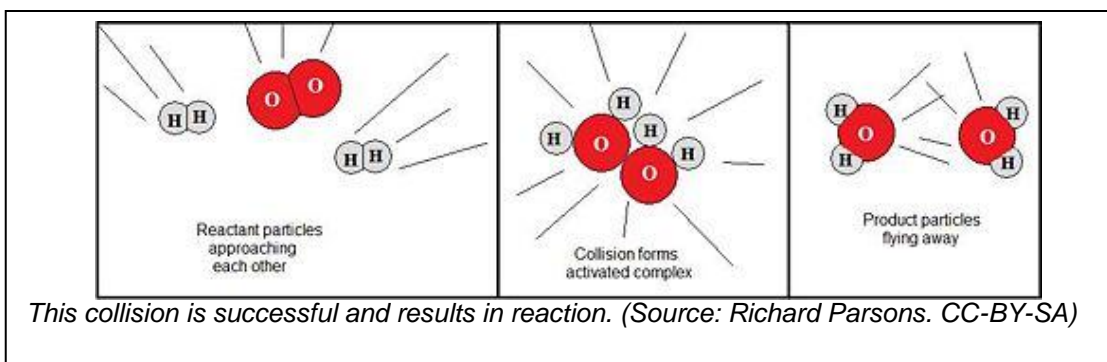
We know that reactions occur when bonds are broken and/or formed. We also know that there are things we can do, such as heating a reaction up, that will speed up or slow down how quickly this change occurs. Chemists had to come up with a theory that satisfactorily could explain these observations.

The **collision theory** states that the rate of a reaction depends on the frequency of effective collision. What conditions are necessary for an “effective collision”? These conditions include:

- 1) The particles must collide with each other.
- 2) The particles must have sufficient energy to break the old bonds.
- 3) The particles must have proper orientation.

Objectives:

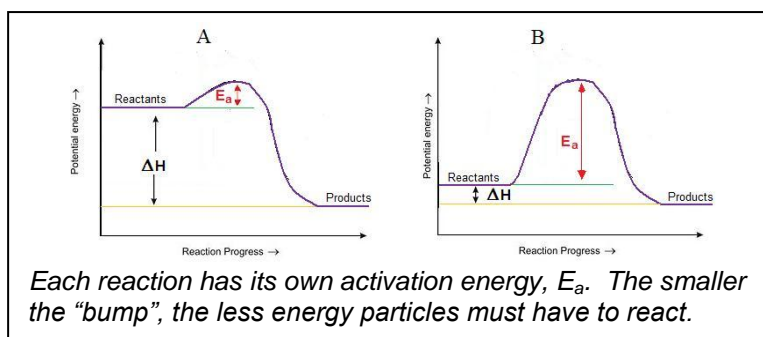
- Relate the frequency of effective collisions to the rate of a reaction.
- Indicate ways to speed up or slow down a reaction and explain in terms of collision theory
- Describe the effect of a catalyst on the rate of a reaction



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. 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.



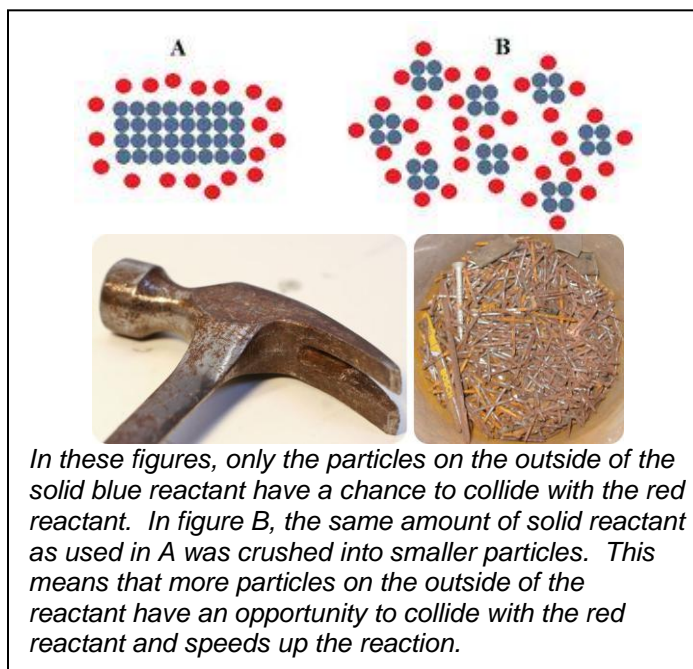
Remember, a successful collision occurs when two reactants collide with enough energy and with the right orientation. That means if we can do things that will increase the number of collisions, increase the number of particles that have enough energy to react and/or increase the number of particles with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

7.2: Reaction Rate

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. Increasing the frequency of collisions will speed up a reaction. 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. 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 Figure 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.



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. 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, and so it is much easier to get the kindling burning than if you had tried to start one big log.

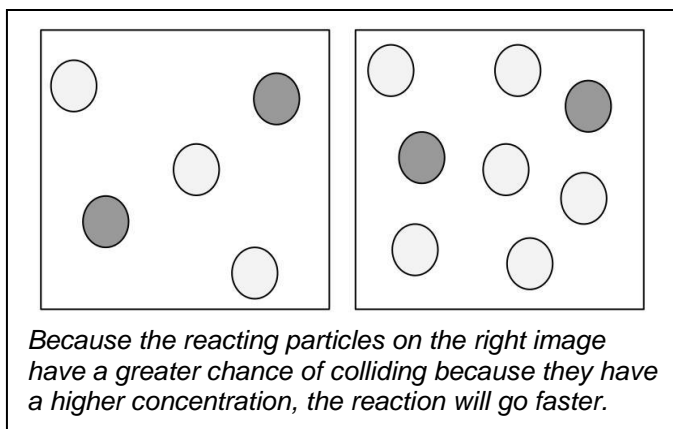
There have been, unfortunately, cases where serious accidents were caused by the failure to understand the relationship between surface area and reaction rate. One such example occurred in flour mills. A grain of wheat is not very flammable. It takes a significant effort to get a grain of wheat to burn. 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.

7.2: Reaction Rate

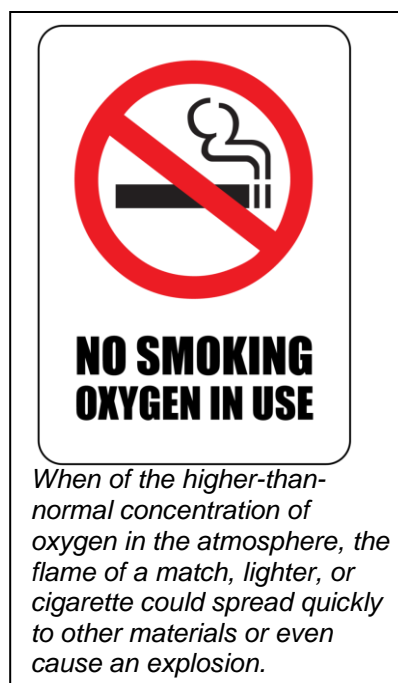
Effect of Concentration on Rate of Reaction

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, 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. If one concentration of one of the reactants is doubled, the number of collisions will also double. 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.



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.



7.2: Reaction Rate

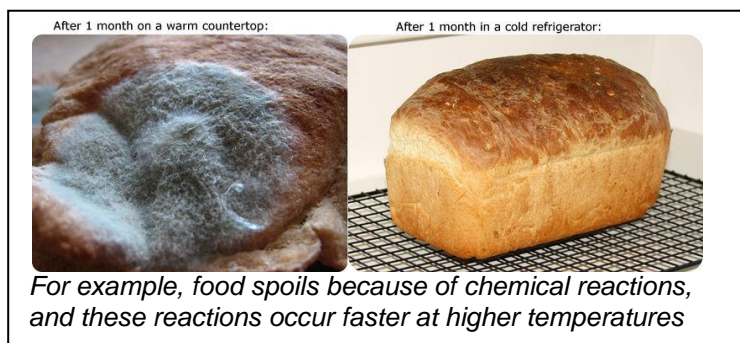
Effect of Temperature on Rate of Reaction

When the temperature is increased, the average velocity of the particles is increased. The result is that the particles will collide more frequently. 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 lower temperatures fewer particles are able to overcome the activation energy and particles merely bounce off each other without reacting.

At any one moment in the atmosphere, there are many collisions occurring between hydrogen and oxygen. However, in the atmosphere these do not react to form water, 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: $O_2(g) + H_2(g) \rightarrow H_2O(l)$

Society uses the effect of temperature on rate every day. We store food in freezers and refrigerators to slow down the processes that cause it to spoil. 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. Some molecules do have enough energy and are colliding so over time even in the refrigerator the milk will spoil. 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.



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

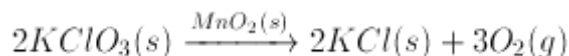
ways of surviving. One method was to store their food in the snow to be used later during their advances to the pole. On some explorations, they buried so much food, that they didn't need to use all of it and it was left. Many years later, when this food was located and thawed, it was found to still be edible.

7.2: Reaction Rate

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.

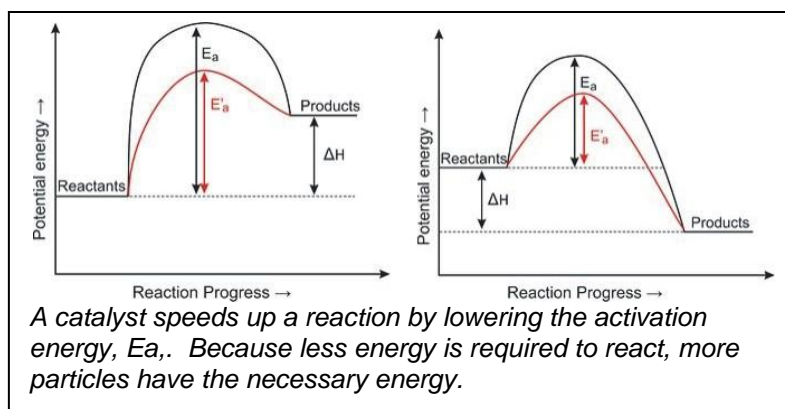
In the reaction of potassium chlorate breaking down to potassium chloride and oxygen, a catalyst is available to make this reaction occur much faster than it would occur by itself under room conditions. The reaction is:



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 and its condition is exactly the same as it was before the reaction. This is part of the definition of a catalyst. Catalysts are 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 arrow.

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.

Look at the energy diagram shown in the figure. Without a catalyst, the activation energy is large – particles must have a lot of energy when they collide in order to react. With a catalyst, the activation energy is much smaller, meaning that colliding particles are more likely to have enough energy to react. There is no effect on the energy of the reactants, the products, or the value of ΔH .



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. 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 within cells, 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.

Catalyst: A substance that speeds up a chemical reaction by lowering the activation energy, without being used up in the process

7.3: Writing Chemical Equations

Various Ways to Represent Chemical Changes

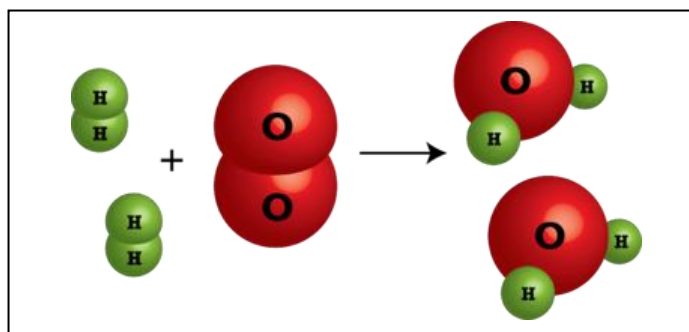
Objectives:

-Correctly write a chemical reaction in words, symbols, or pictures.

We already discussed that 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.

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

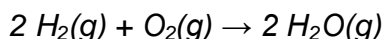
1. They could draw a picture of the chemical reaction.



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.”

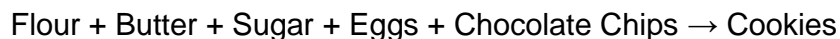
3. They could write the equation in chemical shorthand.



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.

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**. 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.



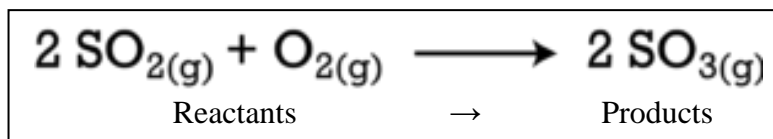
Reactant: The substances present at the beginning of a change

Products: The substances present at the end of a change

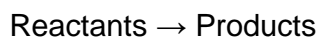
7.3: Writing Chemical Equations

Meanings of Symbols in Chemical Equations

When sulfur dioxide is added to oxygen, sulfur trioxide is produced. Sulfur dioxide and oxygen, $\text{SO}_2 + \text{O}_2$, are reactants and sulfur trioxide, SO_3 , is the product.



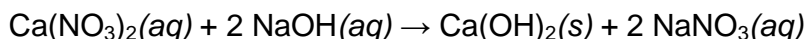
In chemical reactions, the reactants are found before the symbol “→” and the products and found after the symbol “→”. The general equation for a reaction is:



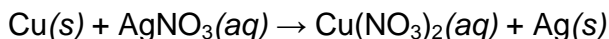
There are a few special symbols that we need to know in order to “talk” in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. You will find there are others but these are the main ones that we need to know.

Common Symbols in Chemical Reactions		
Symbol	Meaning	Example
→	Used to separate reactants from products; can be read as "to produce" or "yields".	$2 \text{H}_2 + \text{O}_2 \rightarrow 2 \text{H}_2\text{O}$
+	Used to separate reactants from each other or products from each other; can be read as "is added to" or “also forms”.	$\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$
(s)	in the solid state	sodium in the solid state = $\text{Na}(s)$
(l)	in the liquid state	water in the liquid state = $\text{H}_2\text{O}(l)$
(g)	in the gaseous state	carbon dioxide in the gaseous state = $\text{CO}_2(g)$
(aq)	in the aqueous state, dissolved in water	sodium chloride solution = $\text{NaCl}(aq)$

In words, 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:



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.



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".

7.3: Writing Chemical Equations

Writing Chemical Equations from Words

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

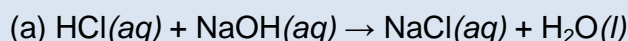
1. Identify the reactants and products. This will help you know what symbols go on each side of the arrow (\rightarrow) and where the + signs go.
2. Write the correct formulas for all substances.
 - a. If an element is listed, you will need to write the symbol for the element. For diatomic elements, write their formulas. (See the table).
 - b. If a compound is listed (its name usually have two words) you will need to use the rules you learned about writing charge-balanced formulas.

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

Doing these two steps will correctly identify products and reactants and their formulas, but this reaction is not actually completely written yet. In order to be useful, a reaction must also be balanced. We will worry about that in the next section. In this section, just focus on getting the correct formulas for all reactants and products.

Follow these steps in the given examples.

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



(b) Gaseous propane, C₃H₈, burns in oxygen gas to produce gaseous carbon dioxide and liquid water.

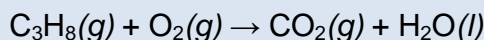
(c) Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

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 (C₃H₈) and oxygen (O₂)

Products: carbon dioxide (CO₂) and water (H₂O)



(c) Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

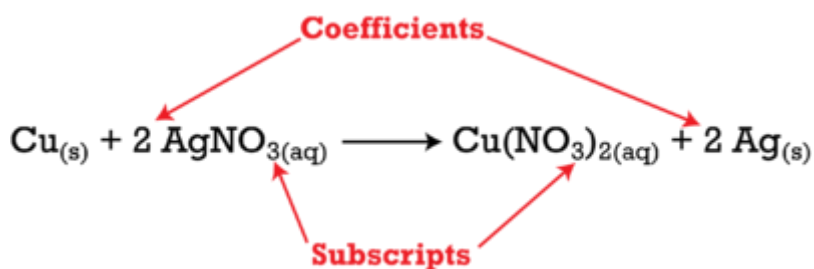


7.4: Balancing Chemical Equations

How to Balance 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 a **balanced equation**. 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.



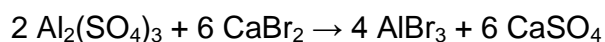
The subscripts are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The coefficients indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates 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*.

1. Write the correct formulas for your reactants and products.
2. 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.

Chemical equations should be balanced with the simplest whole number coefficients that balance the equation. The following equation is *not* balanced as the coefficients could be reduced to smaller whole numbers.



Objectives:

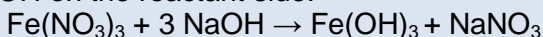
- Balance a chemical reaction using the lowest whole-number coefficients
- Give evidence for the law of conservation of mass
- Describe the importance of a closed system in conserving mass in a chemical reaction

7.4: Balancing Chemical Equations

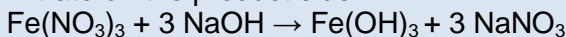
Examples Balancing Equations

Example: Balance the following equation: $\text{Fe}(\text{NO}_3)_3 + \text{NaOH} \rightarrow \text{Fe}(\text{OH})_3 + \text{NaNO}_3$

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



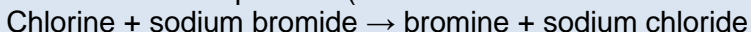
Then we can balance the nitrate ions by inserting a coefficient of 3 in front of the sodium nitrate on the product side.



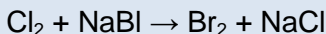
Example: Write a balanced equation for the reaction that occurs between chlorine gas and aqueous sodium bromide to produce liquid bromine and aqueous sodium chloride.

Solution:

Step 1: Write the word equation. (Be careful of diatomic molecules.)

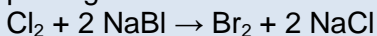


Step 2: Substitute the correct formulas into the equation.



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

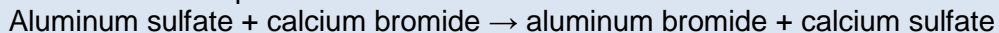
By placing a coefficient of 2 in front of the NaBr, we can balance the bromine atoms and by placing a coefficient of 2 in front of the NaCl, we can balance the Cl atoms.



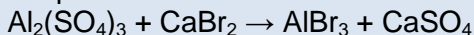
Example: Write a balanced equation for the reaction between aluminum sulfate and calcium bromide to produce aluminum bromide and calcium sulfate.

Solution:

Step 1: Write the word equation.

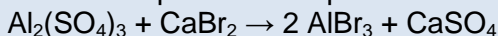


Step 2: Replace the names of the substances in the word equation with formulas.

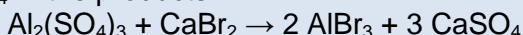


Step 3: Insert coefficients to balance the equation.

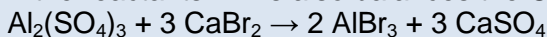
In order to balance the aluminum atoms, we must insert a coefficient of 2 in front of the aluminum compound in the products.



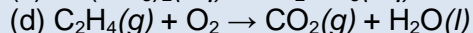
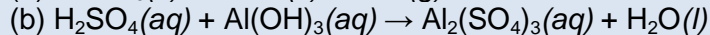
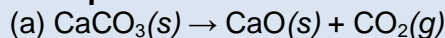
In order to balance the sulfate ions, we must insert a coefficient of 3 in front of the CaSO₄ in the products.



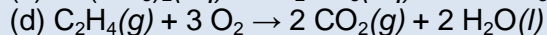
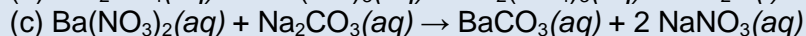
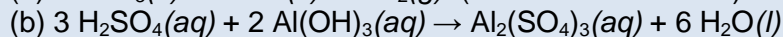
In order to balance the bromine atoms, we must insert a coefficient of 3 in front of the CaBr₂ in the reactants. This also balances the Ca atoms.



Example: Balance each of the following reactions.



Solutions

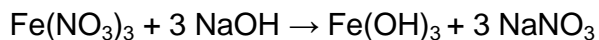


7.4: Balancing Chemical Equations

Conservation of Mass in Chemical Reactions

Previously we discussed 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).

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.



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

$$1 \text{ moles of } \text{Fe}(\text{NO}_3)_3 \times \text{molar mass} = (1 \text{ mol})(241.9 \text{ g/mol}) = 241.9 \text{ g}$$

$$3 \text{ moles of } \text{NaOH} \times \text{molar mass} = (3 \text{ mol})(40.0 \text{ g/mol}) = 120. \text{ g}$$

$$\text{Total mass for reactants} = 241.9 \text{ g} + 120. \text{ g} = 361.9 \text{ g}$$

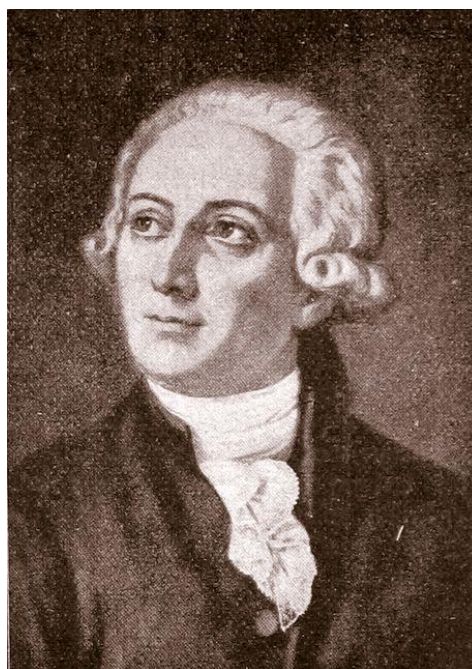
Product Side Mass

$$1 \text{ moles of } \text{Fe}(\text{OH})_3 \times \text{molar mass} = (1 \text{ mol})(106.9 \text{ g/mol}) = 106.9 \text{ g}$$

$$3 \text{ moles of } \text{NaNO}_3 \times \text{molar mass} = (3 \text{ mol})(85.0 \text{ g/mol}) = 255 \text{ g}$$

$$\text{Total mass for products} = 106.9 \text{ g} + 255 \text{ g} = 361.9 \text{ g}$$

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.



Obtained from:
http://upload.wikimedia.org/wikipedia/commons/7/78/Antoine_laurent_lavoisier.jpg

7.5: Types of Reactions

Synthesis Reactions

Objectives:

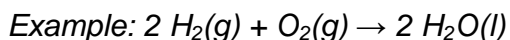
- Label a reaction as synthesis, decomposition, single replacement, double replacement or combustion.
- Predict products of reactions based on their

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.

A **synthesis reaction** is one in which two or more reactants combine to make one type of product.

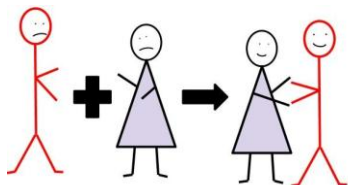


Synthesis reactions occur as a result of two or more simple elements or molecules combining to form a more complex molecule. Look at the example below. Here two elements (hydrogen and oxygen) are combining to form one product (water).



We can always identify a synthesis reaction because there is only one product of the reaction.

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.



A synthesis reaction is similar to forming a couple, which behaves and acts differently than the two single individuals.
CC – Tracy Poulsen

Example:

- (a) Write the chemical equation for the synthesis reaction of silver bromide, AgBr.
(b) Predict the products for the following reaction: $\text{Na}(\text{s}) + \text{O}_2(\text{g})$

Solution:

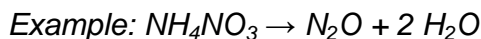
- (a) $2 \text{Ag} + \text{Br}_2 \rightarrow 2 \text{AgBr}$ (b) $\text{Na}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{Na}_2\text{O}$

Decomposition Reactions

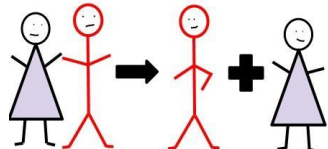
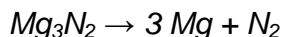
When one type of reactant breaks down to form two or more products, we have a **decomposition reaction**. The best way to remember a decomposition reaction is that for all reactions of this type, there is only one reactant.



In the following example, 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



When studying decomposition reactions, look at the formula for the reactants. In the following example, magnesium nitride decomposes into its elements, magnesium and nitrogen.



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

Example: Write the chemical equation for the decomposition of:

- (a) Al_2O_3 (b) Ag_2S (c) MgO

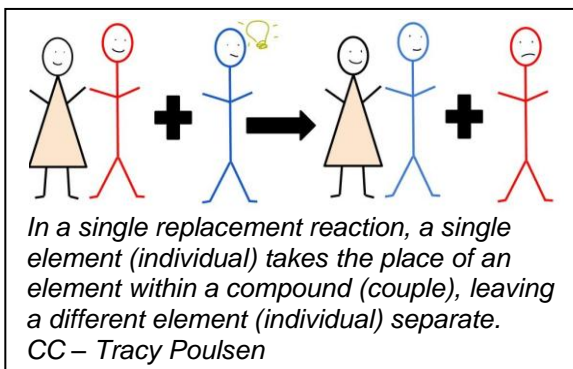
Solution:

- (a) $2 \text{Al}_2\text{O}_3 \rightarrow 4 \text{Al} + 3 \text{O}_2$ (b) $\text{Ag}_2\text{S} \rightarrow 2 \text{Ag} + \text{S}$ (c) $2 \text{MgO} \rightarrow 2 \text{Mg} + \text{O}_2$

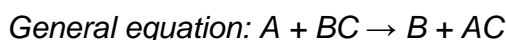
7.5: Types of Reactions

Single Replacement Reactions

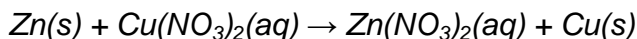
A third type of reaction is the single replacement reaction. In **single replacement reactions** one element reacts with one compound to form products. The single element is said to replace an element in the compound when products form, hence the name



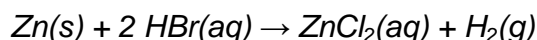
single replacement. Metal elements will always replace other metals in ionic compounds or hydrogen in an acid. Nonmetal elements will always replace another nonmetal in an ionic compound.



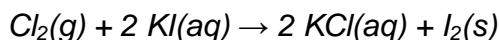
Consider the following examples.



Notice that the metal element, Zn, replaced the metal in the compound $\text{Cu(NO}_3)_2$. 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.



When a metal element is mixed with acid, the metal will replace the hydrogen in the acid and release hydrogen gas as 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^- .



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:

- 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)
- Predict the products for the following reaction: $\text{Fe} + \text{CuSO}_4$ (in this reaction, assume iron forms ions with a +2 charge)
- Predict the products for the following reaction: $\text{Al} + \text{CuCl}_2$
- Complete the following reaction. Then balance the equation: $\text{Al} + \text{HNO}_3 \rightarrow$

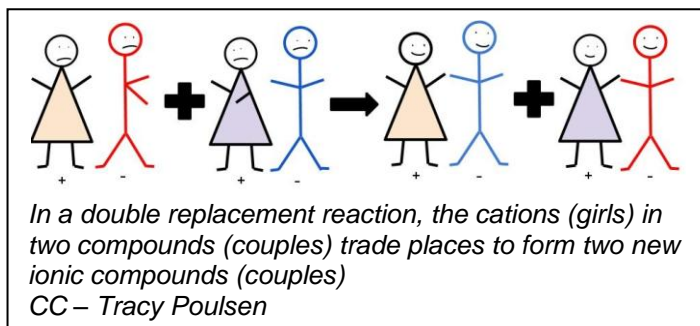
Solution:

- $\text{Zn} + \text{Pb(NO}_3)_2 \rightarrow \text{Pb} + \text{Zn(NO}_3)_2$
- $\text{Fe} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{FeSO}_4$
- $2 \text{Al} + 3 \text{CuCl}_2 \rightarrow 3 \text{Cu} + 2 \text{AlCl}_3$
- $2 \text{Al} + 6 \text{HNO}_3 \rightarrow 2 \text{Al(NO}_3)_3 + 3 \text{H}_2$

7.5: Types of Reactions

Double Replacement Reactions

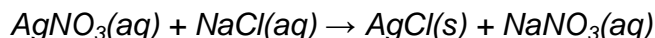
For **double replacement reactions** two ionic compound reactants will react by having the cations exchange places, forming two new ionic compounds. The key to this type of reaction, as far as identifying it over the other types, is that it has two compounds as reactants. This type of reaction is more 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.



For example, when solutions of silver nitrate and sodium chloride are mixed, the following reaction occurs:



This is an example of a precipitate reaction. Notice that two aqueous reactants form one solid, the precipitate, and another aqueous product.

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.

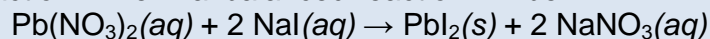
Example:

- Predict the products: $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{NaI}(\text{aq}) \rightarrow$
- 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.
- Predict the products for the following reaction: $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow$
- Predict the products for the following reaction: $\text{FeCl}_3(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow$

Solution:

(a) We will have to look at the charges of each of the cations and anions to see what the products will be. In $\text{Pb}(\text{NO}_3)_2$, made of NO_3^- and Pb^{2+} . NaI is made of Na^+ and I^- .

Switch ions and write the correct subscripts so the total charge of each compound is zero. The Pb^{2+} and I^- will form PbI_2 . The Na^+ and NO_3^- will 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:



- $\text{CaCl}_2(\text{aq}) + 2 \text{KOH}(\text{aq}) \rightarrow \text{Ca}(\text{OH})_2(\text{s}) + 2 \text{KCl}(\text{aq})$
- $\text{AgNO}_3(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{AgCl}(\text{s}) + \text{NaNO}_3(\text{aq})$
- $\text{FeCl}_3(\text{aq}) + \text{KOH}(\text{aq}) \rightarrow \text{Fe}(\text{OH})_3(\text{s}) + \text{KCl}(\text{aq})$

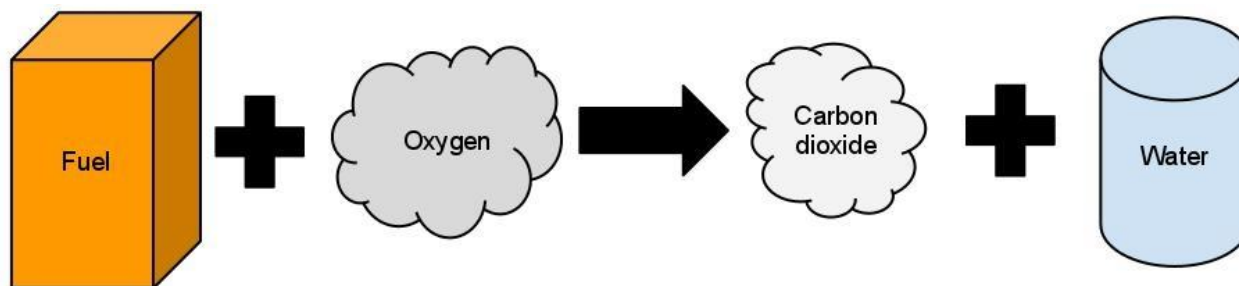
7.5: Types of Reactions

Combustion

In a **combustion reaction** oxygen reacts with fuel to produce carbon dioxide and water. This is what happens when fuel burns. 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

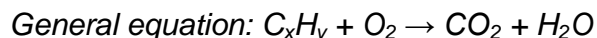


Burning a paper is an example of a combustion reaction

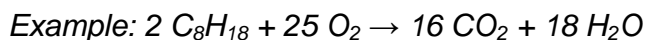


*In a combustion reaction, a fuel (C_xH_y compound) reacts with oxygen (O_2) to form CO_2 and H_2O .
CC – Tracy Poulsen*

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.



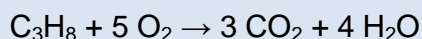
Look at the reaction for the combustion of octane, C_8H_{18} , below. Octane has 8 carbon atoms hence the prefix “oct”.



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.



7.6: Stoichiometry

Coefficients Show Quantity Ratios

Objectives:

-Use coefficients to show the molar quantities of substances in a chemical reaction
 -Calculate the amount of a reactant or product involved in a chemical reaction using the factor-label method

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.

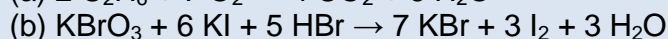
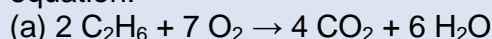
The mole, as you remember, is a quantitative measure that is equivalent to Avogadro's number of particles. So how does this relate to the chemical equation? Look at the chemical equation below.

2 CuSO ₄	+	4 KI	→	2 CuI	+	2 K ₂ SO ₄	+	I ₂
2 formula units CuSO ₄	+	4 formula units KI	→	2 formula units CuI	+	2 formula units K ₂ SO ₄	+	1 molecule I ₂
2 moles CuSO ₄	+	4 moles KI	→	2 moles CuI	+	2 moles K ₂ SO ₄	+	1 moles I ₂

The coefficients used, as we have learned, tell us the relative numbers of each substance in the equation. So for every 2 units of copper (II) sulfate (CuSO₄) 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.

It is important to realize that these coefficients always show numbers, or quantities, of reactants and products. These coefficients do not show mass relationships or the relationships of volumes for various liquids and solids in the reaction.

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.



Solution:

(a) Two molecules of C₂H₆ plus seven molecules of O₂ yields four molecules of CO₂ plus six molecules of H₂O.

Two moles of C₂H₆ plus seven moles of O₂ yields four moles of CO₂ plus six moles of H₂O.

(b) Two formula units of KBrO₃ plus six formula units of KI plus six formula units of HBr yields seven formula units of KBr plus three molecules of I₂ and three molecules of H₂O.

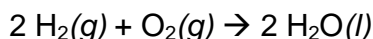
Two moles of KBrO₃ plus six moles of KI plus six moles of HBr yields seven moles of KBr plus three moles of I₂ and three moles of H₂O.

7.6: Stoichiometry

Mole:Mole Ratios

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.

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:



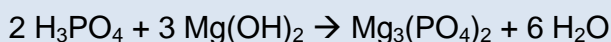
2 moles of H_2 react with 1 mole of O_2 to produce 2 moles of H_2O . Or, an alternate method to represent this information is with mole ratios. The following mole ratios can be obtained from this reaction:

$$\frac{2 \text{ mol H}_2}{1 \text{ mol O}_2} \text{ or } \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2} \text{ or } \frac{2 \text{ mol H}_2\text{O}}{2 \text{ mol H}_2} \text{ or } \frac{2 \text{ mol H}_2\text{O}}{1 \text{ mol O}_2} \text{ or } \frac{2 \text{ mol H}_2}{2 \text{ mol H}_2\text{O}} \text{ or } \frac{1 \text{ mol O}_2}{2 \text{ mol H}_2\text{O}}$$

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.

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, $\text{Mg}(\text{OH})_2$, is present, how many moles of phosphoric acid, H_3PO_4 , would be required for the reaction?



Solution: We need to set up this problem using the same steps of dimensional analysis.

Given: 0.050 mol $\text{Mg}(\text{OH})_2$

Find: mol H_3PO_4

The ratio we need is one that compares mol $\text{Mg}(\text{OH})_2$ 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.

$$0.050 \text{ mol Mg}(\text{OH})_2 \cdot \frac{2 \text{ mol H}_3\text{PO}_4}{3 \text{ mol Mg}(\text{OH})_2} = \mathbf{0.033 \text{ mol H}_3\text{PO}_4}$$

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.

It is also important to remember that the coefficients show the *number* of each molecules compared to the others. Coefficients do not show the mass ratios of substances in a reaction.

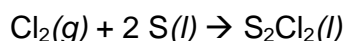
7.6: Stoichiometry

Calculations Using a Mole Map

Balanced reactions do *not* show how masses of different substances are related. Nor do they show how volume or other measurements are related. However, combining the balanced reaction (which does compare moles of different substances) with other ratios will allow us to solve a wide variety of problems. We can 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.

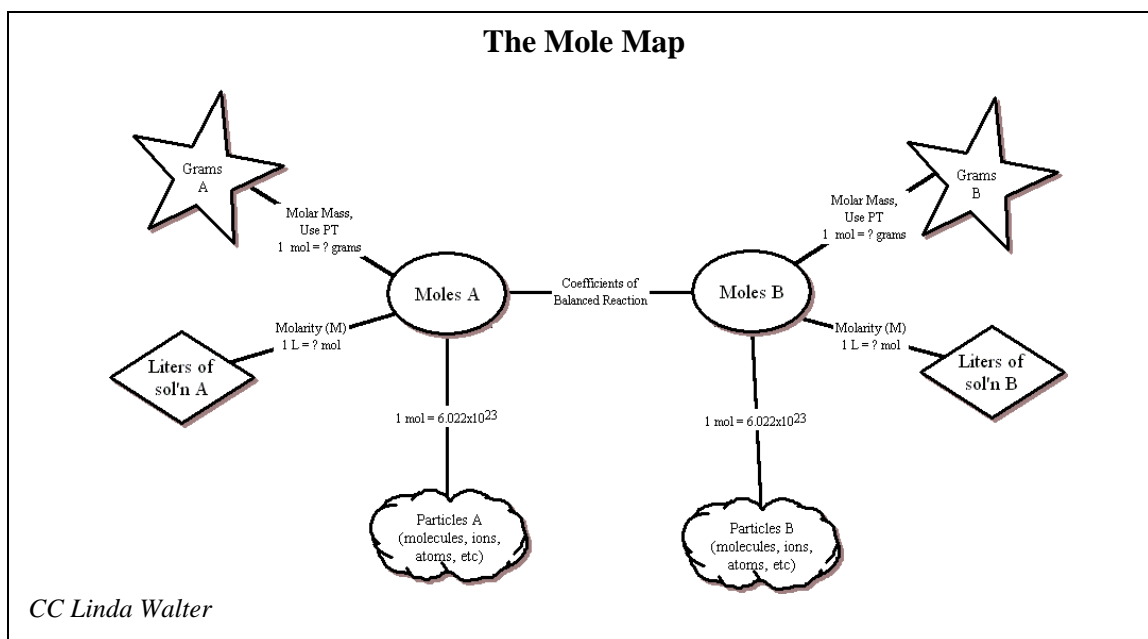
For example, 15.0 g of chlorine gas is bubbled over liquid sulfur to produce disulfur dichloride. How much sulfur, in grams, is needed according to the balanced equation:



1. Identify the **given**: 15.0 g Cl₂
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.

$$15.0 \text{ g Cl}_2 \times \frac{1 \text{ mol Cl}_2}{71.0 \text{ g Cl}_2} \times \frac{2 \text{ mol S}}{1 \text{ mol Cl}_2} \times \frac{32.1 \text{ g S}}{1 \text{ mol S}} = 13.6 \text{ g S}$$

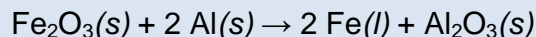
The mole map is a tool we can use to help us to know which ratios to use when solving problems. You use this map much like you would use a road map. You must first find out where you are on the map (your given units) and where you would like to go (your "find" units). The map will then let you know which roads (ratios) to take to get there. Let's see how this works with a couple of example problems.



7.6: Stoichiometry

Examples Using a Mole Map

Example: The thermite reaction is a very exothermic reaction which produces liquid iron, given by the following balanced equation:



If 5.00 g of iron is produced, how much iron(III) oxide was placed in the original container?

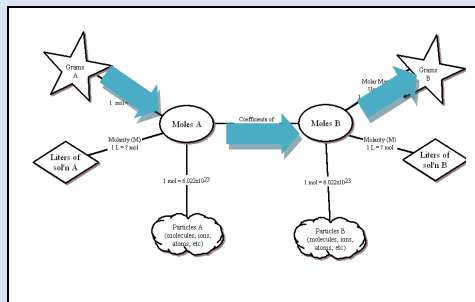
Solution:

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.

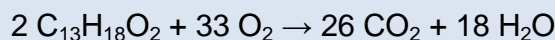
We want to know g Fe_2O_3 . For this problem, “B” stands for Fe_2O_3 . We are heading to grams B.



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:

$$5.00 \text{ g Fe} \cdot \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} \cdot \frac{1 \text{ mol Fe}_2\text{O}_3}{2 \text{ mol Fe}} \cdot \frac{159.7 \text{ g Fe}_2\text{O}_3}{1 \text{ mol Fe}_2\text{O}_3} = 7.17 \text{ g Fe}_2\text{O}_3$$

Example: Ibuprofen is a common painkiller used by many people around the globe. It has the formula $\text{C}_{13}\text{H}_{18}\text{O}_2$. If 200.g of Ibuprofen is combusted how much carbon dioxide is produced? The balanced reaction is:

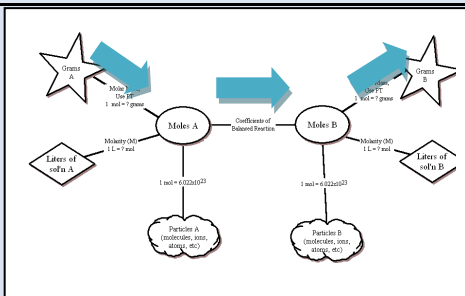


Solution:

Given: 200. g $\text{C}_{13}\text{H}_{18}\text{O}_2$ (g A on the map)

Find: g CO_2 (g B on the map)

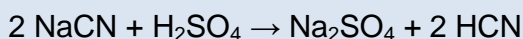
Ratios: The map says we need to use the molar mass of $\text{C}_{13}\text{H}_{18}\text{O}_2$, then the coefficients of the balanced reaction, then the molar mass of CO_2 .



$$200. \text{ g } \text{C}_{13}\text{H}_{18}\text{O}_2 \cdot \frac{1 \text{ mol } \text{C}_{13}\text{H}_{18}\text{O}_2}{206.3 \text{ g } \text{C}_{13}\text{H}_{18}\text{O}_2} \cdot \frac{26 \text{ mol } \text{CO}_2}{2 \text{ mol } \text{C}_{13}\text{H}_{18}\text{O}_2} \cdot \frac{44.01 \text{ g } \text{CO}_2}{1 \text{ mol } \text{CO}_2} = 555 \text{ g } \text{CO}_2$$

7.6: Stoichiometry

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:

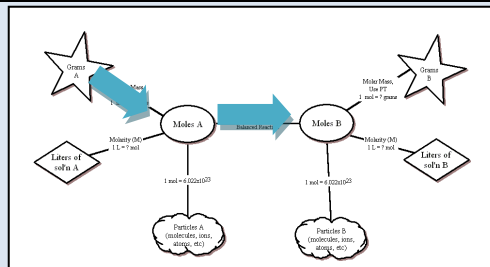


Solution:

Given: 12.5 g HCN (g A on map)

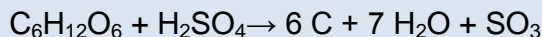
Find: mol H_2SO_4 (mol A on map)

Ratios: The mole map says we need the molar mass of HCN and the coefficients of the balanced reaction.



$$12.5 \text{ g HCN} \cdot \frac{1 \text{ mol HCN}}{27.0 \text{ g HCN}} \cdot \frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol HCN}} = 0.231 \text{ mol H}_2\text{SO}_4$$

Example: How many atoms of carbon would be released from the complete dehydration of 18.0 g of sugar ($\text{C}_6\text{H}_{12}\text{O}_6$) with sulfuric acid? The balanced reaction is:

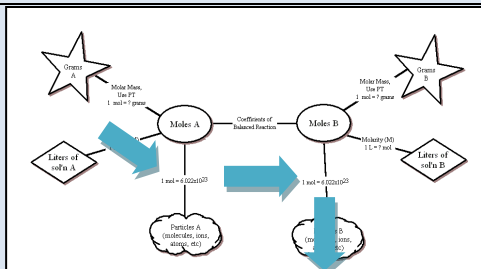


Solution:

Given: 18 g $\text{C}_6\text{H}_{12}\text{O}_6$

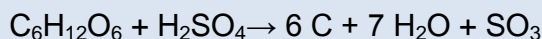
Find: atoms C

Ratios: the mole map says we need the molar mass of the sugar, the balanced reaction, and finally Avogadro's number.



$$18.0 \text{ g C}_6\text{H}_{12}\text{O}_6 \times \frac{1 \text{ mol C}_6\text{H}_{12}\text{O}_6}{180. \text{g C}_6\text{H}_{12}\text{O}_6} \times \frac{6 \text{ mol C}}{1 \text{ mol C}_6\text{H}_{12}\text{O}_6} \times \frac{6.02 \times 10^{23} \text{ atoms C}}{1 \text{ mol C}} = 3.6 \times 10^{23} \text{ atoms C}$$

Example: What mass of carbon is formed when 1.25 g of water are produced?

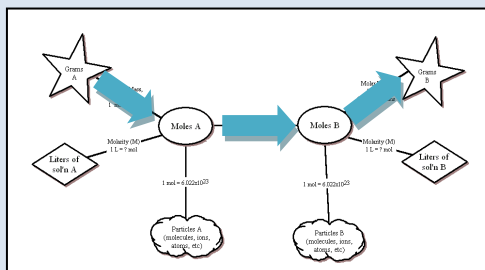


Solution:

Given: 1.25 g H_2O

Find: g C

Ratios: the mole map says we need the molar mass of water, the balanced reaction, and finally the molar mass of carbon.



$$1.25 \text{ g H}_2\text{O} \times \frac{1 \text{ mol H}_2\text{O}}{18.02 \text{ g H}_2\text{O}} \times \frac{6 \text{ mol C}}{7 \text{ mol C}_6\text{H}_{12}\text{O}_6} \times \frac{12.01 \text{ g C}}{1 \text{ mol C}} = 0.71 \text{ g C}$$

7.6: Stoichiometry

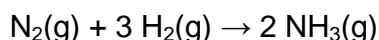
Limiting Reactants

Let us suppose that you are deciding to make some pancakes for a large group of people. The recipe on the box indicates that the following ingredients are needed for each batch of pancakes:

- 1 cup of pancake mix
- $\frac{3}{4}$ cup milk
- 1 egg

You check the pantry and the refrigerator and see that you have the following ingredients available: 8 cups pancake mix, 4 cups milk, and 2 eggs. The question that you must ask is: How many batches of pancakes can I make? The answer is two. Even though you have enough pancake mix and milk oil to make many more batches of pancakes, you are *limited* by the fact that you only have two eggs. As soon as you have made two batches of pancakes, you will be out of eggs and your “reaction” will be complete.

For a chemist, the balanced chemical equation is the recipe that must be followed. Consider the following balanced equation showing the formulation of ammonia:



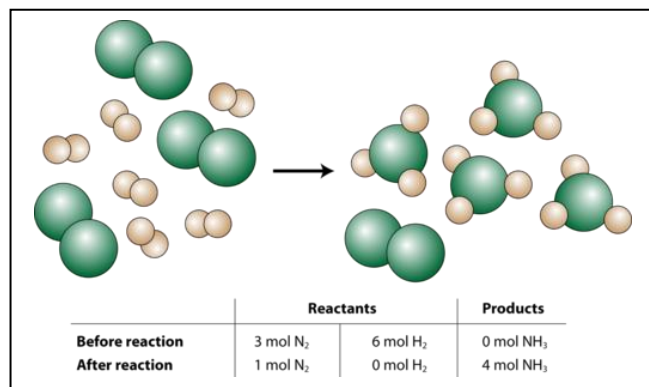
We know that the coefficients of the balanced equation tell us the mole ratio that is required for this reaction to occur. One mole of N_2 will react with three moles of H_2 to form two moles of NH_3 . Now let us suppose that a chemist were to react three moles of N_2 with six moles of H_2 .

So what happened in this reaction? The chemist started with 3 moles of N_2 . You may think of this as being 3 times as much as the “recipe” (the balanced equation) requires since the coefficient for the N_2 is a 1. However, the 6 moles of H_2 that the chemist started with is only two times as much as the “recipe” requires, since the coefficient for the H_2 is a 3 and $3 \times 2 = 6$. So the hydrogen gas will be completely used up while there will be 1 mole of nitrogen gas left over after the reaction is complete. Finally, the reaction will produce 4 moles of NH_3 because that is also two times as much as shown in the balanced equation. The overall reaction that occurred in words:



All the amounts are doubled from the original balanced equation.

The **limiting reactant (or limiting reagent)** is the reactant that determines the amount of product that can be formed in a chemical reaction. The reaction proceeds until the limiting reactant is completely used up. In our example above, the H_2 is the limiting reactant. The **excess reactant (or excess reagent)** is the reactant left over after the reaction is complete. In the above example, the N_2 is the excess reactant.



7.7: Reversible Reactions

Reversible Reactions and Equilibrium

Objectives:

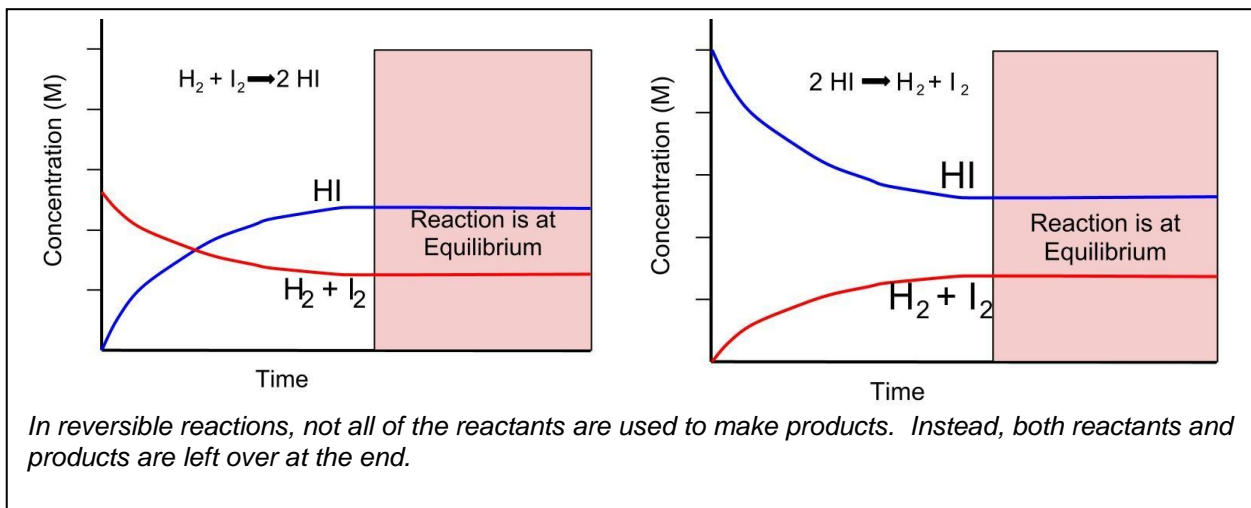
- Describe what is meant by the term "dynamic equilibrium"
- Indicate whether a reaction is at equilibrium

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".



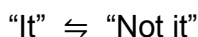
Some reactions, however, are **reversible**, meaning the reaction can go backwards, so that: $A + B \leftarrow C + D$. The direction of the arrow shows that the products, C and D, are reacting to form the reactants, A and B. What if the two reactions, the forward reaction and the reverse reaction, were 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.

Chemists use a double-headed arrow, \rightleftharpoons , 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.

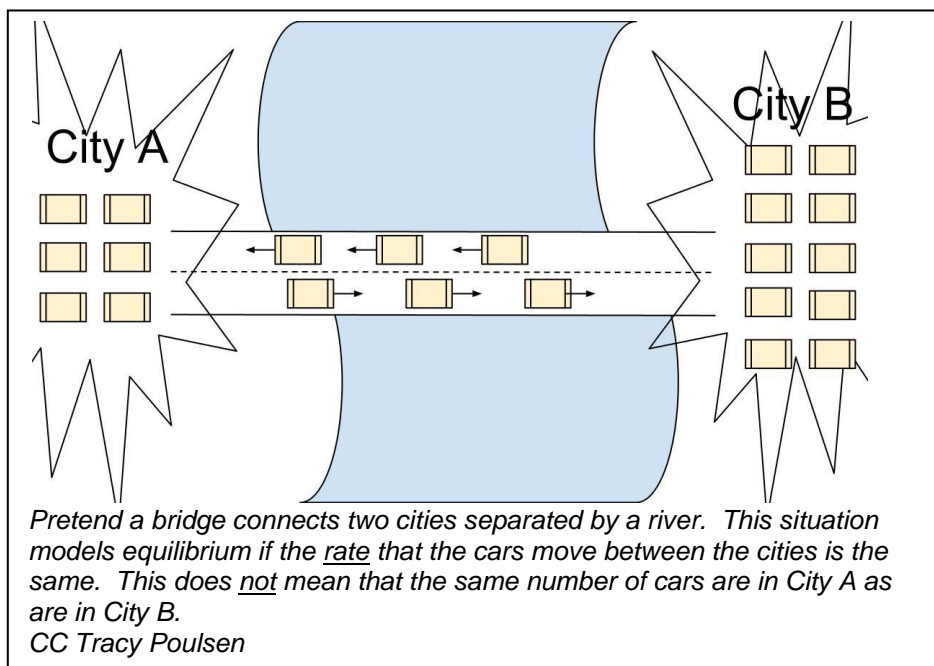
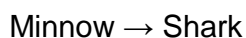


7.7: Reversible Reactions

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:



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:



Equilibrium: A reaction in which the rate of the forward reaction equals the rate of the reverse reaction

Objectives:

- Write equilibrium expressions
- Use equilibrium expressions to calculate the value of K or the concentration of a reactant or product
- Use K to predict the relative amounts of products and reactants at equilibrium

7.8: Equilibrium Constant Expressions

Writing Equilibrium Expressions

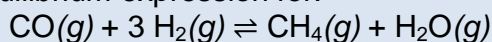
In the previous section, you learned about reactions that can reach a state of equilibrium, in which the concentration of reactants and products aren't changing. If these amounts are changing, we should be able to make a relationship between the amount of product and reactant when a reaction reaches equilibrium.

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:

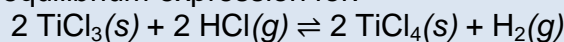


Solution:

$$K = \frac{[\text{CH}_4][\text{H}_2\text{O}]}{[\text{CO}][\text{H}_2]^3}$$

*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:

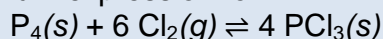


Solution:

$$K = \frac{[\text{H}_2]}{[\text{HCl}]^2}$$

*Note that the solids are left out of the expression completely

Example: Write the equilibrium expression for:



Solution:

$$K = \frac{1}{[\text{Cl}_2]^6}$$

*Note that the only product is a solid, which is left out. That leaves just 1 on top in the numerator

7.8: Equilibrium Constant Expressions

The Equilibrium Constant

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 1×10^{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 1×10^{-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, $\text{SO}_2(g) + \text{NO}_2(g) \rightleftharpoons \text{SO}_3(g) + \text{NO}(g)$ determine the value of K when the equilibrium concentrations are: $[\text{SO}_2]=1.20 \text{ M}$, $[\text{NO}_2]=0.60 \text{ M}$, $[\text{NO}]=1.6 \text{ M}$, and $[\text{SO}_3]=2.2 \text{ M}$.

Solution:

Step 1: Write the equilibrium constant expression:

$$K = \frac{[\text{SO}_3][\text{NO}]}{[\text{SO}_2][\text{NO}_2]}$$

Step 2: Substitute in given values and solve:

$$K = \frac{(2.2)(1.6)}{(1.20)(0.60)} = 4.9$$

Example: Consider the following reaction: $\text{CO}(g) + \text{H}_2\text{O}(g) \rightleftharpoons \text{H}_2(g) + \text{CO}_2(g)$; $K=1.34$

If the $[\text{H}_2\text{O}]=0.100 \text{ M}$, $[\text{H}_2]=0.100 \text{ M}$, and $[\text{CO}_2]=0.100 \text{ M}$ at equilibrium, what is the equilibrium concentration of CO?

Solution:

Step 1: Write the equilibrium constant expression:

$$K = \frac{[\text{H}_2][\text{CO}_2]}{[\text{CO}][\text{H}_2\text{O}]}$$

Step 2: Substitute in given values and solve:

$$1.34 = \frac{(0.100)(0.100)}{[\text{CO}](0.100)}$$

Solving for [CO], we get: $[\text{CO}]=0.0746 \text{ M}$

7.9: Le Châtelier's Principle

Changing a System at Equilibrium

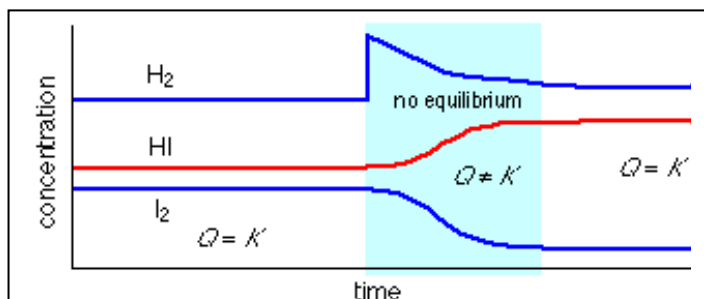
Objectives:

-Predict the effect of changing the concentration or temperature on a system at equilibrium

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.

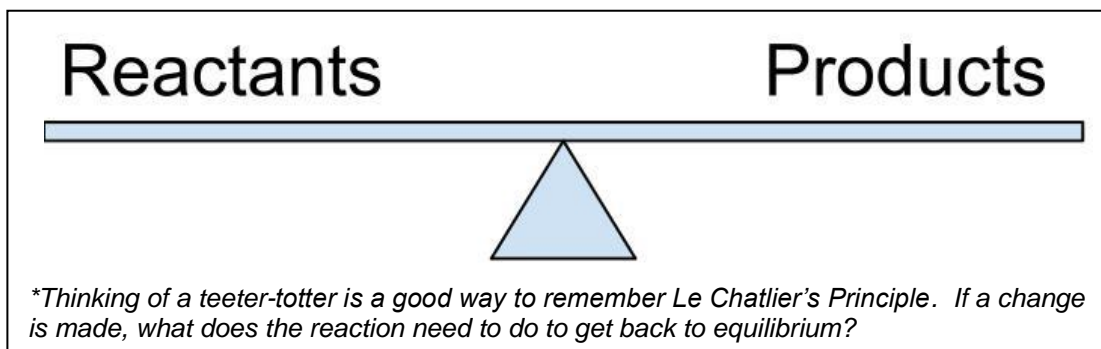
Le Châtelier's Principle



Le Chatlier's Principle is used to predict how a system will re-establish equilibrium after some change has been made to the svstem.

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.

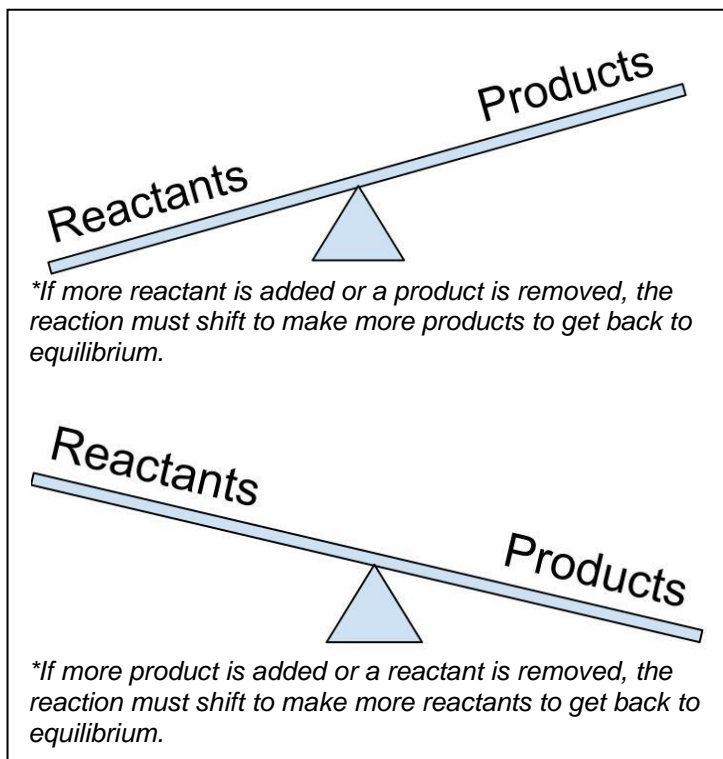


Le Chatlier's Principle: Predicts the effect changing the temperature, pressure, or concentration of a substance on a system in equilibrium

7.9: Le Châtelier's Principle

The Effect of Concentration Changes

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 if 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.

Example: For the reaction: $\text{SiCl}_4(g) + \text{O}_2(g) \rightleftharpoons \text{SiO}_2(s) + 2 \text{Cl}_2(g)$, what would be the effect on the equilibrium system if:

- (a) $[\text{SiCl}_4]$ increases
- (b) $[\text{O}_2]$ increases
- (c) $[\text{Cl}_2]$ increases

Solution:

- (a) $[\text{SiCl}_4]$ increases: The equilibrium would shift to the right
- (b) $[\text{O}_2]$ increases: The equilibrium would shift to the right
- (c) $[\text{Cl}_2]$ increases: The equilibrium would shift left

Example: Here's a reaction at equilibrium. $\text{A}(aq) + \text{B}(aq) \rightleftharpoons \text{C}(aq) + \text{D}(aq)$

- (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?

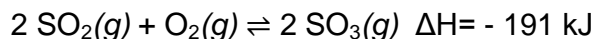
Solution:

- (a) The equilibrium will shift toward the products (forward).
- (b) The equilibrium will shift toward the reactants (backward).

7.9: Le Châtelier's Principle

The Effect of Changing Temperature

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:



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, or that the reaction is **exothermic**. This equation can also be written as:

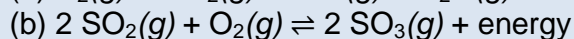
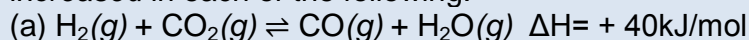


What's important to remember about increasing the temperature of an equilibrium system is energy can be thought as just another product or reactant. In this example, you can clearly see that the heat is a product. Therefore when the temperature of this system is raised, heat is being added and the effect will be the same as increasing any other product. Increasing a product causes the reaction to use up some of the products to make more reactants. 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 $[\text{SO}_3]$ will increase and the $[\text{SO}_2]$ and $[\text{O}_2]$ will decrease.

In **endothermic** reactions, though, heat is a reactant. These reactions would have the opposite effect. If heat is a reactant, adding heat 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.



Solution:

(a) The reaction is endothermic, because ΔH is positive, meaning heat is a reactant. We would write: $\text{heat} + \text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{CO}(g) + \text{H}_2\text{O}(g)$
Increasing temperature for an endothermic reaction, the reactions will shift right producing more products.

(b) The reaction is exothermic, meaning heat is a product. Increasing temperature for an exothermic reaction, the reactions will shift left producing more reactants.

Chapter 7 Summary

7.1: Physical and Chemical Change

- 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 through forming and breaking chemical bonds.
- 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, and formation of a precipitate.

7.2: Reaction Rate

- 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.
- 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.
- With an increase in temperature, there is an increase in energy in a collision and therefore there will be an increase in the reaction rate. A decrease in temperature would have the opposite effect.
- 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.

7.3: Writing Reactions

- Chemical equations have reactants on the left, an arrow that is read as "yields," and the products on the right.
- When writing reactions from words, you must first write the correct formulas for all elements and compounds in the reaction, making sure that the total charge of all compounds is zero using appropriate subscripts.
- Some elements are diatomic, meaning they go around in pairs.

7.4: Balancing Reactions

- 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.

Chapter 7 Summary

7.5: Types of Reactions

- By classifying reactions, chemists are able to predict the products of many reactions.
- Many reactions are one of five types of reactions:

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.

7.6: Stoichiometry

- 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.
- A mole map is a useful tool for determining which ratios are needed to solve a problem involving a balanced chemical equation.

7.7: Equilibrium

- Some reactions don't go to completion. Some reactions are reversible and will reach chemical equilibrium (reactants \rightleftharpoons products).
- A dynamic equilibrium is a state where the rate of the forward reaction is equal to the rate of the reverse reaction.
- At equilibrium, the amounts of products and reactants are not equal, but they are constant.

7.8: Equilibrium Expressions

- 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.

7.9: Le Chatelier's Principle

- Increasing the concentration of a reactant causes the equilibrium to shift to the right producing more products.

Chapter 7 Summary

- 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 Readings / Supplemental Links

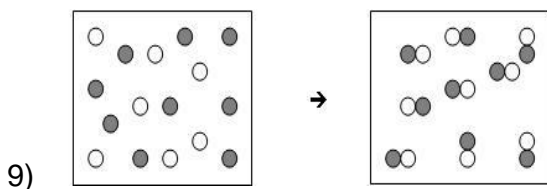
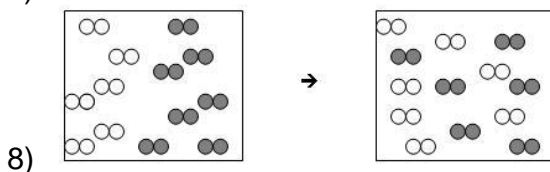
- Balancing Equations Tutorial:
http://www.mpcfaculty.net/mark_bishop/balancing_equations_tutorial.htm
- Balancing Equations Tutorial:
<http://www.wfu.edu/~ylwong/balanceeq/balanceeq.html>
- Equilibrium Animation / Applet:
<http://chemconnections.org/Java/equilibrium/>
- Tutorial: Le Chatlier's Principle:
<http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/lechv17.swf>

7.1: Physical and Chemical Changes Review Questions

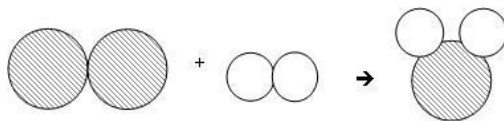
#1-10: Label each of the following as a physical or chemical change.

Justify your answer in each case.

- 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.
- 7) A cake is baked.



Chapter 7 Summary



10)

7.2: Reaction Rate Review Questions

- 11) According to the collision theory, what must happen in order for a reaction to be successful? (Select all that apply.)
- particles must collide
 - particles must have proper geometric orientation
 - particles must have collisions with enough energy
- 12) What would happen in a collision between two particles if particles did not have enough energy or had the incorrect orientation?
- the particles would rebound and there would be no reaction
 - the particles would keep bouncing off each other until they eventually react, therefore the rate would be slow
 - the particles would still collide but the byproducts would form
 - the temperature of the reaction vessel would increase
- 13) Why does higher temperature increase the reaction rate?
- more of the reacting molecules will have higher kinetic energy
 - increasing the temperature causes the reactant molecules to heat up
 - the activation energy will decrease
- 14) When the temperature is increased, what does not change?
- number of collisions
 - activation energy
 - number of successful collisions
 - all of the above change
- 15) Why is the increase in concentration directly proportional to the rate of the reaction?
- The kinetic energy increases.
 - The activation energy increases.
 - The number of successful collisions increases.
 - All of the above.
- 16) Why, using the collision theory, do wood shavings burn more quickly than a log?

#17-19: Limestone (calcium carbonate) reacts with hydrochloric acid in an irreversible reaction, to form carbon dioxide and water as described by the following equation:



What is the effect on the rate if:

- 17) The temperature is lowered?
- 18) Limestone chips are used instead of a block of limestone?
- 19) A more dilute solution of HCl is used?

Chapter 7 Summary

7.3: Writing Chemical Reactions 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.

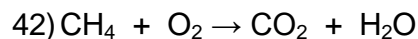
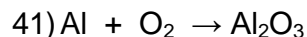
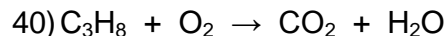
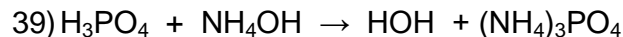
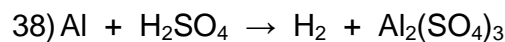
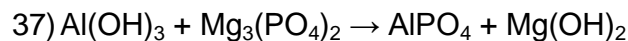
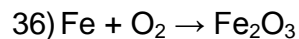
- 20) Solid calcium metal is placed in liquid water to produce aqueous calcium hydroxide and hydrogen gas.
- 21) Gaseous sodium hydroxide is mixed with gaseous chlorine to produce aqueous solutions of sodium chloride and sodium hypochlorite plus liquid water.
- 22) Iron reacts with sulfur when heated to form iron(II) sulfide.
- 23) When aluminum is added to sulfuric acid (H_2SO_4), the solution reacts to form hydrogen gas and aluminum sulfate.
- 24) When aluminum is mixed with iron(III) oxide, they react to produce aluminum oxide and iron.
- 25) Fluorine is mixed with sodium hydroxide to form sodium fluoride, oxygen, and water.
- 26) 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 Review Questions

Balance the following equations.

- 27) $\text{Cu} + \text{O}_2 \rightarrow \text{CuO}$
- 28) $\text{H}_2\text{O} \rightarrow \text{H}_2 + \text{O}_2$
- 29) $\text{Fe} + \text{H}_2\text{O} \rightarrow \text{H}_2 + \text{Fe}_2\text{O}_3$
- 30) $\text{NaCl} \rightarrow \text{Na} + \text{Cl}_2$
- 31) $\text{AsCl}_3 + \text{H}_2\text{S} \rightarrow \text{As}_2\text{S}_3 + \text{HCl}$
- 32) $\text{CaCO}_3 \rightarrow \text{CaO} + \text{CO}_2$
- 33) $\text{H}_2\text{S} + \text{KOH} \rightarrow \text{HOH} + \text{K}_2\text{S}$
- 34) $\text{XeF}_6 + \text{H}_2\text{O} \rightarrow \text{XeO}_3 + \text{HF}$
- 35) $\text{Cu} + \text{AgNO}_3 \rightarrow \text{Ag} + \text{Cu}(\text{NO}_3)_2$

Chapter 7 Summary



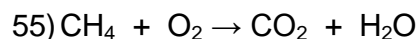
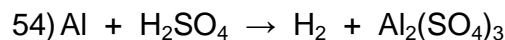
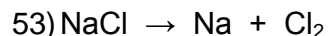
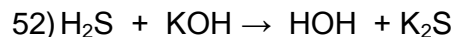
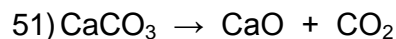
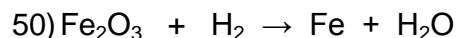
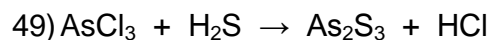
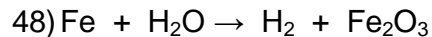
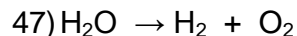
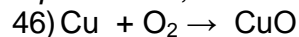
43) When the following equation is balanced, what is the coefficient found in front of the O_2 ? $\text{P}_4 + \text{O}_2 + \text{H}_2\text{O} \rightarrow \text{H}_3\text{PO}_4$

44) When properly balanced, what is the **sum** of all the coefficients in the following chemical equation? $\text{SF}_4 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3 + \text{HF}$

45) 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 Review Questions

#46-55: Classify each type of reaction as synthesis, decomposition, single replacement, double replacement or combustion.

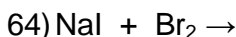
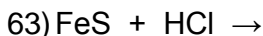
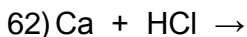
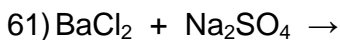
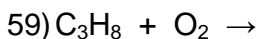
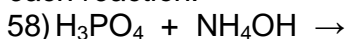


Chapter 7 Summary

56) When dodecane, $C_{10}H_{22}$, burns in excess oxygen, what will be the products?

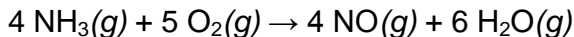
57) When iron rods are placed in liquid water, a reaction occurs. Hydrogen gas evolves from the container and iron(III) oxide forms onto the iron rod. Classify the type of reaction and write a balanced chemical equation for the reaction.

#58-64: Classify each of the following reactions and predict products for each reaction.

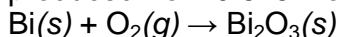


7.6: Stoichiometry Review Questions

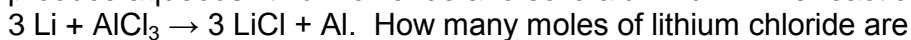
65) Given the reaction between ammonia and oxygen to produce nitrogen monoxide, how many moles of water vapor can be produced from 2 mol of ammonia? The balanced reaction is:



66) When properly balanced, how many moles of bismuth(III) oxide can be produced from 0.625 mol of bismuth? The unbalanced reaction is:



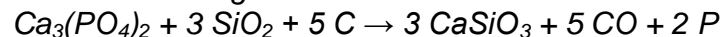
67) Solid lithium reacts with an aqueous solution of aluminum chloride to produce aqueous lithium chloride and solid aluminum. The reaction is:



How many moles of lithium chloride are formed if 5.0 mol aluminum were produced?

Chapter 7 Summary

#68-69: For the given balanced reaction:



68) How many moles of silicon dioxide are required to react with 2.0 mol of carbon?

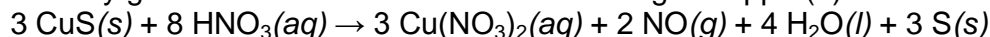
69) How many moles of calcium phosphate are required to produce 0.5 mol of calcium silicate?

#70-71: For the given balanced reaction, $4 \text{FeS} + 7 \text{O}_2 \rightarrow 2 \text{Fe}_2\text{O}_3 + 4 \text{SO}_2$

70) How many moles of iron(III) oxide are produced from 1.27 mol of oxygen?

71) How many moles of iron(II) sulfide are required to produce 3.28 mol of sulfur dioxide?

72) 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?



73) When properly balanced, what mass of iodine was needed to produce 2.5 g of sodium iodide in the equation below? $\text{I}_2(aq) + \text{Na}_2\text{S}_2\text{O}_3(aq) \rightarrow \text{Na}_2\text{S}_4\text{O}_6(aq) + \text{NaI}(aq)$

74) Determine the mass of lithium hydroxide produced when 0.38 grams of lithium nitride reacts with water according to the following equation: $\text{Li}_3\text{N} + 3\text{H}_2\text{O} \rightarrow \text{NH}_3 + 3\text{LiOH}$

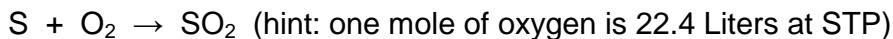
75) If 3.01×10^{23} formulas of cesium hydroxide are produced according to this reaction:



76) How many liters of oxygen are necessary for the combustion of 425 g of sulfur, assuming that the reaction occurs at STP? The balanced

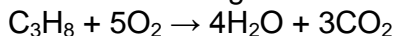
Chapter 7 Summary

reaction is:



77) If I have 2.0 grams of carbon monoxide, how many molecules of carbon monoxide are there?

78) What mass of oxygen is needed to burn 3.5 g of propane (C_3H_8) is burned according to the following equation:



79) How many grams of water are produced if 5 moles of oxygen react according to the following reaction? $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$

7.7: Reversible Reactions Review

Questions

80) 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?

81) The word "equilibrium" comes from the word "equal". What does the term equal mean in this definition?

#80	
Time (min)	$[\text{HC}_2\text{H}_3\text{O}_2]$ 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

#82-86: Indicate whether each of the following statements is true or false for a system in equilibrium.

82) The amount of products is equal to the amount of reactants.

83) The amount of product is not changing.

84) The amount of reactant is not changing.

85) Particles (atoms/molecules) are not reacting.

86) The rate of the forward reaction is equal to the rate of the reverse reaction.

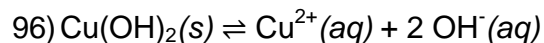
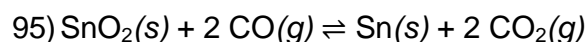
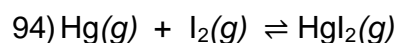
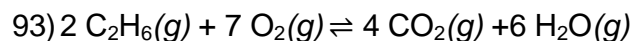
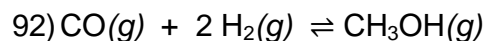
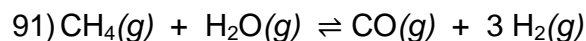
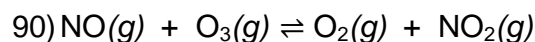
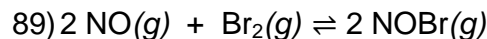
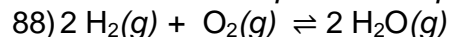
7.8: Equilibrium Constant Expressions Review

Questions

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87) Which phases of substances are not included in the equilibrium expression?

#88-96: Write an equilibrium expression for each reaction:



97) What does a large value for K imply?

98) Consider the following equilibrium system: $2 \text{NO}(g) + \text{Cl}_2(g) \rightleftharpoons 2 \text{NOCl}(g)$. At a certain temperature, the equilibrium concentrations are as follows: $[\text{NO}] = 0.184 \text{ M}$, $[\text{Cl}_2] = 0.165 \text{ M}$, $[\text{NOCl}] = 0.060 \text{ M}$. What is the equilibrium constant for this reaction?

99) For the reaction: $\text{MgCl}_2(s) + \frac{1}{2} \text{O}_2(g) \rightleftharpoons 2 \text{MgO}(s) + \text{Cl}_2(g)$. The equilibrium constant was found to be 3.86 at a certain temperature. If $[\text{O}_2] = 0.560 \text{ M}$ at equilibrium, what is the concentration of $\text{Cl}_2(g)$?

100) Consider the equilibrium: $\text{CO}(g) + \text{H}_2\text{O}(g) \rightleftharpoons \text{H}_2(g) + \text{CO}_2(g)$.
a) Write an equilibrium expression for this reaction.

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- b) If $[\text{CO}] = 0.200\text{M}$, $[\text{H}_2\text{O}] = 0.500\text{M}$, $[\text{H}_2] = 0.32\text{M}$ and $[\text{CO}_2] = 0.42\text{M}$, find K .
- 101) Hydrogen sulfide decomposes according to the equation: $2\text{H}_2\text{S}(g) \rightleftharpoons 2\text{H}_2(g) + \text{S}_2(g)$.
- a) Write an equilibrium expression for this reaction.
- b) At equilibrium, the concentrations of each gas are as follows: $[\text{H}_2\text{S}] = 7.06 \times 10^{-3}\text{M}$, $[\text{H}_2] = 2.22 \times 10^{-3}\text{M}$ and $[\text{S}_2] = 1.11 \times 10^{-3}\text{M}$. What is K_{eq} ?
- 102) Given the following system in equilibrium: $2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)$
- 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 $[\text{SO}_2] = 0.0500\text{M}$ and $[\text{O}_2] = 0.0500\text{M}$, what is the concentration of SO_3 at equilibrium?

7.9: Le Chatelier's Principle Review Questions

#103-105: For the reaction: $\text{N}_2\text{O}_5(s) \rightleftharpoons \text{NO}_2(g) + \text{O}_2(g)$, what would be the effect on the equilibrium if:

103) $[\text{NO}_2]$ decreases

104) $[\text{NO}_2]$ increases

105) $[\text{O}_2]$ increases

#106-108: For the reaction: $\text{C}(s) + \text{H}_2\text{O}(g) \rightleftharpoons \text{CO}(g) + \text{H}_2(g)$, what would be the effect on the equilibrium system if:

106) $[\text{H}_2\text{O}]$ increases

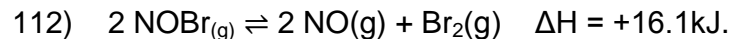
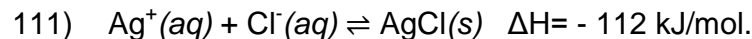
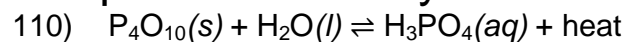
107) $[\text{CO}]$ increases

108) $[\text{H}_2]$ decreases

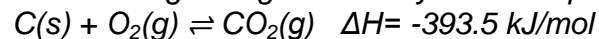
#109-112: Predict the effect on the equilibrium position if the temperature is increased in each of the following.

109) $\text{H}_2(g) + \text{I}_2(g) \rightleftharpoons 2\text{HI}(g)$ $\Delta H = + 51.9\text{ kJ}$

Chapter 7 Summary



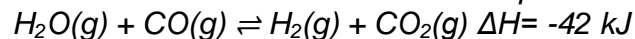
#113-114: In the following reaction, what would be the effect of each of the following changes to the system at equilibrium?



113) increase O_2

114) increase the temperature

#115-118: Predict the effect on the equilibrium:



when each of the following changes are made to the equilibrium system.

115) Temperature is increased

116) $[\text{CO}_2]$ decreases

117) $[\text{H}_2\text{O}]$ increases

118) $[\text{H}_2]$ decreases

#119-120: Predict the effect on the chemical equilibrium

$2 \text{SO}_3(\text{g}) + \text{heat} \rightleftharpoons 2 \text{SO}_2(\text{g}) + \text{O}_2(\text{g})$, when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?

119) Temperature is increased

120) $[\text{O}_2]$ decreases

#121-122: Predict the effect on the chemical equilibrium:

$\text{N}_2\text{O}_4(\text{g}) + \text{heat} \rightleftharpoons 2 \text{NO}_2(\text{g})$, when each of the following changes are made to the equilibrium system. What will the effect be on the amount of product produced?

121) Temperature is decreased

122) $[\text{N}_2\text{O}_4]$ decreases

Chapter 8: Acids & Bases



What made the pits in this gargoyle?

This gargoyle, on Notre Dame Cathedral in Paris, has pits and rounded edges, which are the results of acid rain. Acid rain damages statues and architecture in developed nations.

8.1: Classifying Acids and Bases

Properties of Acids

Objectives:

-Identify acids and bases according to their properties

-Distinguish between Arrhenius acids, Arrhenius bases, Bronsted-Lowry acids and Bronsted-Lowry bases.

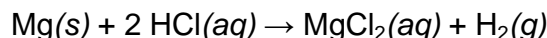
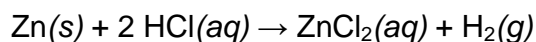
-Give the conjugate acid or base of a Bronsted-Lowry base or acid.

Acids are a special group of compounds with a set of common properties. This helps to distinguish them from other compounds. 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.



Lemon juice is an acid, and all acids taste sour.

Acids react with many metals to produce hydrogen gas. For some examples, look at the reactions below:



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 ($\text{HC}_2\text{H}_3\text{O}_2$) 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.



Both nitric acid and phosphoric acid are used to make fertilizer.



Hydrochloric acid is used to clean swimming pools, bricks, and concrete.



Sulfuric acid is an important component of car batteries.

8.1: Classifying Acids and Bases

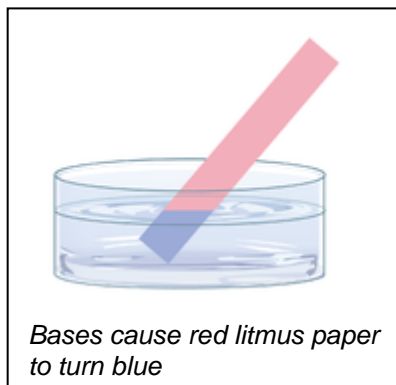
Properties of Bases

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.

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.



Many cleaning products contain bases such as sodium hydroxide.



Concrete contains the base calcium hydroxide.



Many deodorants contain the base aluminum hydroxide.

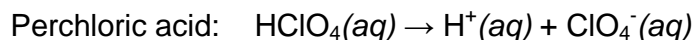
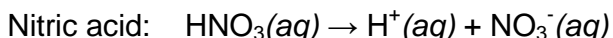
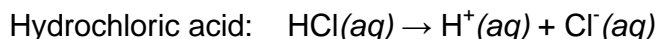
8.1: Classifying Acids and Bases

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

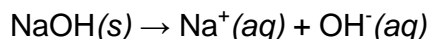
Take a look at all of the following chemical equations. What do you notice about them? What is common for each of the equations below?



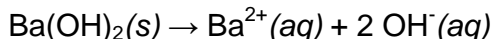
One of the distinguishable features about acids is the fact that acids produce H^+ ions in solution. If you notice in all of the above chemical equations, all of the compounds dissociated to produce H^+ ions. This is the one main, distinguishable characteristic of acids and the basis for the Arrhenius 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:



Barium hydroxide produces a similar reaction when dissociating in water:



The production of OH^- ions is the definition of bases according to the Arrhenius.



Svante Arrhenius
(image obtained from:
<http://upload.wikimedia.org/wikipedia/commons/6/6c/Arrhenius2.jpg>; public domain)

Acid (Arrhenius): A substance that produces H^+ ions in solution

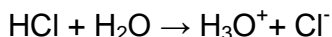
Base (Arrhenius): A substance that produced OH^- ions in solution

8.1: Classifying Acids and Bases

Brønsted-Lowry Acids & Bases

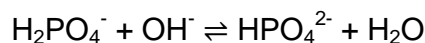
Arrhenius provided chemistry with the first definition of acids and bases, but like a lot of scientific ideas, these definitions have been refined over time. Two chemists, named Brønsted and Lowry, working on similar experiments as Arrhenius, derived a more generalized definition for acids and bases. As the Brønsted-Lowry definition unfolded, the number of acids and bases that were able to fit into each category increased.

A Brønsted-Lowry **acid** is a substance that is a proton (H^+) donor and a Brønsted-Lowry **base** is a proton (H^+) accepter. This theory would treat hydrochloric acid added to water as follows:



The acid is losing (donating) H^+ to the water. Water in this case is acting as a Brønsted-Lowry base, by accepting H^+ to become H_3O^+ . The ion H_3O^+ is called hydronium. For most purposes, H^+ is equivalent to H_3O^+ .

Let's look at another example of Brønsted-Lowry acids & bases.



In this case, the H_2PO_4^- is donating H^+ to OH^- . H_2PO_4^- is the acid, and OH^- is the base.

There is one more aspect of the Brønsted-Lowry theory that was a significant breakthrough to acid-base chemistry. Brønsted and Lowry said that in acid-base reactions, there are actually pairs of acids and bases in the reaction itself. According to Brønsted-Lowry, for every acid there is a conjugate base associated with that acid. The conjugate base is the result of the acid losing (or donating) a proton. Therefore, if you look at the figure below, you can see on the left, the acid and on the right the conjugate base.

Look at the previous acid/base reaction. Note that it is reversible. What happens if we treat this reaction backwards? The H_2O becomes the acid, donating H^+ to the base, HPO_4^{2-} . The Brønsted-Lowry definition includes terms for these reverse components. A **conjugate base** is what remains of an acid after it has donated the H^+ . In the reverse reaction, this substance will act as a base. A **conjugate acid** is what remains after the base has accepted H^+ . In the reverse reaction, this substance will act as an acid. In the reaction above, HPO_4^{2-} is the conjugate base and H_2O is the conjugate acid.

<u>Acid</u>	<u>Conjugate Base</u>
HCl	Cl^-
HBr	Br^-
HNO_3	NO_3^-
$\text{HC}_2\text{H}_3\text{O}_2$	$\text{C}_2\text{H}_3\text{O}_2^-$

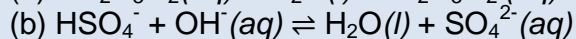
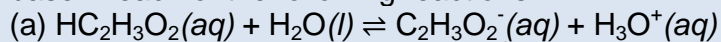
<u>Base</u>	<u>Conjugate Acid</u>
NH_3	NH_4^+
OH^-	HOH
H_2O	H_3O^+
CO_3^{2-}	HCO_3^-

Notice that the difference between the acid and its conjugate base is simply a difference of a proton. The conjugate base has one less proton (or H^+). For every base in the acid-base reaction, there must be a

8.1: Classifying Acids and Bases

corresponding conjugate acid. The conjugate acid is the result of the base gaining (or accepting) the proton.

Example: Identify the Brønsted Lowry acid, base, conjugate acid, and conjugate base in each of the following reactions.



Solution:

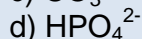
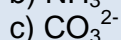
(a) $\text{HC}_2\text{H}_3\text{O}_2$ is the acid. It donates H^+ to H_2O and becomes $\text{C}_2\text{H}_3\text{O}_2^-$. $\text{C}_2\text{H}_3\text{O}_2^-$ is therefore the conjugate base. In the reverse reaction, $\text{C}_2\text{H}_3\text{O}_2^-$ will accept H^+ to become $\text{HC}_2\text{H}_3\text{O}_2$.

H_2O is the base, accepting H^+ to become H_3O^+ . H_3O^+ is the conjugate acid of this reaction.

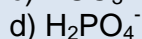
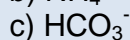
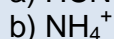
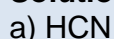
(b) HSO_4^- is the acid, donating H^+ to the base, OH^- .

SO_4^{2-} is the conjugate base and H_2O is the conjugate acid.

Example: What is the conjugate acid of each of the following:

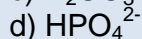
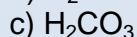
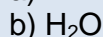


Solution:

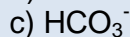


*Note that each conjugate acid has one more H atom as well as an increase in charge by +1.

Example: What is the conjugate base of each of the following:



Solution:



*Note that each conjugate base has one fewer H atom as well as a decrease in charge by 1.

Acid (Brønsted-Lowry): A substance that donates protons (or H^+ ions)

Base (Brønsted-Lowry): A substance that accepts protons (or H^+ ions)

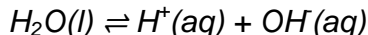
Conjugate Acid: The substance produced after a base has accepted an H^+ ion

Conjugate Base: The substance produced after an acid has donated an H^+ ion

8.2: The pH Scale

The 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 and any solution that contains water. Water self-ionizes according to the following reaction:



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, $[H^+] = 1 \times 10^{-7}$ M and $[OH^-] = 1 \times 10^{-7}$ M. If you substitute these values into the equilibrium expression, you find that $K_w = 1 \times 10^{-14}$. 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:

$$K_w = 1 \times 10^{-14} = [H^+][OH^-] \text{ @ } 25^\circ\text{C}$$

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 $[H^+] > [OH^-]$ or $[H^+] > 1 \times 10^{-7}$ the solution is acidic.
- If $[H^+] < [OH^-]$ or $[H^+] < 1 \times 10^{-7}$ the solution is basic.

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 1×10^{-4} M. What would $[OH^-]$ be?

Solution: substitute what we know into the equilibrium expression:

$$K_w = 1 \times 10^{-14} = [H^+][OH^-]$$

$$1 \times 10^{-14} = [1 \times 10^{-4}][OH^-]$$

To isolate $[OH^-]$, divide by sides by 1×10^{-4} .

$$\text{This leaves, } [OH^-] = 1 \times 10^{-10} \text{ M}$$

Note that because $[H^+] > [OH^-]$, the solution must be acidic.

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 1.26×10^{-11} ? What is $[H^+]$ and is the solution acidic, basic, or neutral?

Solution:

$$K_w = 1 \times 10^{-14} = [H^+][OH^-]$$

$$1 \times 10^{-14} = [H^+][1.26 \times 10^{-11}]$$

To isolate $[H^+]$, divide by sides by 1.26×10^{-11} .

$$\text{This leaves, } [H^+] = 7.94 \times 10^{-4} \text{ M}$$

Also, the solution must be acidic because $[H^+] > [OH^-]$

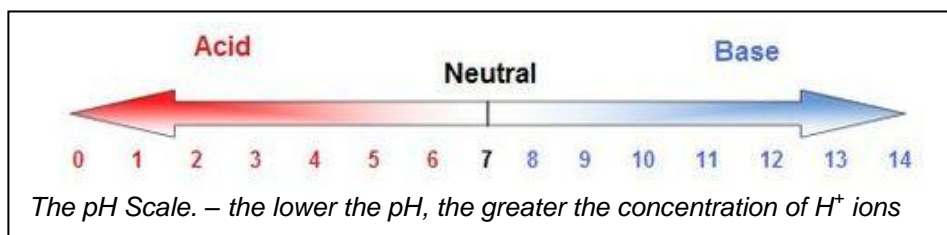
Objectives:

- Calculate the concentration of H^+ and OH^- ions
- Calculate the pH of solution
- Compare the relative concentration of H^+ ions of solutions with different pH values

8.2: The pH Scale

Calculating pH

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 1.0×10^{-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:

$$pH = -\log (0.01)$$
$$pH = 2$$

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} .

$$\text{This leaves, } [H^+] = 1.39 \times 10^{-11} \text{ M}$$

We can now find the pH

$$pH = -\log (1.39 \times 10^{-11})$$
$$pH = 10.9$$

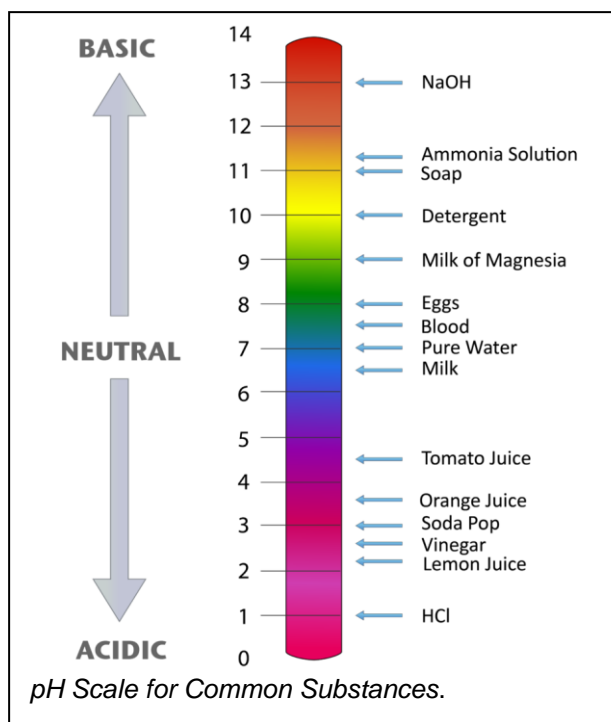
8.2: The pH Scale

Understanding the pH Scale

The pH scale developed by Sørensen is a logarithmic scale, which means that a difference of 1 in 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=1 is 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.

Have you ever cut an onion and had your eyes water up? 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.



Example: Compare lemon juice (pH=2.5) to milk (pH=6.5). Answer each of the following:

- Label each as acidic, basic, or neutral
- Which has a higher concentration of H^+ ions?
- How many times more H^+ does that solution have?

Solution:

- 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)
- The lower the pH, the higher the concentration of H^+ ions. Therefore, lemon juice has more H^+ .
- 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 $10 \times 10 \times 10 \times 10$ or 10,000 times more H^+ ions than milk.

pH Scale: a logarithmic scale used to compare the acidity of various solutions; lower numbers on the scale indicate a greater concentration of H^+ ions in solution

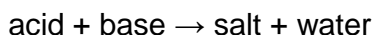
8.3: Reactions Between Acids & Bases

Acid-Base Neutralization Reactions

Objectives:

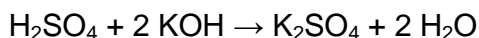
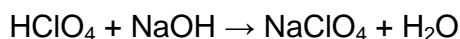
-Write balanced acid/base reactions

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.



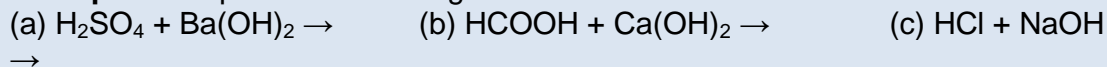
Acids are a combination of hydrogen ions (H^+) and an anion. Examples include HCl , HNO_3 , and $\text{HC}_2\text{H}_3\text{O}_2$. Bases can be a combination of metal cations and hydroxide ions, OH^- . Examples include NaOH , KOH , and $\text{Mg}(\text{OH})_2$. According to the Arrhenius definitions 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?



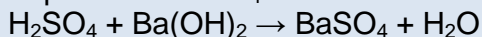
No matter what the acid or the base may be, the products of this type of reaction will always be a salt and water. The H^+ ion from the acid will neutralize the OH^- ion from the base to form water. The other product is a salt formed when the cation of the base combines with the anion of the acid. Remember, the total charge on the salt MUST be zero. You must have the correct number of cations and anions to cancel out the charges.

Example: Complete the following neutralization reactions.

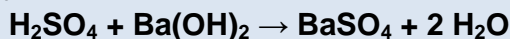


Solution:

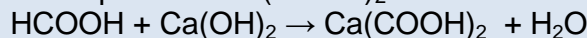
(a) The H^+ in H_2SO_4 will combine with the OH^- part of $\text{Ba}(\text{OH})_2$ to make water (H_2O). The salt produced is what is formed when Ba^{2+} (the cation from the base) combines with SO_4^{2-} (the anion from the acid). These have charges of +2 and -2, so the formula for this compound is BaSO_4 .



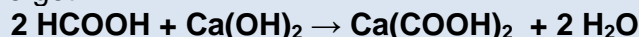
After balancing, we get:



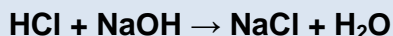
(b) The H^+ in HCOOH will combine with the OH^- part of $\text{Ca}(\text{OH})_2$ to make water (H_2O). The salt produced is what is formed when Ca^{2+} (the cation from the base) combines with COOH^- the anion from the acid). These have charges of +2 and -1, so the formula for this compound is $\text{Ca}(\text{COOH})_2$.



After balancing, we get:



(c) The H^+ in HCl will combine with the OH^- part of NaOH to make water (H_2O). 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 .



The reaction is already balanced, so we are done.

Neutralization Reaction: A reaction between an acid and base, frequently producing water and a salt

8.4: Titrations

Indicators

An **indicator** is a substance that changes color at a specific pH. 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 ($\text{pH} > 7$), if the blue litmus turns red the solution is acidic ($\text{pH} < 7$).

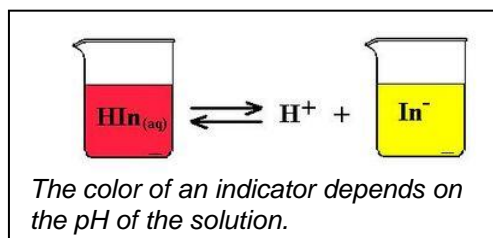


Hydrangeas are also a natural indicator. The petals will change colors based on the pH of the soil.

The juice from red cabbage can be used to prepare an indicator paper. It contains the chemical anthocyanin, 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

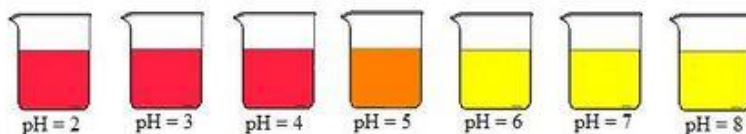
the soil. If you are travelling around and see a hydrangea plant with blue flowers, the soil is acidic, the creamy white flowers indicate the soil is neutral, and the pink flowers mean the soil is basic.

There are two requirements for a substance to function as an acid-base indicator; 1) the substance must form an equilibrium affected by the addition of H^+ ions and 2) the two forms of the compound (acid and conjugate base) must have different colors.



The color of an indicator depends on the pH of the solution.

For the example above, HIn is red and In^- is 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 $\text{pH}=5$, 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.

Objectives:

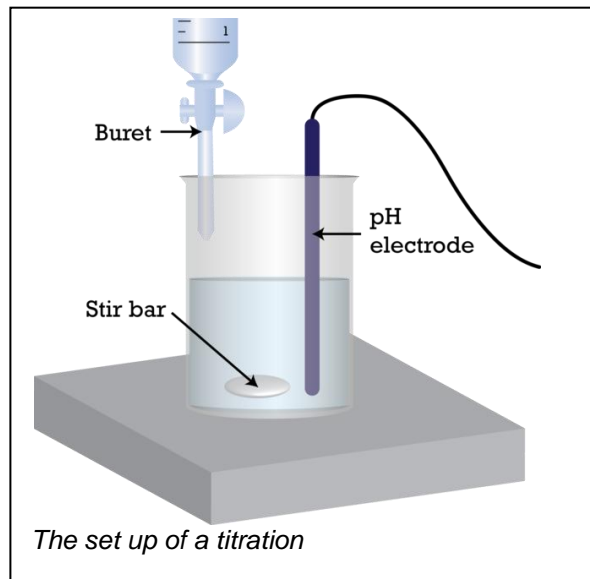
- Describe the purpose of an indicator
- Describe the purpose and set-up of an acid/base titration
- Calculate the concentration of an acid or base given data from a titration

Indicator: a substance whose color changes depending on the pH

8.4: Titrations

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).

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.

Titration: A lab procedure used to find the concentration of a solution by reacting it with another solution

Equivalence Point: The point in a titration in which equals moles of acid and base have reacted

8.4: Titrations

The Mathematics of Titration

For most acid/base neutralization reactions, the balanced reactions show one molecule of acid reacting with one molecule of the base. Whenever we have a 1:1 ratio of acid (H^+ ions) : base (OH^- ions), we find calculate the amounts and concentrations of each substance needed to reach the equivalence point by the following equation:

$$(M_a)(V_a)=(M_b)(V_b)$$

Where

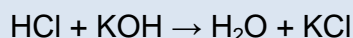
- M_a is the molarity of the acid
- V_a is the volume of the acid
- M_b is the molarity of the base
- V_b is the volume of the base

In a titration we are looking for the equivalence point, the point where the number of acid molecules is equal to the number of base molecules. This equation works, because the left side calculates the number of moles of acid and the right side calculates the number of moles of 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.



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.

$$\begin{aligned}M_a &= 0.125 \text{ M} \\V_a &= 10.0 \text{ mL} \\M_b &= 0.100 \text{ M} \\V_b &=?\end{aligned}$$

$$\begin{aligned}(M_a)(V_a) &= (M_b)(V_b) \\(0.125 \text{ M})(10.0 \text{ mL}) &= (0.100 \text{ M})(V_b) \\V_b &= \mathbf{12.5 \text{ mL}}\end{aligned}$$

Therefore, for this weak acid-strong base titration, the volume of base required for the titration is 12.5 mL.

Example: What is the concentration of an acid if it takes 17.52 mL of a 0.1025 M solution of base to neutralize a 25.00 mL sample of the acid? Assume a 1:1 ratio of the acid and base in the balanced reaction.

Solution:

Using the information in the problem, we know: $M_a=?$; $V_a=25.00 \text{ mL}$; $M_b=0.1025 \text{ M}$; and $V_b=17.52 \text{ mL}$.

$$\begin{aligned}(M_a)(V_a) &= (M_b)(V_b) \\(M_a)(25.00 \text{ mL}) &= (0.1025 \text{ M})(17.52 \text{ mL}) \\M_a &= \mathbf{0.07183 \text{ M}}\end{aligned}$$

Chapter 8 Summary

8.1: Classifying Acids and Bases

- Acids turn blue litmus paper red, taste sour, and react with metals to produce hydrogen gases.
- Common acids include vinegar ($\text{HC}_2\text{H}_3\text{O}_2$), 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 ($\text{Mg}(\text{OH})_2$) 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.
- Brønsted-Lowry acid is a substance that is a proton (H^+) donor and a Brønsted-Lowry base is a proton (H^+) acceptor.

8.2: The pH Scale

- Water ionizes slightly according to the equation $\text{H}_2\text{O}(l) \rightleftharpoons \text{H}^+(aq) + \text{OH}^-(aq)$
- The equilibrium constant for the dissociation of water is:
 $K_w = 1 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$
- $\text{pH} = -\log [\text{H}^+]$
- As the pH of a solution decreases, its H^+ concentration increases
- The pH scale is a logarithmic scale, meaning that each change of one on the pH scale changes the concentration of H^+ ions by a factor of ten.

8.3: Reactions Between Acids and Bases

- A neutralization reaction between an acid and a base will produce a salt and water.

8.4: Titrations

- 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.

8.1: Classifying Acids and Bases Review Questions

#1-6: 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

Chapter 8 Summary

- 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?

#8-10: Give the conjugate acid of each of the following bases.

- 8) PO_4^{3-}
- 9) HCO_3^-
- 10) NH_3

#11-13: Give the conjugate base of each of the following acids.

- 11) HNO_2
- 12) HI
- 13) H_2PO_4^-

8.2: The pH Scale Review Questions

- 14) In saturated limewater, $[\text{H}^+] = 3.98 \times 10^{-13}$ M.
 - a) Find $[\text{OH}^-]$
 - b) What is the pH?
 - c) Is the solution acidic, basic, or neutral?
- 15) In butter, $[\text{H}^+] = 6.0 \times 10^{-7}$ M.
 - a) Find $[\text{OH}^-]$
 - b) What is the pH?
 - c) Is the solution acidic, basic, or neutral?
- 16) In peaches, $[\text{OH}^-] = 3.16 \times 10^{-11}$ M.
 - a) Find $[\text{H}^+]$
 - b) What is the pH?
 - c) Is the solution acidic, basic, or neutral?
- 17) During the course of the day, human saliva varies between being acidic and basic. If $[\text{OH}^-] = 3.16 \times 10^{-8}$ M,
 - a) Find $[\text{H}^+]$
 - b) What is the pH?

Chapter 8 Summary

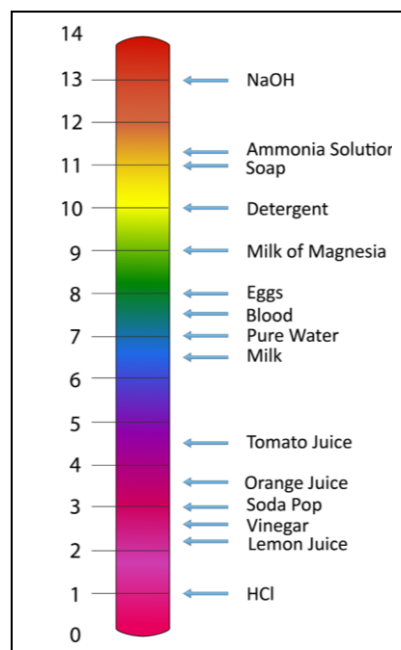
- c) Is the solution acidic, basic, or neutral?
- 18) A solution contains $4.33 \times 10^{-8} \text{M}$ hydroxide ions. What is the pH of the solution?
- 19) A solution contains a hydrogen ion concentration of $6.43 \times 10^{-9} \text{M}$. What is the pH of the solution?
- 20) If the pH of one solution is 5 less than another solution, how does the amount of H^+ in each solution compare? Which has more H^+ ? How many times more?

#21-23: Use the diagram of the pH scale. For each pair of solutions give, indicate which has a greater concentration of H^+ ions. Also indicate how many times more H^+ that solution has.

21) Detergent vs. pure water

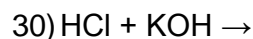
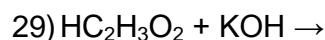
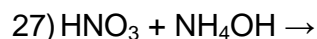
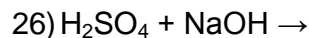
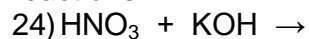
22) Orange juice vs. milk

23) Soda pop vs. milk of magnesia



8.3: Reactions Between Acids and Bases Review Questions

#21-33: Write a balanced reaction for each of the following neutralization reactions:



Chapter 8 Summary

31) Milk of magnesia, $\text{Mg}(\text{OH})_2$ 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.

32) Hydrochloric acid (HCl) reacts with barium hydroxide.

33) Sodium hydroxide reacts with perchloric acid (HClO_4).

8.4: Titrations Review Questions

34) What is an indicator? What is it used for?

35) What is an equivalence point?

36) 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 ?

37) 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.

38) What volume of 0.45 M hydrochloric acid must be added to 15.0 mL of .997 M potassium hydroxide to neutralize the base? ($\text{HCl} + \text{KOH} \rightarrow \text{H}_2\text{O} + \text{KCl}$)

39) What volume of .20 M HI is needed to neutralize 25 mL of .50 M KOH ?

40) What is the molarity of sodium hydroxide if .174L of the solution is neutralized by .20L of 1.2 M HCl ? ($\text{HCl} + \text{NaOH} \rightarrow \text{H}_2\text{O} + \text{NaCl}$)

41) Suppose we used .150L of 0.500M NaOH and .250L of vinegar (acetic acid solution) of an unknown concentration. What is the molarity of the vinegar? (Balanced reaction is:
 $\text{NaOH}_{(\text{aq})} + \text{HC}_2\text{H}_3\text{O}_2_{(\text{aq})} \rightarrow \text{NaC}_2\text{H}_3\text{O}_2_{(\text{aq})} + \text{H}_2\text{O}_{(\text{l})}$)

Chapter 9: Energy Changes



We use energy all the time – to move our bodies, run our cell phones, etc. Energy can be produced in several ways such as wind mills, solar panels, and batteries. But that energy had to come from somewhere, one of the sources is chemical potential energy. In this chapter, we will look at sources of energy.

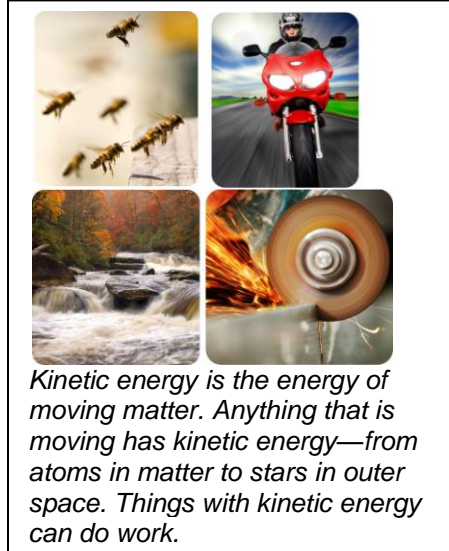
9.1: Conservation of Energy

What is Energy?

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.

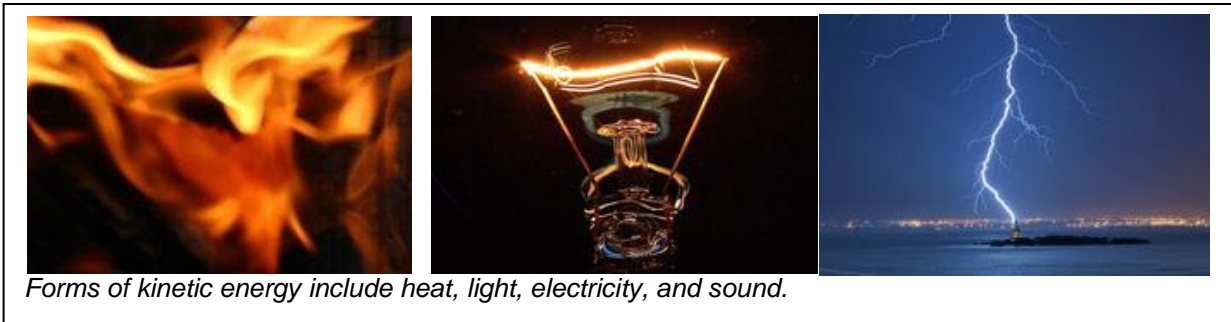
Kinetic Energy

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. Forms of kinetic energy include heat, light, sound, and electricity.



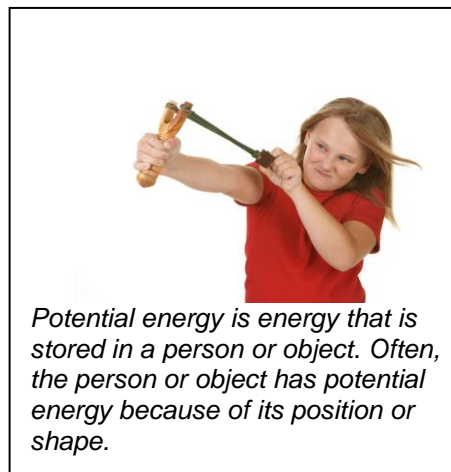
Objectives:

- List types of kinetic and potential energies
- Describe what is meant by the law of conservation of energy

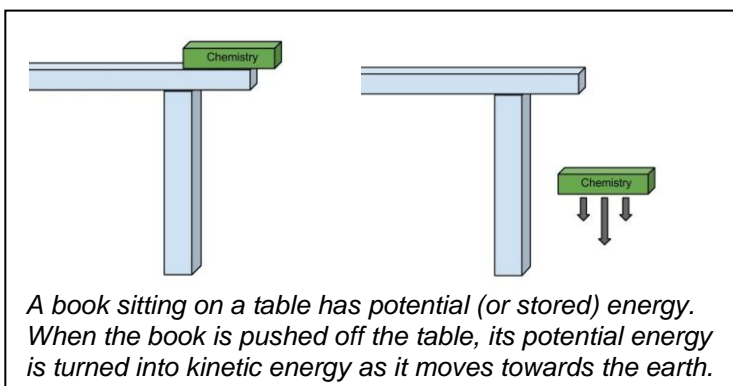


Potential Energy

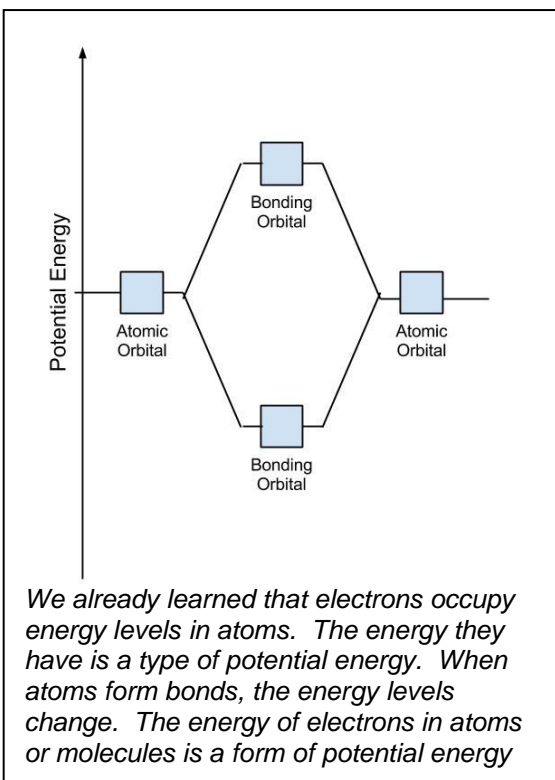
Potential energy is stored energy that remains available until we choose to use it. Think of a battery in a flashlight. If you leave a flashlight on, the battery will run out of energy within a couple of hours. If, instead, you only use the flashlight when you need it and turn it off when you don't, 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.



9.1: Conservation of Energy



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 and the energy of the electrons within the atoms. 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.

The Law of Conservation of 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**.

Kinetic Energy: energy associated with motion

Potential Energy: energy associated with position; stored energy

Law of Conservation of Energy: Energy cannot be created or destroyed but can change from one form to another form

9.2: Endothermic and Exothermic Changes

All Chemical Reactions Involve Changes in Energy

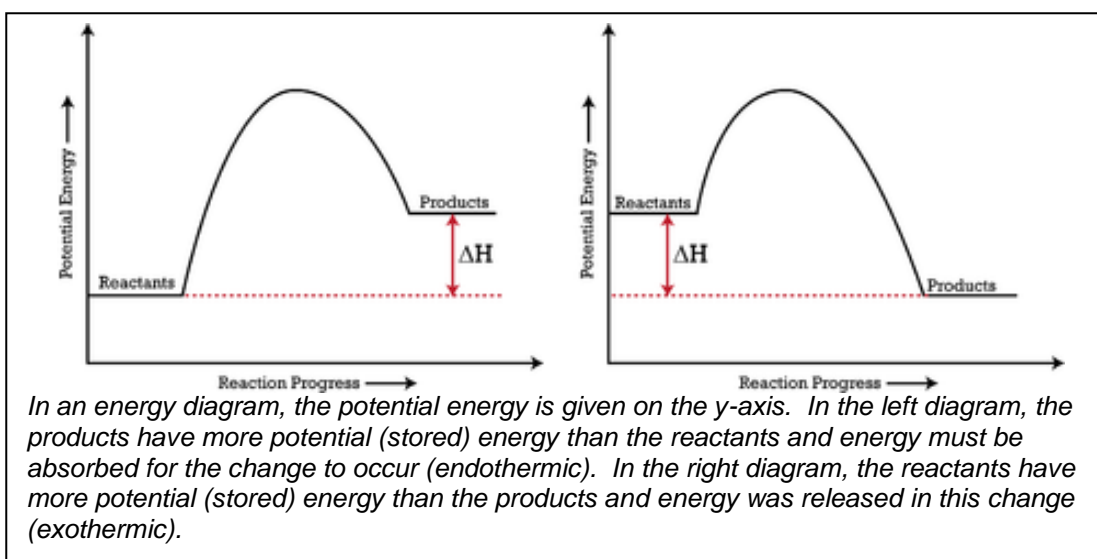
Remember that all chemical reactions involve a change in the bonds of the reactants. The bonds in the reactants are broken and the bonds of the products are formed. Chemical bonds have **potential energy** or "stored energy". Because we are changing the bonding, this means we are also changing how much 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.

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**

Objectives:

-Label a reaction as endothermic or exothermic based on energy diagrams, temperature changes, or the absorbance or release of energy



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.

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.

Exothermic Reaction: a reaction in which kinetic energy is released

Endothermic Reaction: a reaction in which kinetic energy is absorbed

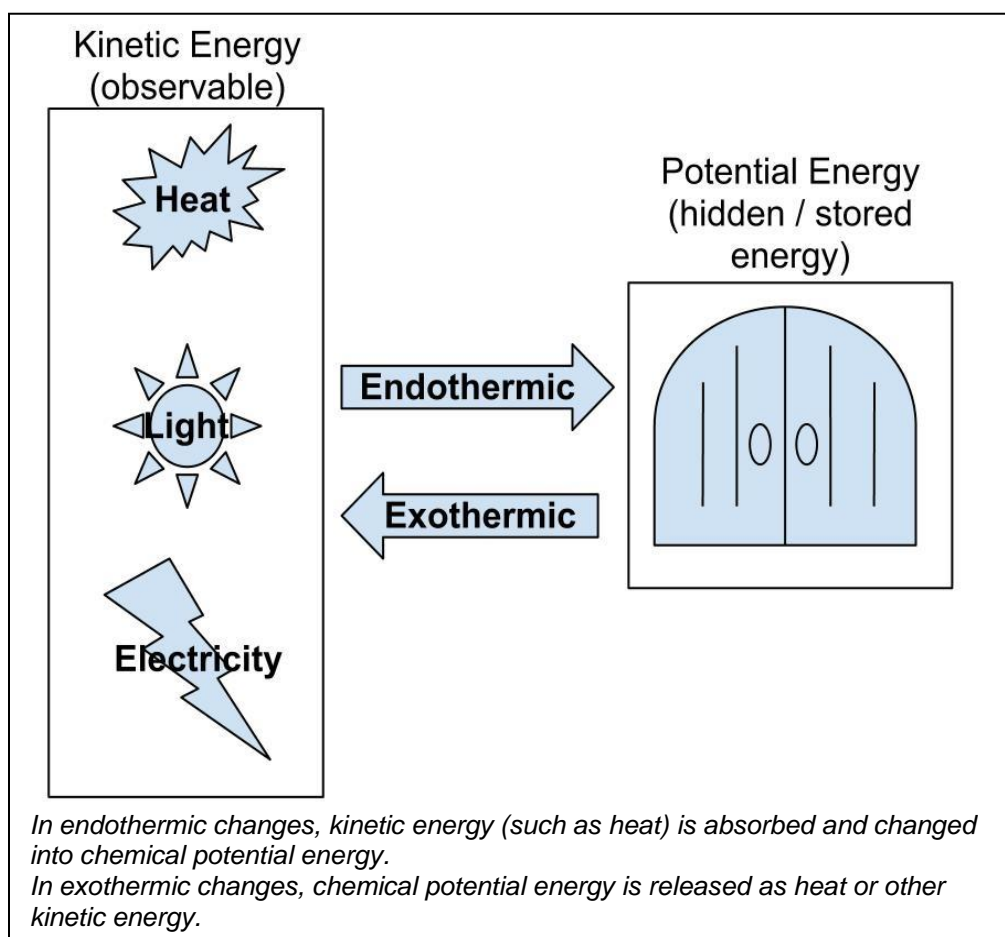
9.2: Endothermic and Exothermic Changes

Identifying Endothermic and Exothermic Changes

Exothermic reactions are reactions that produce kinetic energy. This energy given off comes from some of the potential energy (stored energy) being turned into kinetic energy. This means that heat, light, or electricity is given off.

If you were to measure the temperature change of an exothermic reaction, the temperature would increase over time, as some of the hidden energy turns into heat energy, and, therefore, increases the temperature.

All **endothermic reactions** absorb kinetic energy and the energy is transformed into potential (or stored) energy. These are reactions in



which you must add heat, light or electricity.

If you were to measure the temperature change of an endothermic reaction, the temperature would decrease. The reaction absorbs heat from the surroundings as the heat is transformed into potential energy and the temperature decreases as the heat decreases.

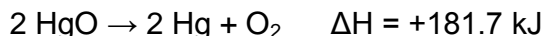
ΔH , Change in Enthalpy

Another way of classifying a reaction as endothermic or exothermic was already presented to you. 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

9.2: Endothermic and Exothermic Changes

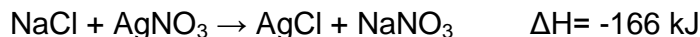
exothermic and energy is produced with the rest of the products and given off into the surroundings.

Consider the following equation:



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.

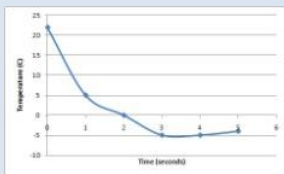
Contrast the previous reaction to the next reaction:



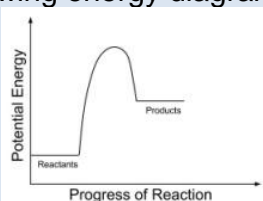
This is an example of a chemical reaction in which energy is released. This means that there is less energy stored in the bonds of the products than there was in the bonds in the reactants. Therefore, extra energy was left over when the reactants become the products. The negative sign (-) on the ΔH tells us that this reaction had extra energy and released this energy into its surroundings. This reaction is exothermic. 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) Gasoline burning
b) The temperature of a reaction is measured and the following graph is produced:



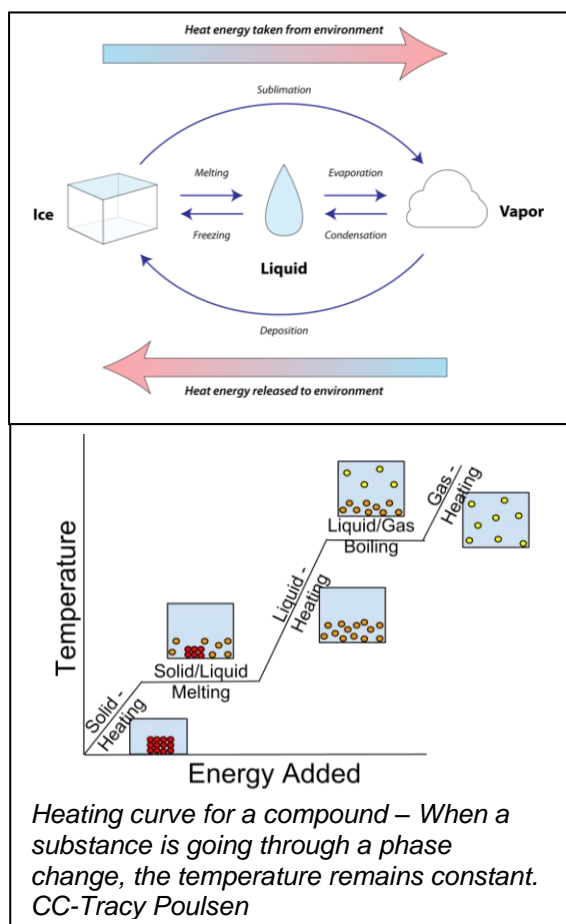
- c) A battery operates to run a cell phone
d) $2 \text{SO}_2 + \text{O}_2 \rightarrow 2 \text{SO}_3 \quad \Delta H = -46.8 \text{ kJ/mol}$
e) A reaction produces the following energy diagram:



Solution:

- a) Exothermic – when you burn something, it feels hot to you because it is given off heat into the surroundings
b) Endothermic – the temperature drops as the reactants absorb energy from their surroundings in an endothermic reaction
c) Exothermic – batteries produce electricity. If kinetic energy (including electricity) is given off, the change is exothermic.
d) Exothermic – because ΔH is negative, meaning some of the potential energy was released
e) Endothermic – the reactants have less potential energy than the products, meaning the reaction had to absorb kinetic energy and turn it into potential energy

9.2: Endothermic and Exothermic Changes Energy in Physical Changes



Physical changes also involve changes in energy. In chemical changes, the energy changes occur due to differences in the potential energies of the bonds in the reactants and products. This isn't the case in physical changes, as bonds are not being broken or formed. In physical changes, such as melting and boiling, the energy changes occur in the motion of the particles.

As we add heat to a solid, the motion of the particles increases, that is, the average kinetic energy of the particles increases. At some temperature, the motion of the particles becomes great enough to overcome the attractive forces. The thermal energy that was added to the solid up to this point was absorbed by the solid as kinetic energy. That means that the speed of the molecules increased. In order for the molecules to actually separate from each other (increasing the distance between objects that attract each other), more energy must be added and this energy is absorbed by the particles as potential energy. The potential energy absorbed by a solid as it changes to a liquid is called the heat of fusion or the heat of melting. It is very important to recognize that once the

temperature of a solid has been raised to the melting point by absorbing thermal energy as kinetic energy, it is still necessary for the solid to absorb more thermal energy in the form of potential energy as the molecules separate.

The boiling point of a liquid is the temperature at which the particles have sufficient molecular motion to exist in the form of a gas (at the given pressure). Once again, however, in order for the particles to separate to the gaseous form, they must absorb a sufficient amount of potential energy. The amount of potential energy necessary for a liquid to turn into a gaseous form is called the heat of vaporization.

Example: Label each of the following processes as endothermic or exothermic.

- a) water boiling
- c) ice forming on a pond

Solution:

- a) endothermic – you must put a pan of water on the stove and give it heat in order to get water to boil. You are adding heat/energy, so the reaction is endothermic.
- 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.

9.3: Electrochemistry

Reactions that Transfer Electrons May Make Electricity

Electricity is an important form of energy that you use every day. It runs your calculators, cell phones, dishwashers, and watches. **Electricity** involves moving electrons through a wire and using the energy of these electrons.

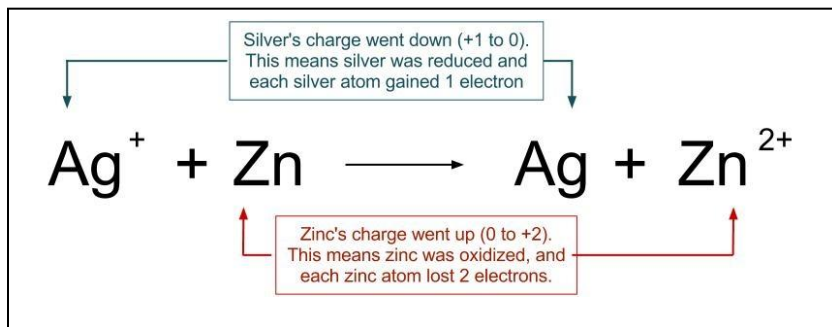
Batteries are one way of producing this type of energy. Many important chemical reactions involve the exchange of one or more electrons, and, therefore we can use this movement of electrons as electricity.

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.



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 zinc atoms are losing electrons to become zinc +2 ions and are being oxidized and the charge of zinc is increasing. Whenever, a chemical reaction involves electrons being transferred from

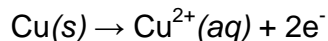
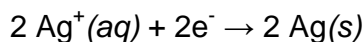
Objectives:

- Identify conditions necessary for a chemical reaction to produce electricity
- Identify the parts of a battery including the anode, cathode, and salt bridge
- Identify reactions that use electricity, such as electrolysis or electroplating

9.3: Electrochemistry

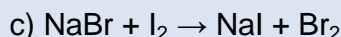
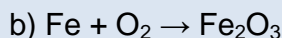
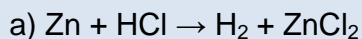
one substance to another, the reaction is an **oxidation – reduction reaction** (or a redox reaction).

Half-reactions 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:



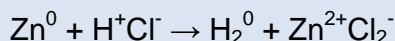
Both oxidation and reduction must occur at the same time so the electrons are donated and absorbed nearly simultaneously. The two half-reactions may be added together to represent a complete reaction. In order to add the half-reactions, the number of electrons donated and the number of electrons accepted must be equal.

Example: For each reaction, identify what is oxidized and what is reduced.



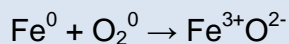
Solution: In order to determine what is being oxidized and reduced, we must look at charges of atoms and see if they increase or decrease. (Remember, elements have no charge. In a compound, we can use our periodic table and what we learned in chapter 4 to assign charges). If the charge increases, the atom was oxidized. If the charge decreases, the atom was reduced.

a) This reaction written with charges is:



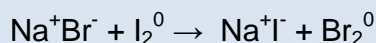
Zn is oxidized because it went from 0 to +2. H is reduced because it went from +1 to 0. Cl was neither oxidized nor reduced.

b) This reaction written with charges is:



Fe 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:



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.

Oxidation: the change in which electrons are lost and the charge goes up

Reduction: the change in which electrons are gained and the charge goes down

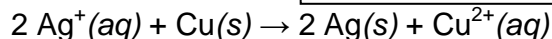
Oxidation-Reduction Reaction: a reaction in which electrons are transferred as one atom loses electrons to another atom

9.3: Electrochemistry

Batteries Produce Electricity in Chemical Changes

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**.

If a strip of copper is placed in a solution of silver nitrate, the following reaction takes place:

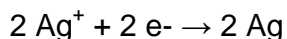
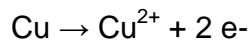


In this reaction, copper atoms are donating electrons to silver ions so the silver ions are reduced to silver atoms and copper atoms are oxidized to copper(II) ions.

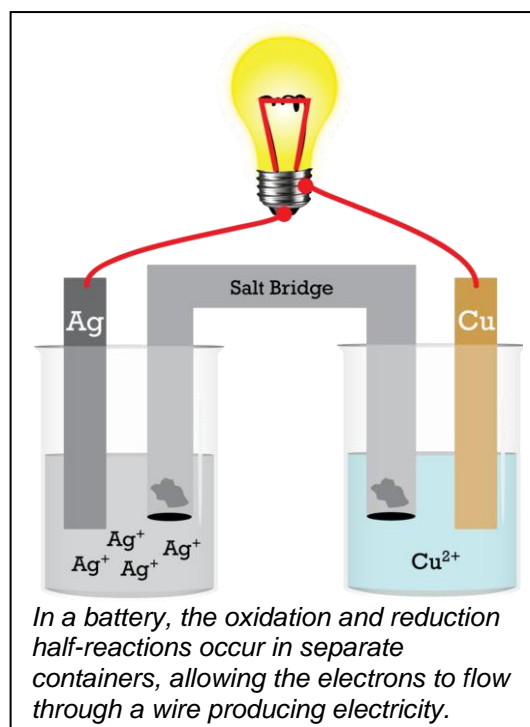
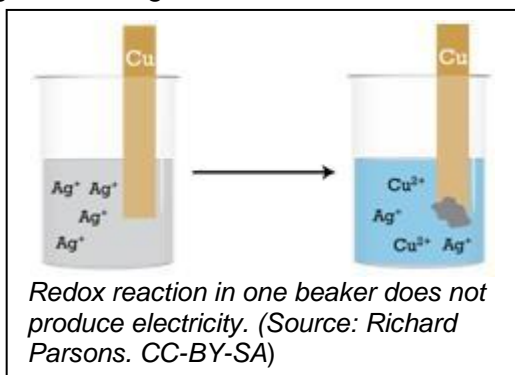
As the reaction occurs, an observer would see the solution slowly turn blue (Cu^{2+} ions are blue in solution) and a mass of solid silver atoms would build up on the copper strip.

The reaction we just described is not set up in such a 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, $2 \text{Ag}^+(aq) + \text{Cu}(s) \rightarrow 2 \text{Ag}(s) + \text{Cu}^{2+}(aq)$, 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.

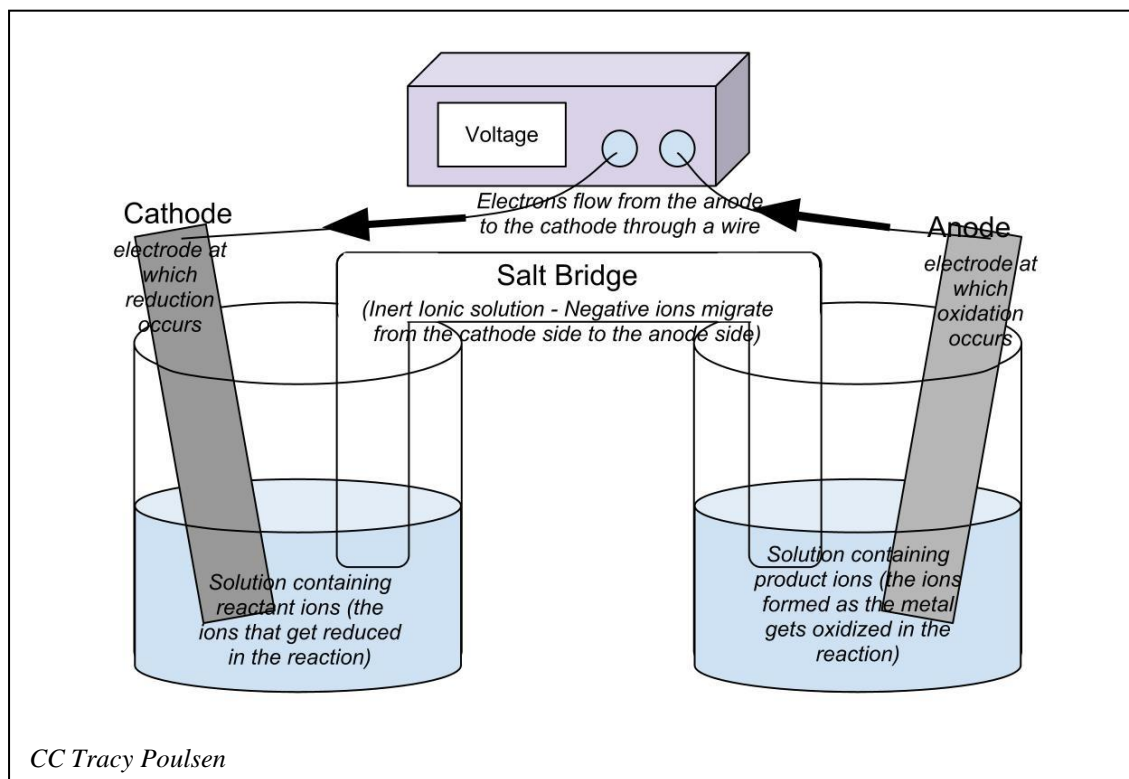


The wire connects the two halves of the reaction, allowing electrons to flow from one metal strip to the other. In



9.3: Electrochemistry

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.



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.

Battery: a device that uses an oxidation-reduction reaction to produce electrical current

9.3: Electrochemistry

Using Electricity in Chemical Changes

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.

Reduction occurs at the positive electrode: $\text{Na}^+ + \text{e}^- \rightarrow \text{Na}$

Oxidation occurs at the negative electrode: $2 \text{Cl}^- \rightarrow \text{Cl}_2 + 2 \text{e}^-$

The overall reaction for this reaction is: $2 \text{Na}^+ + 2 \text{Cl}^- \rightarrow 2 \text{Na} + \text{Cl}_2$

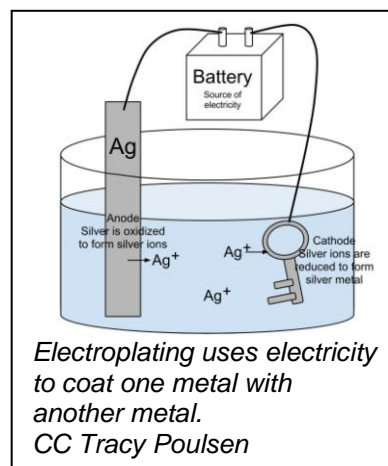
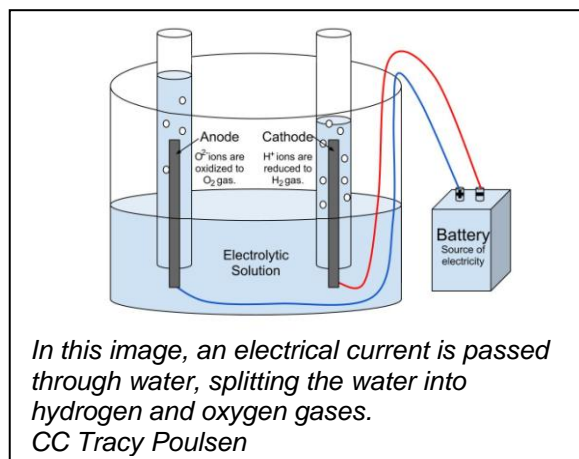
The use of electrolysis to coat one material with a layer of metal is called

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. 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. 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: $\text{Ag}^+ + \text{e}^- \rightarrow \text{Ag}$

Half-reaction at the anode: $\text{Ag} \rightarrow \text{Ag}^+ + \text{e}^-$



Electrolysis: a process that uses electrical energy to cause a chemical reaction to occur

Chapter 9 Summary

9.1: Conservation of Energy

- Forms of kinetic energy include heat, light, electricity, and sound. Kinetic energy can be used to do work.
- Potential energy is a stored energy based on position. Forms of potential energy include chemical potential energy and gravitational potential energy.
- The law of conservation of energy states that energy cannot be created or destroyed. It can only be transferred into other forms of energy.

9.2: Endothermic and Exothermic Changes

- Reactions that absorb heat, light, or electricity are endothermic.
- Other ways to identify endothermic changes are if the temperature decreases or the change in enthalpy (ΔH) has a positive value.
- On an energy diagram, if the products have more potential energy than the reactants, the reaction is endothermic.
- Reactions that release heat, light, or electricity are exothermic.
- Other ways to identify exothermic changes are if the temperature increases or the change in enthalpy (ΔH) has a negative value.
- On an energy diagram, if the products have less potential energy than the reactants, the reaction is exothermic.
- Many physical changes are also endothermic or exothermic. Changing from solid to liquid to gas is endothermic and absorbs heat. Changing in the opposite direction is exothermic.

9.3: Electrochemistry

- Reactions in which there is a transfer of electrons are said to be oxidation-reduction reactions or a redox reactions.
- 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.
- The electrode where oxidation occurs is called the anode and the electrode where reduction occurs is called the cathode.
- Electrolysis reactions use electricity to cause a reaction to occur.
- In electroplating, the object to be plated is made the cathode

Further Reading / Supplemental Links

- Battery simulation: Galvanic Cells
<http://www.mhhe.com/physsci/chemistry/essentialchemistry/flash/galvan5.swf>
- The principles of electrochemical cell design are explained through batteries, sensors, and a solar-powered car. [The Busy Electron \(http://www.learner.org/vod/vod_window.html?pid=807\)](http://www.learner.org/vod/vod_window.html?pid=807)

Chapter 9 Summary

9.1: Conservation of Energy Review Questions

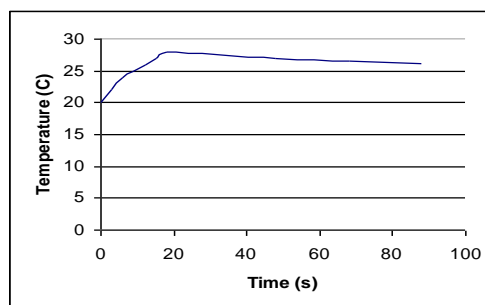
- 1) What does the law of conservation of energy state?
- 2) What is kinetic energy?
- 3) List 3 forms of kinetic energy.
- 4) What is potential energy?
- 5) List 2 forms of potential energy.

9.2: Endothermic and Exothermic Changes Review Questions

- 6) Define endothermic and exothermic reactions.

#7-15: Label each of the following processes as endothermic or exothermic

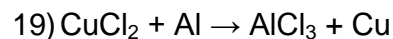
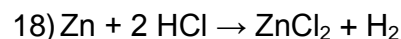
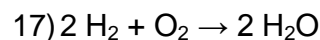
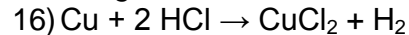
- 7) natural gas burning
- 8) melting chocolate
- 9) fireworks exploding
- 10) Water condensing and freezing to form frost on your windshield
- 11) Photosynthesis (plants using light to make sugar)
- 12) Sugar is dissolved in water in a test tube and the test tube feels cold.
- 13) Gasoline is burned in a car engine.
- 14) Water is converted to steam
- 15) 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.



Chapter 9 Summary

9.3: Electrochemistry Review Questions

#16-19: Identify the element oxidized and the element reduced in the following reactions:

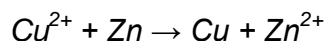


20) What is electricity or electrical current?

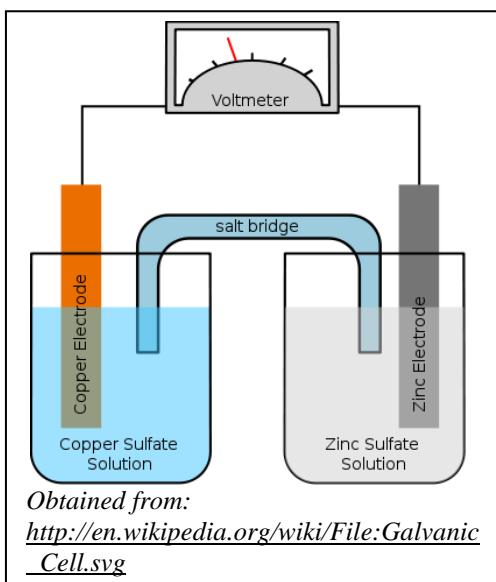
21) What is an anode?

22) What is a cathode?

Consider the battery diagrammed. The overall reaction that occurs in this battery is:



Answer each of the following about the battery:



Obtained from:

http://en.wikipedia.org/wiki/File:Galvanic_Cell.svg

23) What is being oxidized?

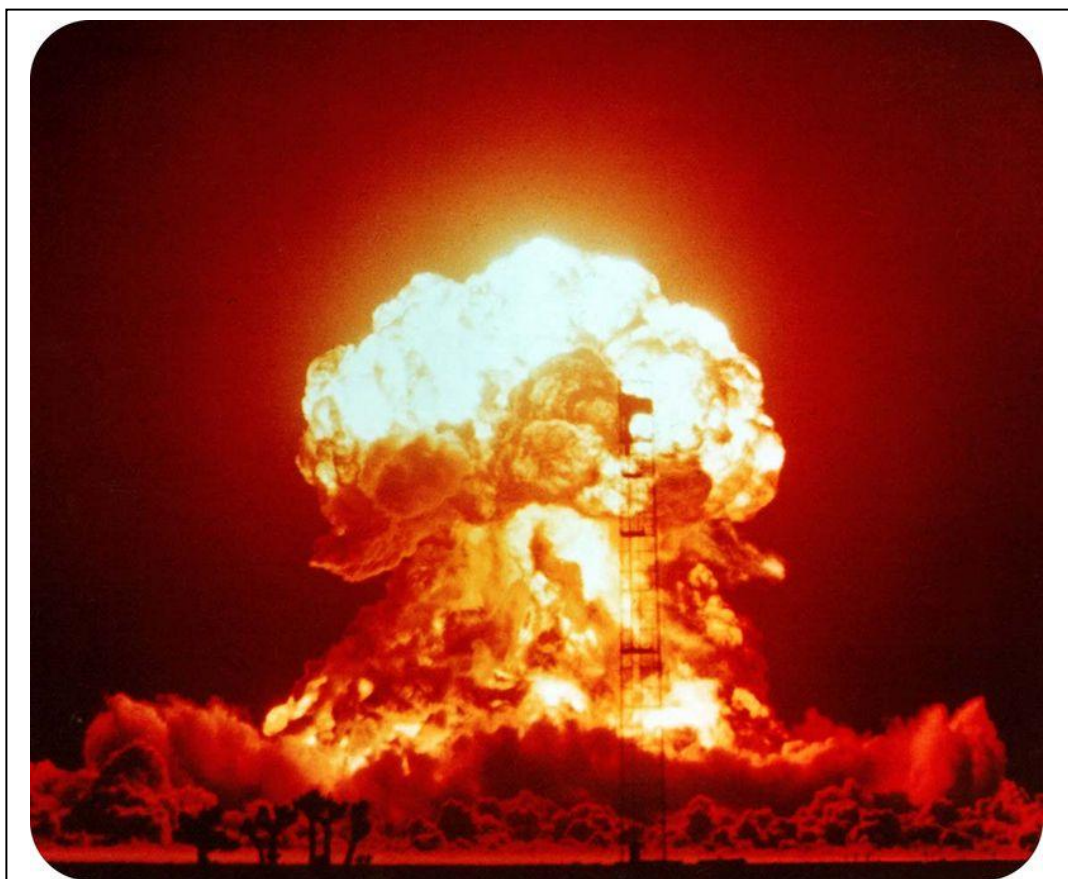
24) What is being reduced?

25) Label the anode. How do you know this is the anode?

26) Label the cathode. How do you know this is the cathode?

27) Which direction will electrons flow through the wire and voltmeter? (Draw an arrow on the diagram.)

Chapter 10: Nuclear Chemistry



This mushroom cloud was produced in a 1953 nuclear bomb test in Nevada. There's no doubt that the explosion gave off a huge amount of energy. Although chemical changes can release or absorb energy, nuclear changes involve much bigger energy changes. In nuclear changes the nucleus of an atom changes, but in chemical changes the electron energy levels are changing.

10.1: Radioactivity

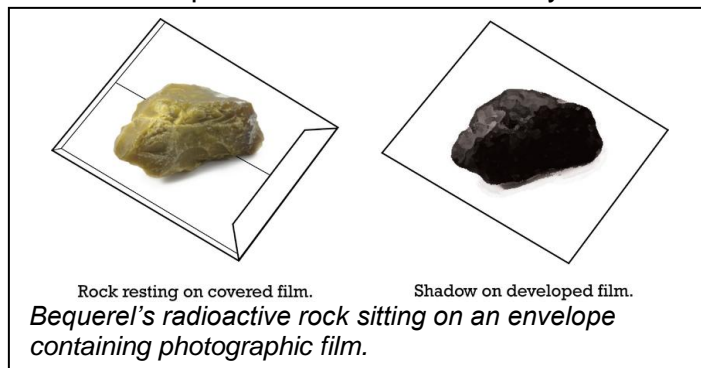
The Discovery of Radioactivity

Objectives:

- Give evidence for the release of radiation
- Compare the amount of energy involved in nuclear and chemical equations.

No one could have known in the 1800s that the discovery 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. At the time of its discovery photography was a strange thing.

Even stranger was the discovery by Roentgen, that radiation other than visible light could expose photographic film. He called this invisible type of light x-rays. When Becquerel heard about Roentgen's discovery, he wondered if his fluorescent minerals would give the same x-rays. He 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, Fortunately, the weather didn't cooperate and Becquerel had to

leave the crystals and film stored in a drawer for several cloudy days. 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. Becquerel had discovered radioactivity.

The Curies and Radium



One of Becquerel's assistants, a young Polish scientist named Maria Skłodowska (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 Joachimsthal 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.

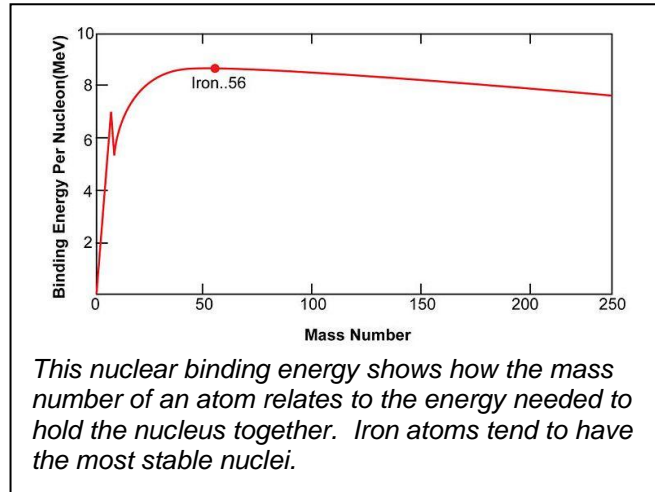
By 1902, the world was aware of a new phenomenon called radioactivity and of new elements which exhibited natural radioactivity. For this work, Becquerel and the Curies shared the 1903 Nobel Prize and for subsequent work, Marie Curie received a second Nobel Prize. She is

the only person ever to receive two Nobel Prizes in science.

10.1: Radioactivity

Energy of Nuclear Changes

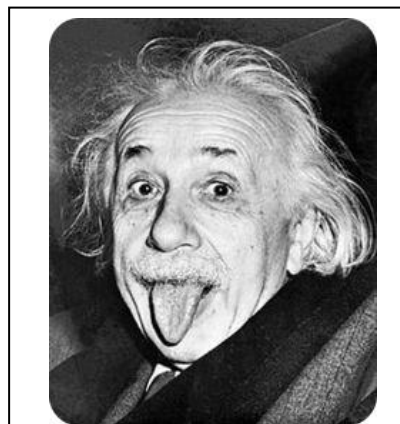
A nucleus (with one exception, hydrogen-1) consists of some number of protons and neutrons pulled together in an extremely tiny volume. Since protons are positively charged and like charges repel, it is clear that protons cannot remain together in the nucleus unless there is a powerful force holding 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 **radioactivity**.

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)!

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 1×10^3 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 1×10^8 kJ/mol. That means that nuclear changes involve almost a million times more energy per atom than chemical changes!!! That's a lot.



Einstein's famous equation $E=mc^2$ shows that a very small amount of mass from the nucleus can be converted into huge amounts of energy.

Nuclear Binding Energy: the energy involved in holding the nucleus together
Radioactivity: the process through which an unstable nucleus emits high energy particles as it changes to become a more stable nucleus

10.2: Types of Radioactive Decay

Three Common Types of Radiation

Objectives:

-Compare the properties alpha, beta, and gamma radiation.

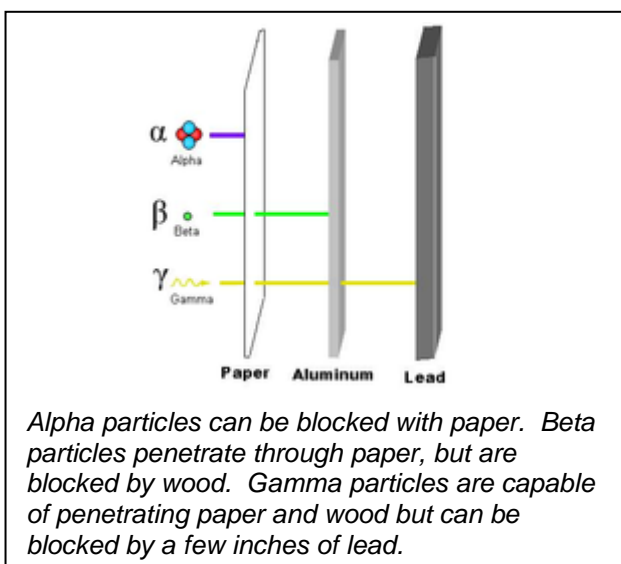
-Write nuclear reactions for alpha and beta decay reactions

Many nuclei are **radioactive**; that is, the unstable nuclei decompose by emitting particles and in doing so, become a different nucleus. In ordinary chemical reactions, atoms of one element never 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.

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

Radiation can cause damage to living tissues. 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.*

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

10.2: Types of Radioactive Decay

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. 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.

<i>Particle</i>	<i>Symbol</i>	<i>Mass</i>	<i>Penetrating Power</i>	<i>Ionizing Power</i>	<i>Shielding</i>
<i>Alpha</i>	α	4 amu	Very Low	Very High	Paper Skin
<i>Beta</i>	β	1/2000 amu	Intermediate	Intermediate	Aluminum
<i>Gamma</i>	γ	0 (energy only)	Very High	Very Low	2 inches lead

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.

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.



You may have seen this sign before—maybe in a hospital. The sign means there is danger of radiation in the area.

Alpha Radiation: the type of radiation emitted from the nucleus equivalent to a helium nucleus

Beta Radiation: the type of radiation released as a neutron is changed into a proton emitting a high energy electron

Gamma Radiation: high energy electromagnetic radiation

10.2: Types of Radioactive Decay

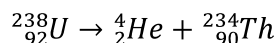
Writing Nuclear Reactions

These types of equations are called nuclear equations. When writing these, there are some general rules that will help you:

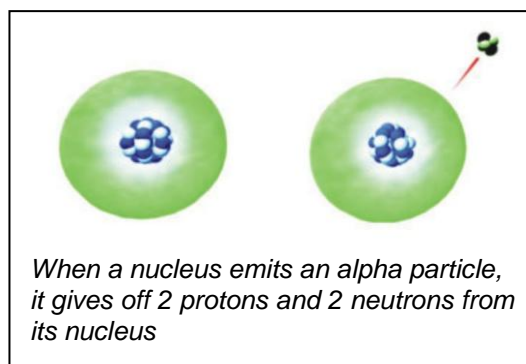
- The sum of the mass numbers (top numbers) on the reactant side equal the sum of the atomic numbers on the product side
- The atomic numbers (bottom numbers) on the two sides of the reaction will also be equal.

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



In this nuclear change, the uranium atom (${}_{92}^{238}\text{U}$) transmuted into an atom of thorium (${}_{90}^{234}\text{Th}$) and, in the process, gave off an alpha particle. Look at the symbol for the alpha particle: ${}_2^4\text{He}$. 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.

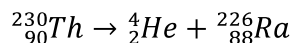


In this equation,

$$\text{mass number: } 238 = 4 + 234.$$

$$\text{atomic number: } 92 = 2 + 90$$

Another alpha particle producer is thorium-230.



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, ${}_2^4\text{He}$.

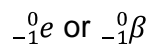
Beta Decay

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. To simplify things, we can picture a neutron breaking into two pieces with the

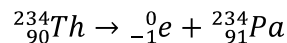
10.2: Types of Radioactive Decay

pieces being 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



Thorium-234 is a nucleus that undergoes beta decay. Here is the nuclear equation for this beta decay.

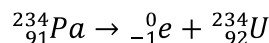


Note that both the mass numbers and the atomic numbers add up properly:

$$\text{mass number: } 234 = 0 + 234$$

$$\text{atomic number: } 90 = -1 + 91$$

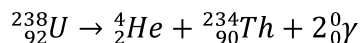
The mass numbers of the original nucleus and the new nucleus are the same because a neutron has been lost, but a proton has been gained and so the sum of protons plus neutrons remains the same. The atomic number in the process has been increased by one since the new nucleus has one more proton than the 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.



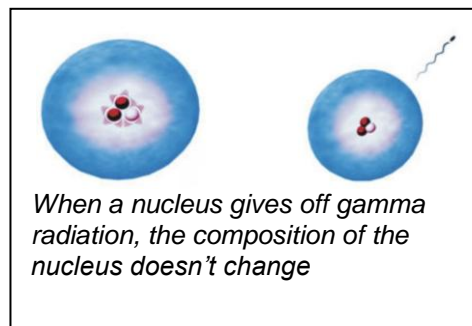
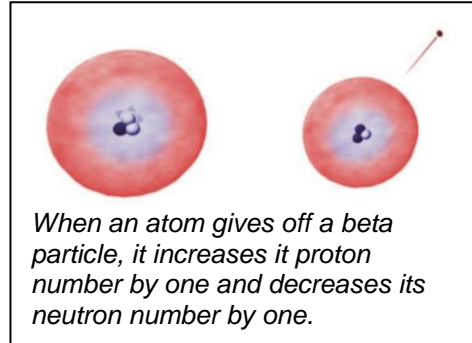
Once again, the atomic number increases by one and the mass number remains the same; confirm that the equation is correctly balanced.

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.



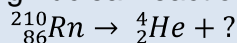
Virtually all of the nuclear reactions in this chapter also emit gamma rays, but for simplicity the gamma rays are generally not shown. Giving off a gamma ray does not by itself change the nucleus. The same element is present both before and after the change.



10.2: Types of Radioactive Decay

Examples Writing Nuclear Reactions

Example: Complete the following nuclear reaction by filling in the missing particle.



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: ${}_{84}^{206}\text{Po}$. 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. $4 + ?$

$$\text{Mass numbers: } 210 = 4 + ?$$

$$\text{Atomic numbers: } 86 = 2 + ?$$

We are left with ${}_{84}^{206}\text{Po}$

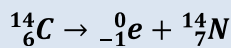
Example: Write each of the following nuclear reactions.

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 ${}_{-1}^0e$. 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:



b) Alpha particles have the symbol ${}_2^4\text{He}$. Emitting an alpha particle causes the atomic number to decrease by 2 and the mass number to decrease by 4. We are left with:



Decay Series



Several elements listed in the periodic table are not found in nature. Many of them were produced in the cyclotron in the radiation laboratory at the University of California at Berkeley under the direction of Glenn Seaborg.

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.

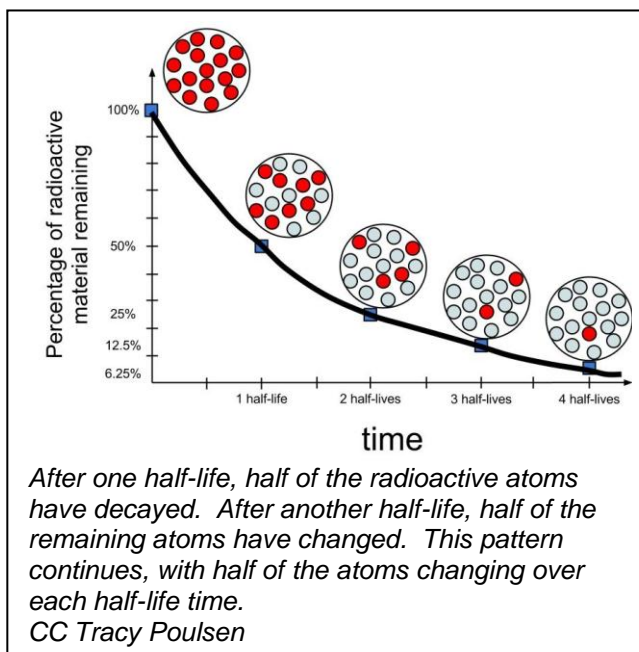
10.3: Rate of Radioactive Decay

Rate of Decay depends on the Half-Life

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 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.



Objectives:

- Describe what is meant by the term half-life
- Indicate factors that influence the half-life of an isotope
- Find the half-life of an isotope from a graph
- Predict the amount of a radioactive isotope remaining after some time

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-251 in 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.

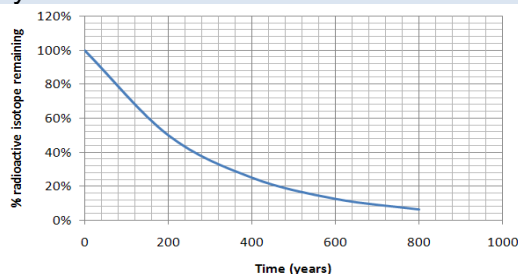
Remember, the half-life is the time it takes for half of your sample, not matter how much you have, to remain. Each half-life will follow the same general pattern as Cf-251. The only difference is the length of time it takes for half of a sample to decay.

Half-life: the amount of time it takes for half of a sample of radioactive atoms to decay

10.3: Rate of Radioactive Decay

Examples of Half-Life Problems

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.

10.3: Rate of Radioactive Decay

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 4.5×10^9 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.

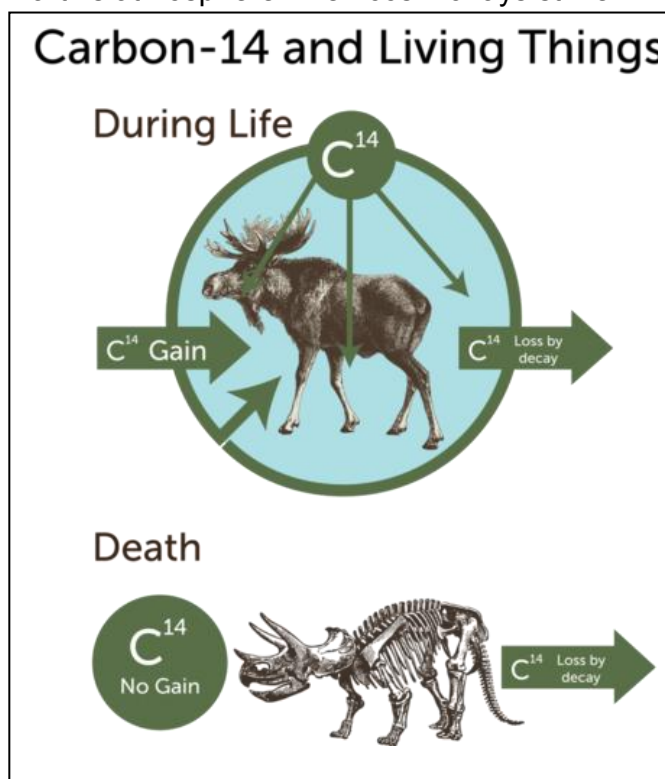
Carbon Dating

One of the most familiar types of radioactive dating is carbon-14 dating. Carbon-14 forms naturally in Earth's atmosphere when cosmic rays strike atoms of nitrogen-14.

Living things take in and use carbon-14, just as they do carbon-12. The carbon-14 in living things gradually decays to nitrogen-14. However, as it decays, it is constantly replaced because living things keep taking in carbon-14. As a result, there is a constant ratio of carbon-14 to carbon-12 in organisms as long as they are alive. After organisms die, the carbon-14 they already contain continues to decay, but it is no longer replaced. Therefore, the carbon-14 in a dead organism constantly declines at a fixed rate equal to the half-life of

carbon-14. Half of the remaining carbon-14 decays every 5,700 years. If you measure how much carbon-14 is left in a fossil, you can determine how many half-lives (and how many years) have passed since the organism died.

These procedures have been used to determine the age of organic artifacts and determine, for instance, whether art works are real or fake. Carbon-dating has its limits, though. It can only be used for materials that were once living. Additionally, after about 60,000 years too little Carbon-14 is left to be accurately measured so this technique will no longer work.



Carbon Dating: the process of using the amount of C-14 in a once-living sample and the half-life of C-14 to find the age of the sample

Objectives:

- Identify sources of radiation
- Describe biological hazards of radiation
- Give medical uses of radiation

10.4: Applications of Radioisotopes

We Are All Exposed to Radiation

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 sources in the earth (such as naturally occurring radioactive elements which are incorporated in building materials and also in the human body) ; sources from space in the form of cosmic rays; and 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.

Hazards 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.

Uses of 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.



This machine scans a patient's body and detects radiation

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.

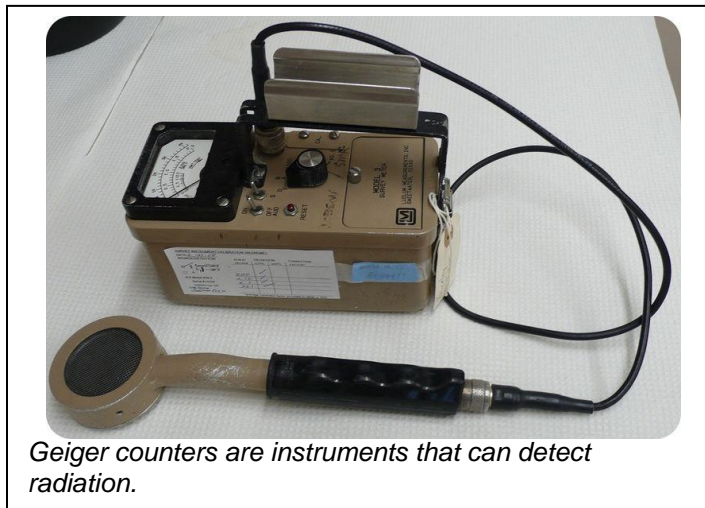
10.4: Applications of Radioisotopes

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.

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 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 several types of cancer, including: breast cancer, colorectal cancer, head and neck cancer, lung cancer, and prostate cancer.

10.5: Fission & Fusion

Objectives:

- Compare fission and fusion reactions
- Give applications of fission and fusion reactions

Fusion

Nuclei that are larger than iron-56 may undergo nuclear reactions in which they break up into two or more smaller nuclei.

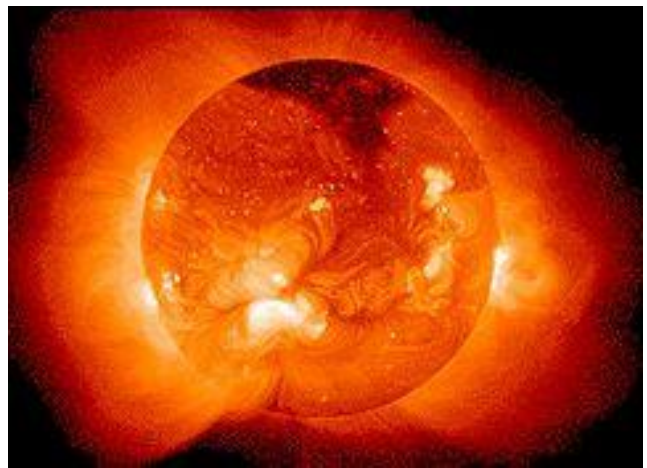
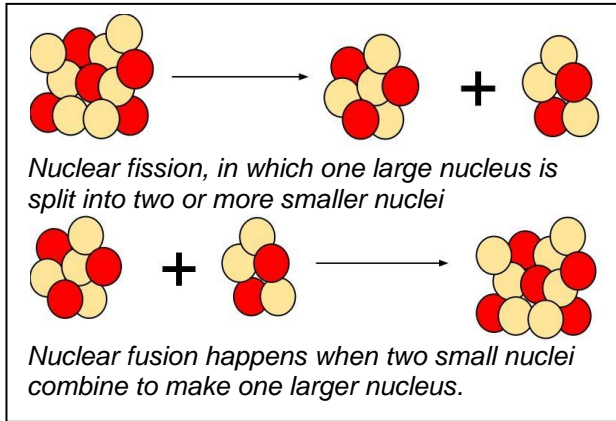
These reactions are called **fission** reactions.

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 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.

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 that won't melt under the conditions of a fusion reaction. Electricity-producing fusion reactors are still a distant dream.



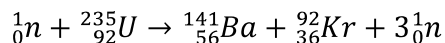
The energy that comes from the sun and other stars is produced by fusion. (Source: <http://commons.wikimedia.org/wiki/File:Sun-in-X-ray>. Public Domain)

Fusion: the process in which two nuclei combine to make one larger nucleus
Fission: the process in which one nucleus is split into two smaller nuclei

10.5: Fission & Fusion

Fission and Chain Reactions

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.



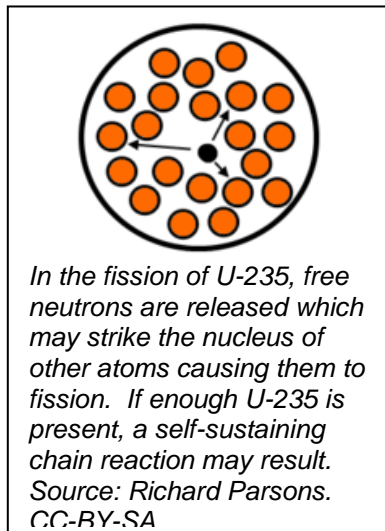
The products shown are only one of many sets of products from the disintegration of a U-235 nucleus. Over 35 different elements have been observed in the fission products of U-235.

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.

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.

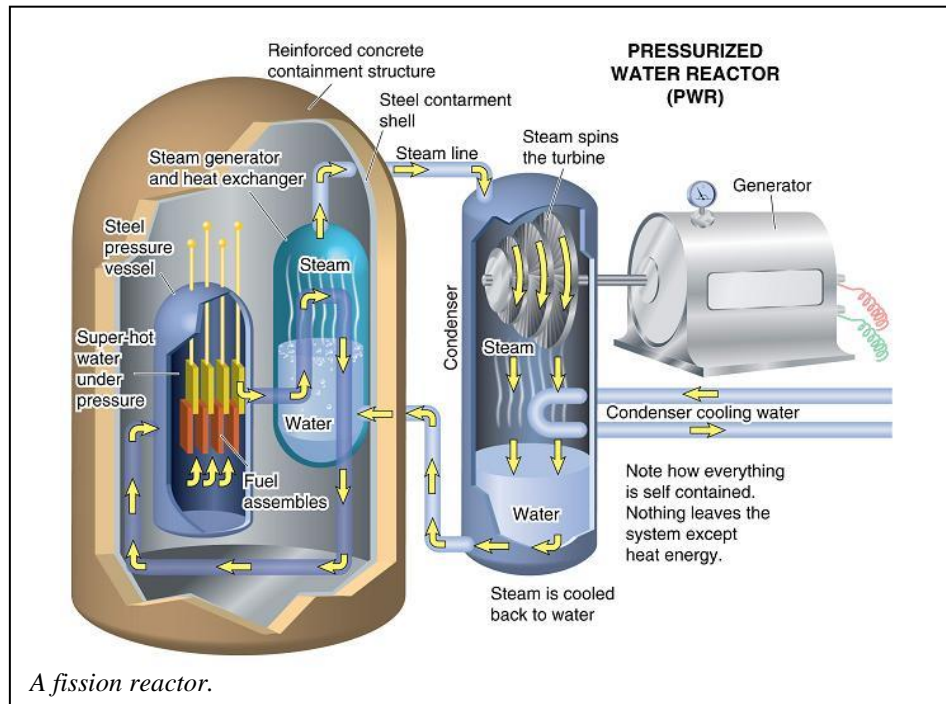
U-235 is the isotope of uranium that will undergo fission. Once enough U-235 is acquired, it is placed in a series of long cylindrical tubes called fuel rods. These fuel cylinders are bundled together with control rods made of neutron-absorbing material. (In the United States, all public nuclear power plants contain less than a critical mass of U-235 and therefore, could never produce a nuclear explosion.) The amount of heat generated



10.5: Fission & Fusion

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.

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.



A fission reactor.

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 CO₂ annually, accounting for 40% of U.S. CO₂ emissions and 10% of global emissions. These coal burning plants also produce 64% of the sulfur dioxide emissions and 26% of the nitrous oxide emissions, which cause acid rain.

10.6: The Origin of Elements

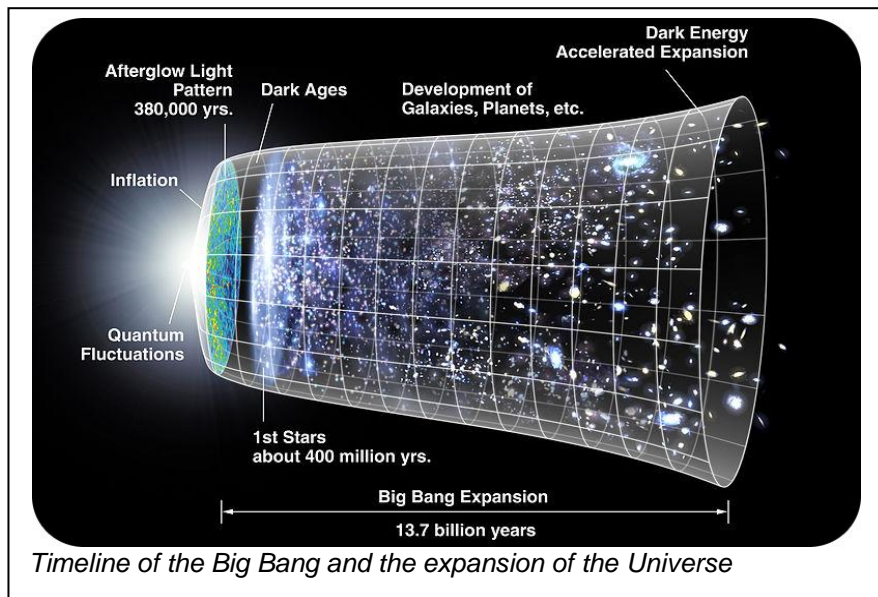
The Big Bang Theory

The **Big Bang** is the currently accepted theory of the early development of the universe. The theory suggests that the universe was originally an extremely hot and dense point in the space at some time in the past and has since cooled by expanding to the present state. The universe continues to expand today. Remember, theories are supposed to explain observations in science and are also used to make predictions. The theory is supported by the most comprehensive data and is able to explain a wide range of observations.

According to the best available measurements the big bang occurred about 13.75 billion years ago. The Universe would have cooled sufficiently to allow energy to be converted into subatomic particles (protons, neutrons, and electrons and many other particles). While protons and neutrons would have formed the first atomic nuclei only a few minutes after the Big Bang, it would then have taken thousands of years for electrons to lose enough energy to form neutral atoms. The first element produced would be hydrogen. Giant clouds of these primordial elements would then form stars and galaxies. Other elements were formed by fusion within the stars.

Objectives:

- Give evidence for the big bang theory
- Give the relative amounts of elements in the universe and on earth
- Describe how elements are formed in the big bang, stars, and super nova.



Evidence for the Big Bang Theory

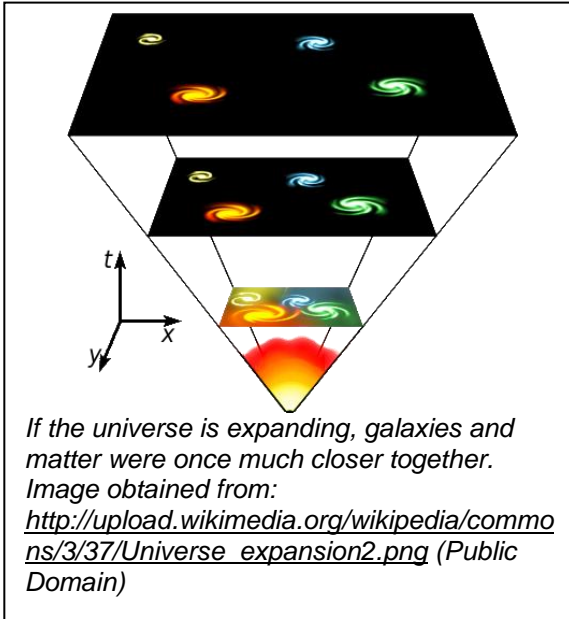
Many scientists have contributed to gathering evidence and developing theories to contribute to our understanding of the origin of the universe and the Big Bang Theory. Georges Lemaître, a Belgian priest, physicist, and astronomer was the first person to propose the theory of the expansion of the Universe 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

10.6: The Origin of Elements

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

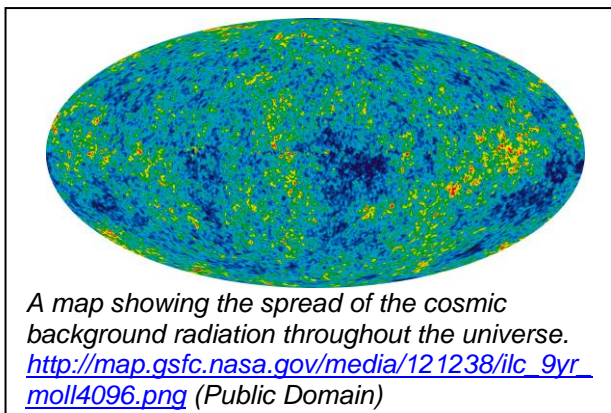


away from each other and that galaxies that are further away are moving faster, as first suggested by Lemaître in 1927. Hubble's evidence is now known as red shift. This discovery was the first observational support for the Big Bang Theory. If the distance between galaxies is increasing today, then galaxies 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. In 1964 Arno Penzias and Robert Wilson serendipitously discovered the cosmic background radiation. Their discovery confirmed the predictions. After cosmic microwave background radiation was discovered in 1964 and the analysis matched the amount of missing radiation from the Big Bang,

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.

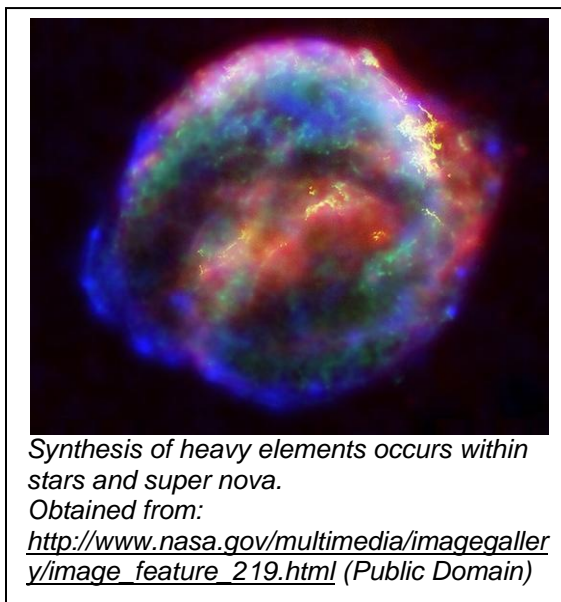
10.6: The Origin of Elements

The Formation of Elements

According to predictions made by the Big Bang Theory, scientists theorized 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. This is indeed what scientists have found.

As the universe expanded and cooled it allowed for the formation of protons and neutrons, forming mostly hydrogen and helium atoms. These two elements remain the most abundant in the universe, with hydrogen atoms making up roughly 74% and helium making up roughly 24% of the mass of the universe. Very abundant hydrogen and helium are products of the Big Bang

Elements heavier than hydrogen and helium were mostly produced later within stars. As the hydrogen and helium condensed together to form stars, the hydrogen and helium were able to fuse to make the larger elements. 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. Only about 2% (by mass) of the Milky Way galaxy's disk is composed of heavy elements.



The more distant galaxies are being viewed as they appeared in the past, so their abundances of elements appear closer to the primordial mixture. As physical laws and processes appear common throughout the visible universe, however, it is expected that these galaxies will likewise have evolved similar abundances of elements.

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 in 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.

Chapter 10 Summary

10.1: Radioactivity

- Henri Becquerral, Marie Curie, and Pierre Curie shared the discovery of radioactivity.
- A nuclear reaction is one that changes the structure of the nucleus of an atom.
- Nuclear changes involve millions of times more energy per atoms than chemical changes (such as burning, etc)

10.2: Types of Radioactive Decay

- The two most common modes of natural radioactivity are alpha decay and beta decay.
- Alpha particles are equivalent to helium nuclei, containing 2 protons and 2 neutrons. They have a charge of +2 and a mass of 4 amu.
- Beta particles are ejected from the nucleus when a neutron changes into a proton.
- The beta particle is equivalent to an electron with a charge of -1 and a mass of approximately 1/2000 amu.
- Most nuclear reactions emit energy in the form of gamma rays, which is high energy electromagnetic radiation.
- The atomic numbers and mass numbers in a nuclear equation must be balanced.

10.3: Rate of Radioactive Decay

- 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.

10.4: Applications of Radioisotopes

- There are many sources of naturally occurring radiation.
- Radiation can cause damage to organisms including tissue, DNA, and other damage.
- Radiation has several medical uses, including identifying and treatment of many diseases.

10.5: Fission and Fusion

- 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.
- Nuclear power plants use fission to

10.6: The Origin of Elements

- The Big Bang Theory proposes that all matter in the universe was once contained in a small point, but has since expanded and cooled.

Chapter 10 Summary

- The theory is supported by scientists as it provides a satisfactory explanation for the observations that the universe is expanding today, cosmic background radiation, etc.
- The theory also explains the observation that universe is composed mostly of hydrogen and helium, which formed shortly after the big bang
- The theory also provides as explanation for where elements heavier than hydrogen were formed, through fusion into heavier elements in stars and super nova
- The earth consists mostly of much heavier elements including oxygen, silicon, aluminum, and iron.

Further Reading / Supplementary Links

- Tutorial: Vision Learning: Nuclear Chemistry:
http://visionlearning.com/library/module_viewer.php?mid=59&l=&c3=

10.1: Radioactivity Review Questions

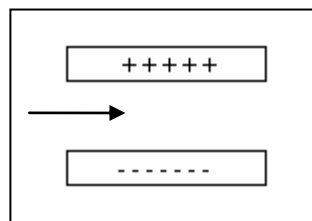
- 1) What is radiation?
- 2) What is the nuclear binding energy?
- 3) In nuclear changes, what part of the atom is changing? In chemical changes, what part of the atom is changing?
- 4) Compared to ordinary chemical reactions (such as burning wood), how much energy is given off in nuclear reactions?

10.2: Types of Radioactive Decay Review Questions

#5-7: Match the following descriptions to the appropriate type of radiation.

- | | |
|-------------------|--|
| 5) alpha particle | a. high energy electromagnetic radiation |
| 6) beta particle | b. a high speed electron |
| 7) gamma ray | c. a helium nucleus |

- 8) 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.



Chapter 10 Summary

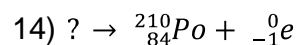
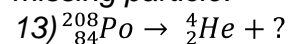
9) When a nucleus gives off a beta particle, what effect does this have on the number of protons, neutrons, and total number of particles in the nucleus?

10) 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?

11) Which of the three common emissions from radioactive sources requires the heaviest shielding?

12) Which type of radiation is most dangerous to living tissues?

#13-14: Complete the following nuclear equations by supplying the missing particle.



#15-20: Write each of the following nuclear reactions as a nuclear equation.

15) The radioactive isotope iodine-131 (used to study thyroid function) decays by beta emission.

16) Thorium-230 decays by alpha emission.

17) ${}^{234}\text{Th}$ decays by beta emission.

18) Hydrogen-3 decays by beta emission.

19) The alpha decay of radon-198.

20) The beta decay of uranium-237.

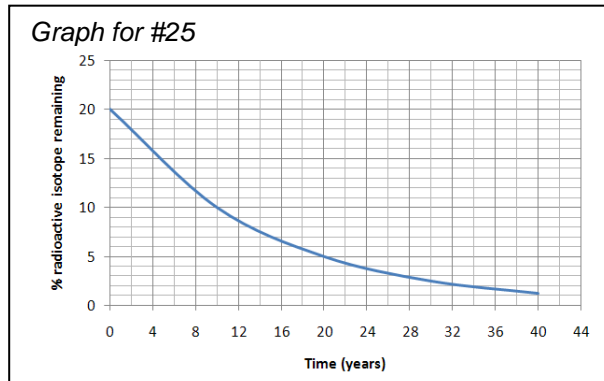
10.3: Rate of Radioactive Decay Review Questions

21) The half-life of radium-226 is about 1600 years. How many grams of a 2.00 gram sample will remain after 4800 years?

Chapter 10 Summary

- 22) 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?
- 23) A radioactive isotope decayed from 24.0 grams to 0.75 grams in 40.0 years. What is the half-life of the isotope?
- 24) The half-life of C-14 is 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?

- 25) What is the half-life of the isotope which produced the given graph over time?



10.4: Applications of Radioisotopes Review Questions

#26-30: For each of the following, indicate which type of nuclear change is used (fission, fusion, or nuclear decay/radiation):

- 26) Nuclear power plants
- 27) Biological tracers
- 28) Energy from the stars
- 29) Cancer treatment
- 30) Nuclear warheads

10.5: Fission and Fusion Review Questions

- 31) What is fission?
- 32) What is fusion?

Chapter 10 Summary

- 33) What is the primary physical difference between a nuclear electricity generating plant and a coal-burning electricity generating plant?
- 34) What do the control rods in a nuclear reactor do and how do they do it?
- 35) Is it possible for a nuclear explosion to occur in a nuclear reactor? Why or why not?

10.6: The Origin of Elements Review Questions

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

Chapter 11: Behavior of Gases



A hiker pauses to view the impressive peak of Mount Everest, the tallest mountain in the world. At the top of Mount Everest, the air is very thin. Climbers may need oxygen tanks to get enough oxygen to breathe, even though oxygen is the second most plentiful gas in the atmosphere.

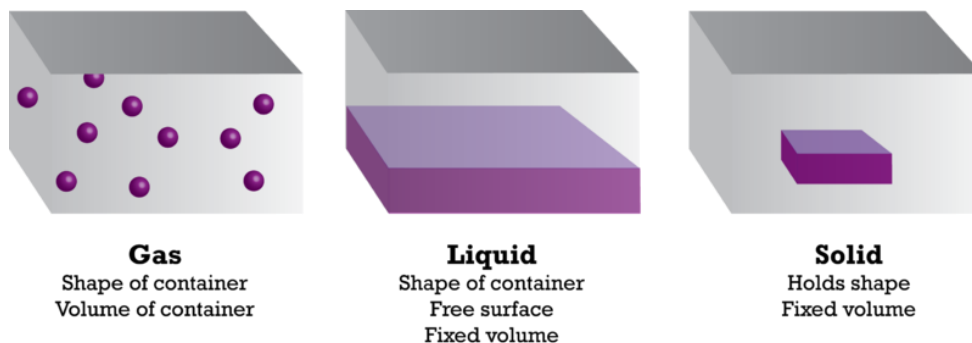
11.1: Kinetic Theory of Gases

Properties of Gases

Objectives:

- Summarize the kinetic theory of matter
- Convert between various units for volume, temperature, and pressure
- Describe why units of Kelvin must always be used when performing calculations involving temperature
- Compare the kinetic energy and speed of various gas particles at the same temperature

The most common states of matter on Earth are solids, liquids, and gases. How do these states of matter differ? Their properties are contrasted in the diagram below.



In gases, because there is little attraction between particles, they spread out to take the shape and volume of the container. Particles are very far apart. In liquids, the particles have strong enough attractive forces to keep the particles together. Liquids will take the shape of the container, but have a fixed volume. Particles in solids have very little motion, as particles are close together and have a strong attraction for each other. Solids have both fixed shape and volume.

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.

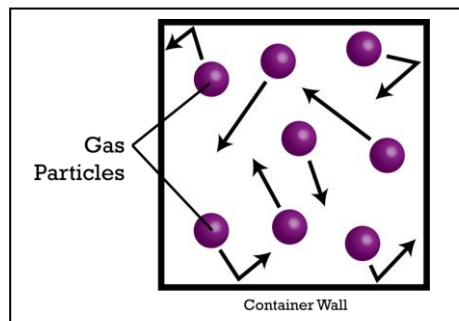
The two most common ways of expressing volume are using mL and L. You will need to be able to convert between these two units. The relationship is as follows:

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

11.1: Kinetic Theory of Gases

Pressure of Gases

Did you ever use a bicycle pump? When you push down on the handle, it forces air out through the hose, and the air enters the tire through a tiny opening. Like other gases, air can flow and take the shape of its container. The air that enters the tire from the pump quickly spreads out to fill the entire tire evenly. As the tire fills with air, it feels firmer. That's because the air exerts pressure against the inside surface of the tire.



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.

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/in². 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.

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.

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.

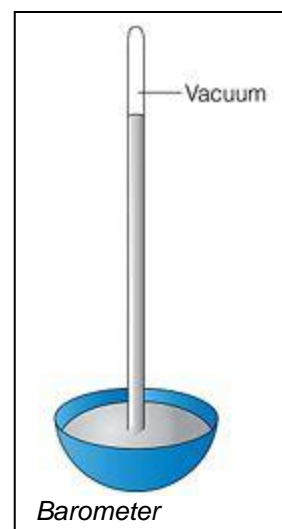
$$pressure = \frac{force}{area}$$

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

11.1: Kinetic Theory of Gases

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:

$$1.00 \text{ atm} = 760. \text{ mmHg}$$

Example: Convert 425 mmHg to atm.

Solution

The conversion factor is $760. \text{ mmHg} = 1.00 \text{ atm}$

$$425 \text{ mmHg} \times \frac{1 \text{ atm}}{760 \text{ mmHg}} = 0.559 \text{ atm}$$

11.1: Kinetic Theory of Gases

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 $KE = \frac{1}{2}mv^2$, 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. 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. 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.

That means that if the temperature, in Kelvin, is doubled the kinetic energy of the particles is also doubled. This relationship 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. 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 $H_2 = 2(1.008 \text{ g/mol}) = 2.016 \text{ g/mol}$

Mass of $O_2 = 2(16.00 \text{ g/mol}) = 32.00 \text{ g/mol}$

Mass of $N_2 = 2(14.01 \text{ g/mol}) = 28.02 \text{ g/mol}$

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.

11.2: Gas Laws

Boyle's Law: Pressure vs. Volume

Objectives:

- Describe how changing the volume, temperature, or pressure of a gas sample affects one of the other variables
- Perform calculations using one of many gas laws

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**.

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 the figure. 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.

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.

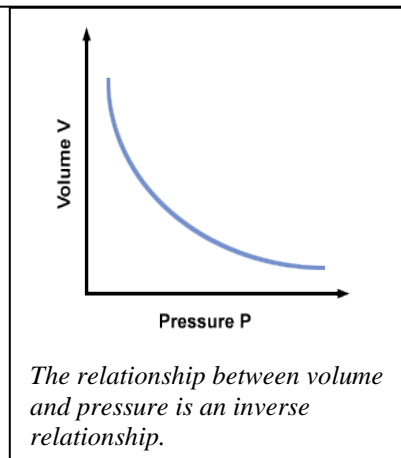
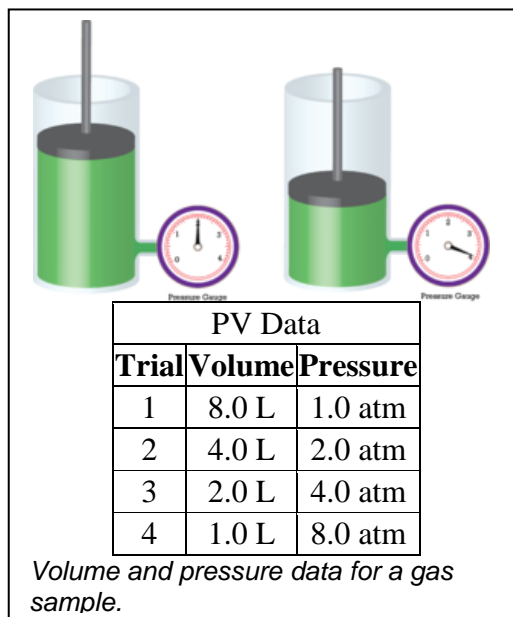
Boyle's Law can be summarized in the following equation:

$$P_1V_1 = P_2V_2$$

Where:

- P_1 =the initial pressure
- V_1 =the initial volume
- P_2 =the final pressure
- V_2 =the final volume

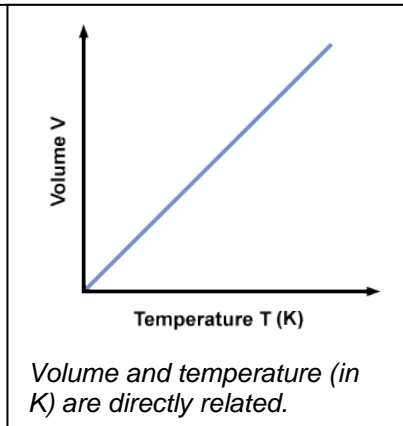
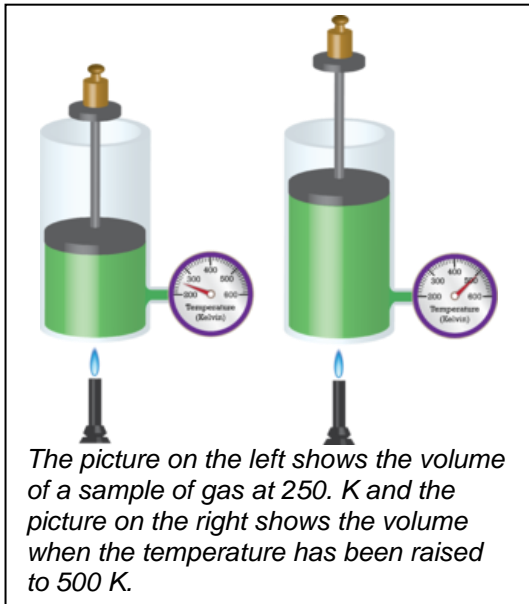
For this equation, the units used for pressure are unimportant, as long as both pressures have the same unit (either mmHg or atm) and each volume has the same unit (either mL or L).



11.2: Gas Laws

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). (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**.



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 direct 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:

$$\frac{V_1}{V_2} = \frac{T_1}{T_2}$$

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.

11.2: Gas Laws

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.

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

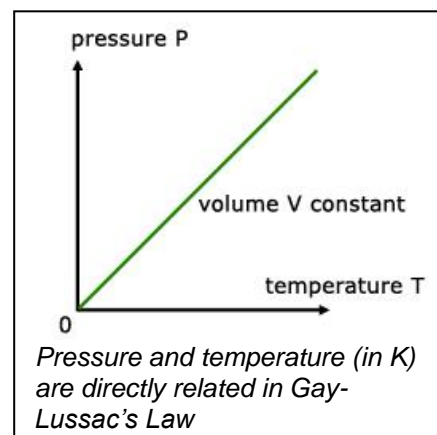
Temperature and pressure data.

Note that if the temperature doubles from 200. K to 400. K, the pressure also doubles.

After a series of temperatures and pressures have been measured, a data table like the others can be produced.

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.

$$\frac{P_1}{P_2} = \frac{T_1}{T_2}$$



This relationship is ONLY true if the temperature is measured in Kelvin. However, the units of pressure are irrelevant, as long as the two pressures are measured in the same units.

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

11.2: Gas Laws

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 form shown below and is known as the Combined Gas Law.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

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 temperatures out of Boyle's Law, you get:

$$P_1V_1 = P_2V_2$$

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.

$$P_1=300 \text{ mmHg}$$

$$V_1=400. \text{ L}$$

$$T_1=293 \text{ K (remember, ALL temperatures must be in Kelvin)}$$

$$P_2=760 \text{ mmHg (standard pressure)}$$

$$V_2=?$$

$$T_2=273 \text{ K}$$

Step 2: Solve the combined gas law for the unknown variable.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$
$$\frac{(300 \text{ mmHg})(400\text{L})}{293 \text{ K}} = \frac{(760 \text{ mmHg})V_2}{273\text{K}}$$

$$V_2=147 \text{ L}$$

11.2: Gas Laws

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.

$$\frac{(1.00 \text{ atm})(1.00\text{L})}{273 \text{ K}} = \frac{(2.00 \text{ atm})(1.50 \text{ L})}{T_2}$$

$T_2=819$ 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

$V_2=?$

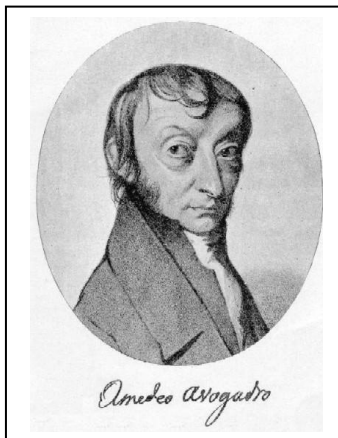
Temperature is constant, so it cancels out of the combined gas law.

Step 2: Solve the combined gas law for the unknown variable. (Or, recognize this is Boyle's Law and start with that equation.)

$$P_1V_1 = P_2V_2$$
$$(500 \text{ mmHg})(500 \text{ mL}) = (300 \text{ mmHg})V_2$$

$V_2=833$ mL

Avogadro's Law



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.

11.2: Gas Laws

The Ideal Gas Law

We have considered several laws that describe the behavior of gases. These relationships show how the volume, pressure, or temperature of a gas change as one of the other variables is kept constant. These patterns can be combined to form the **ideal gas law**. This law describes a gas at one point in time, instead of when conditions are changing.

$$PV = nRT$$

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.

$$P=1.50 \text{ atm} \quad V=5.56 \text{ L} \quad n=? \quad T=273 \text{ K (must be in K)}$$

Step 2: Solve the ideal gas law for the unknown variable.

$$PV = nRT$$

$$(1.50 \text{ atm})(5.56 \text{ L}) = n \left(0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}} \right) (273 \text{ K})$$

$$n=0.372 \text{ mol}$$

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.

$$P=? \quad V=500 \text{ mL} = 0.500 \text{ L} \quad n=2.00 \text{ mol} \quad T=100^\circ\text{C} = 373 \text{ K}$$

Step 2: Solve the ideal gas law for the unknown variable.

$$PV = nRT$$

$$P(0.500 \text{ L}) = (2.00 \text{ mol}) \left(0.0821 \frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}} \right) (373 \text{ K})$$

$$P=122 \text{ atm}$$

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Chapter 11 Summary

11.1: Kinetic Theory of Gases

- Molecules in gases spread out to take up the shape and volume of the container in which they are placed.
- Particles in a gas are very far apart compared to liquids and solids.
- The kinetic theory is used to explain the properties and behavior of gases. The kinetic theory states:
 - 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.
 - Temperature is directly proportional to the average kinetic energy of gas molecules.
 - Molecules of a gas are so far apart, on average, that the volume of the molecules themselves is negligible compared to the volume of the gas.
 - Molecular collisions with container walls cause the gas to exert pressure.
- Common units of measuring pressure include atmospheres (atm) and millimeters of mercury (mmHg). They are related by the following ratio: $1 \text{ atm} = 760 \text{ mmHg}$
- Because gases at the same temperature have the same kinetic energy, lighter gases will have higher velocities than heavier gases, at the same temperature.
- In the Kelvin scale, 0 K means the particles have no kinetic energy. Doubling the temperature in Kelvin doubles the kinetic energy of particles.

11.2: Gas Laws

- For a sample of a gas at constant temperature, volume is inversely proportional to pressure. As one of these variables doubles, the other is reduced to half its original value.
- For a sample of a gas at constant pressure, volume is directly proportional to temperature. Doubling the temperature in Kelvin doubles the volume and vice versa.
- For a sample of a gas at constant volume, pressure is directly proportional to temperature. Doubling the temperature in Kelvin doubles the pressure and vice versa.
- Standard conditions of temperature and pressure are 0°C and 1.0 atm.
- Avogadro's Law states that equal volumes of gases under the same conditions of temperature and pressure contain equal numbers of molecules.
- At STP, one mole of any gas occupies 22.4
- The Universal Gas Law: $PV=nRT$

Chapter 11 Summary

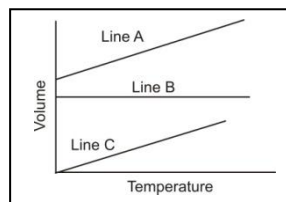
11.1: Kinetic Theory of Gases Review Questions

#1-5: Make each of the following unit conversions.

- 1) 2.5 atm to mmHg
- 2) 810 mmHg to atm
- 3) 75°C to Kelvin
- 4) 400 K to °C
- 5) 405 mL to L
- 6) Why is it necessary to use Kelvin temperatures in all gas law calculations that involve temperature?
- 7) If molecules of H₂, O₂, and N₂ are all placed in the same container at the same temperature, which molecules will have the greatest average kinetic energy? Which molecules have the greatest average speed?
- 8) If you have two balloons, one filled with CO₂ and the other with N₂, which molecules will have the greatest average kinetic energy? Which molecules have the greatest average speed?

11.2: Gas Laws Review Questions

- 9) When a sample of gas is placed in a larger container at the same temperature, what happens to the total pressure of the molecules hitting the walls? Explain why this happens in terms of kinetic theory.
- 10) 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?
- 11) A graph is made illustrating Charles's Law. Which line in the picture at right would be appropriate assuming temperature is measured in Kelvin?



Chapter 11 Summary

- 12) A sample of gas has its temperature increased from -43°C to 47°C at constant pressure. If its volume at -43°C was 500. mL, what is its volume at 47°C ?
- 13) A sample of gas has a volume of 500. mL at STP. Find its volume at 47°C and 800. mmHg.
- 14) A gas is confined in a rigid container and exerts a pressure of 250. mmHg at a temperature of 17°C . To what temperature must this gas be cooled in order for its pressure to become 216 mmHg? Express this temperature in $^{\circ}\text{C}$.
- 15) A sample of gas has a volume of 100. L at 17°C and 800 mmHg. To what temperature must the gas be cooled in order for its volume to become 50.0 L at a pressure of 0.83 atm?
- 16) 2.20 L of air at 1.10 atm is allowed to expand to fill a 6.30 L container. Find the final pressure.
- 17) A balloon is inflated to a volume of 1.25 L. If the temperature of the air inside is cooled from 35°C to 15°C , what will be the final volume of the balloon?

Chapter 11 Summary

- 18) A 5.25 L sample of oxygen gas at 720. mmHg and 35°C is compressed to a volume of 3.80 L while being heated to 78°C. What is the new pressure?
- 19) 2.0 liters of hydrogen is held in a rigid container at STP. If the temperature is dropped to -65.0°C , what is the resulting pressure in mmHg?
- 20) What is the volume of one mole of any gas at one atmosphere of pressure at room temperature 20.0°C ?

Glossary

Glossary

A

Absolute Zero:	the temperature at which molecules stop moving and therefore, have zero kinetic energy (See sections 5.1 and 11.1)
Alkali earth metals:	group 2A of the periodic table (See section 3.1)
Alkali metals:	group 1A of the periodic table (See section 3.1)
Alpha decay:	Alpha decay is a common mode of radioactive decay in which a nucleus emits an alpha particle (a helium-4 nucleus). (See section 10.1-10.2)
Alpha particle:	An alpha particle is a helium-4 nucleus, composed of 2 protons and 2 neutrons (See section 10.1-10.2)
Anion:	negative ion; formed by gaining electrons (See sections 2.3 and 4.3)
Anode:	The electrode at which oxidation occurs. (See section 9.3)
Arrhenius acid:	a substance that produces H ⁺ ions in solution (See section 8.1)
Arrhenius base:	a substance that produces OH ⁻ ions in a solution (See section 8.1)
Atom:	Democritus' word for the tiny, indivisible, solid objects that he believed made up all matter in the universe (See section 2.1)
Atomic mass unit (amu):	a unit of mass equal to one-twelfth the mass of a carbon-twelve atom (See section 2.4)
Atomic mass:	the weighted average of the masses of the isotopes of an element (See section 2.4)
Atomic number:	the number of protons in the nucleus of an atom (See section 2.3)
Avogadro's number:	The number of objects in a mole; equal to 6.02×10^{23} . (See section 5.5)

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. (See section 10.1, 10.4)
Balanced chemical equation:	a chemical equation in which the number of each type of atom is equal on the two sides of the equation (See section 7.4)
Battery:	A group of two or more cells that produces an electric current. (See section 9.3)
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. (See section 10.2)
Beta particle:	A beta particle is a high speed electron, specifically an electron of nuclear origin. (See section 10.2)
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 (See section 10.6)
Boiling point elevation:	the amount the boiling point of a solution increases from the boiling point of a pure solvent (See section 6.3)

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 (See section 7.2)
Cathode:	electrode at which reduction occurs (See section 9.3)
Cation:	positive ion; formed by losing electrons (See sections 2.3 and 4.3)
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. (See section 7.1)
Chemical reaction:	the process in which one or more substances are changed into one or more

Glossary

	new substances (See section 7.1)
Coefficient:	a small whole number that appears in front of a formula in a balanced chemical equation (See section 7.4)
Colligative property:	a property that is due only to the number of particles in solution and not the type of the solute (See section 6.3)
Combustion reaction:	a reaction in which oxygen reacts with another substance to produce carbon dioxide and water. (See section 7.5)
Compound:	a substance that is made up of more than one type of atom bonded together (See section 4.1)
Concentrated:	a solution in which there is a large amount of solute in a given amount of solvent (See section 6.2)
Concentration:	the measure of how much of a given substance is mixed with another substance (See section 6.2)
Controlled experiment:	An experiment that compares the results of an experimental sample to a control sample (See section 1.1)
Conversion factor:	a ratio used to convert one unit of measurement into another. (See section 5.4)
Cosmic background radiation:	energy in the form of radiation leftover from the early big bang (See section 10.6)
Covalent bond:	A type of bond in which electrons are shared by atoms. (See sections 4.2 and 4.5)
Covalent compound:	two or more atoms (typically nonmetals) forming a molecule in which electrons are being shared between atoms. (See sections 4.2 and 4.5)
D	
Dalton's Atomic Theory:	the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom (See section 2.1)
Decomposition reaction:	a reaction in which one reactant breaks down to form two or more products (See section 7.5)
Dilute:	a solution in which there is a small amount of solute in a given amount of solvent (See section 6.2)
Double replacement reaction:	a reaction in which two reactants form products by having the cations exchange places with the anions (See section 7.5)
Ductile:	can be drawn out into thin wires (See section 3.3)
E	
Electrochemical cell:	An arrangement of electrodes and ionic solutions in which a redox reaction is used to make electricity; a battery (See section 9.3)
Electrolysis:	A chemical reaction brought about by an electric current. (See section 9.3)
Electron configuration:	a list that represents the arrangement of electrons of an atom. (See section 2.6)
Electron:	a negatively charged subatomic particle, responsible for chemical bonding (See sections 2.2, 2.3, 2.5, 2.6, and 4.2)
Electronegativity:	the ability of an atom in a molecule to attract shared electrons (See sections 3.4 and 4.7)
Electroplating:	A process in which electrolysis is used as a means of coating an object with a layer of metal. (See section 9.3)
Electrostatic attraction:	The force of attraction between opposite electric charges. (See section 4.2)
Element:	a substance that is made up of only one type of atom Subatomic particles: particles that are smaller than the atom (See sections 2.1 and 4.1)
Endothermic:	reactions in which energy is absorbed, heat can be considered as a reactant (See section 9.2)
Equilibrium constant (K):	A mathematical ratio that shows the concentrations of the products divided by concentration of the reactants. (See section 7.8)
Equilibrium:	A state that occurs when the rate of forward reaction is equal to the rate of the reverse reaction. (See section 7.7)

Glossary

Equivalence point:	the point in the titration where the number of moles of acid equals the number of moles of base (See section 8.3)
Exothermic reaction:	A reaction in which heat is released, or is a product of a reaction. (See section 9.2)
Experiment:	A controlled method of testing a hypothesis. (See section 1.1)
Extrapolation:	the process of creating data points beyond the end of the graph line, using the basic shape of the curve as a guide (See section 1.2)
F	
Fission:	A nuclear reaction in which a heavy nucleus splits into two or more smaller fragments, releasing large amounts of energy. (See section 10.5)
Formula unit:	the empirical formula of an ionic compound; shows the lowest possible ratio (See section 4.4)
Freezing point depression:	the amount the freezing point of a solution decreases from the freezing point of a pure solvent (See section 6.3)
Fusion:	A nuclear reaction in which nuclei combine to form more massive nuclei with the simultaneous release of energy. (See section 10.5)
G	
Gamma ray:	the highest energy on the spectrum of electromagnetic radiation. (See section 10.1-10.2)
Graph:	a pictorial representation of patterns using a coordinate system (See section 1.2)
Group (family):	a vertical column in the periodic table, have similar chemical properties (See section 3.1)
H	
Half-life:	the time interval required for a quantity of material to decay to half its original value. (See section 10.3)
Halogens:	group 7A of the periodic table; reactive non-metals (See section 3.1)
Hydrocarbon:	an organic substance consisting of only hydrogen and carbon (See section 7.5)
Hypothesis:	A tentative explanation that can be tested by further investigation. (See section 1.3)
I	
Immiscible:	liquids that do not have the ability to dissolve in each other (See section 6.1)
Indicator:	a substance that changes color at a specific pH and is used to indicate the pH of the solution (See section 8.4)
International System of Units (SI Units)	the internationally agreed upon standard metric system (See section 5.1)
Interpolation:	the process of estimating values between measured values (See section 1.2)
Ion:	An atom or group of atoms with an excess positive or negative charge, lost or gained electrons (See sections 2.3 and 4.3)
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. (See sections 4.2-4.4)
Ionic compound:	a positively charged particle (typically a metal) bonded to a negatively charged particle (typically a nonmetal) held together by electrostatic attraction (See sections 4.2-4.4)
Ionic Formula:	includes the symbols and number of each ion (atom) present in a compound in the lowest whole number ratio (See sections 4.2-4.4)
Ionization energy:	the energy required to remove the most loosely held electron from a gaseous atom or ion (See section 3.4)
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 (See section 2.3)

Glossary

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. (See section 7.9)

M

Malleable: can be hammered into thin sheets (See section 3.3)

Mass number: the total number of protons and neutrons in the nucleus of an atom (See section 2.3)

Mass: a measure of the amount of matter in an object (See section 5.1)

Mendeleev: the Russian chemist credited with organizing the periodic table in the form we use today. (See section 3.1)

Metric system: international decimal-based system of measurement. (See section 5.1)

Miscible: liquids that have the ability to dissolve in each other (See section 6.1)

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 (See sections 4.1 and 6.1)

Molality: The ratio of the number of moles of solute per kilogram of solvent (See section 6.2)

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. (See section 5.5)

Molarity: the number of moles of solute per liter of solution (See section 6.2)

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 (See section 7.6)

Mole: An Avogadro's number of objects; 6.02×10^{23} particles (See section 5.5)

Molecular geometry: The specific three-dimensional arrangement of atoms in molecules. (See section 4.6)

N

Neutralization: a reaction between an acid and a base that produces water and a salt (See section 8.3)

Neutron: a subatomic particle with no charge (See sections 2.2-2.3)

Noble gases: group 8A of the periodic table; extremely non-reactive (See section 3.1)

Nuclear charge: the number of protons in the nucleus (See section 2.3)

Nucleus: the small, dense center of the atom (See sections 2.2-2.3)

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. (See section 4.2)

Oxidation: a loss of electrons, resulting in an increased charge or oxidation number (See section 9.3)

P

periodic law: states that the properties of the elements recur periodically as their atomic numbers increase (See sections 3.1, 3.4)

Periodic table: a tabular arrangement of the chemical elements according to atomic number. (See section 3.1-3.4)

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. (See section 7.1)

Polar covalent bond: A covalent bond in which the electrons are not shared equally because one atom attracts them more strongly than the other. (See section 4.7)

Potential energy: The energy of position or stored energy, including bond energy. (See section 9.1)

Products: materials present at the end of a reaction, shown on the right of the arrow

Glossary

	in a chemical equation (See section 7.3)
Proton:	a positively charged subatomic particle (See section 2.2-2.3)
R	
Reactants:	the starting materials in a reaction, shown left of the arrow in a chemical equation (See section 7.3)
Reduction:	gaining electrons, resulting in a decreased charge or oxidation number (See section 9.3)
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. (See section 5.2)
Single replacement reaction:	a reaction in which an element reacts with a compound to form products (See section 7.5)
Slope:	the ratio of the change in one variable with respect to the other variable. (See section 1.2)
Solute:	the substance in a solution present in the least amount, dissolved by the solvent (See section 6.1)
Solution:	a homogeneous mixture of substances (See section 6.1)
Solvent:	the substance in a solution present in the greatest amount (See section 6.1)
Stoichiometry:	the calculation of quantitative relationships of the reactants and products in a balanced chemical equation (See section 7.6)
Subscripts:	part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element (See section 4.2)
Synthesis reaction:	a reaction in which two or more reactants combine to make one product (See section 7.5)
T	
Temperature:	the average kinetic energy of the particles that make up a material (See sections 5.1 and 11.1)
Theory:	A well-established explanation based on extensive experimental data (See section 1.3)
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 (See section 8.4)
Transition elements:	groups 3 to 12 of the periodic table (See section 3.1)
V	
Valence electrons:	the electrons in the outermost energy level of an atom. (See section 3.3)
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. (See section 4.6)
W	
Weight:	the force of attraction between the object and the earth (or whatever large body it is resting on) (See section 5.1)

References

Polyatomic ions

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, $\text{C}_2\text{H}_3\text{O}_2^-$	Oxalate, $\text{C}_2\text{O}_4^{2-}$ Silicate, SiO_3^{2-} Thiosulfate, $\text{S}_2\text{O}_3^{2-}$	
Permanganate, MnO_4^-	Chromate, CrO_4^{2-} Dichromate, $\text{Cr}_2\text{O}_7^{2-}$	
Cyanide, CN^- Thiocyanate, SCN^-		

Conversion Factors

$$1 \text{ mole} = 6.02 \times 10^{23}$$

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)
.6214 miles (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

Equations

$$\text{slope} = \frac{\text{rise}}{\text{run}} = \frac{(y_2 - y_1)}{(x_2 - x_1)}$$

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

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

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{m}{V}$$

$$\text{pH} = -\log [\text{H}^+]$$

$$K_w = 1 \times 10^{-14} = [\text{H}^+][\text{OH}^-]$$

$$(M_a)(V_a) = (M_b)(V_b)$$