YOU Be The Chemist® CHALLENGE

2009–2010
STUDY GUIDE

CHEMICAL EDUCATIONAL FOUNDATION

CREATED BY THE CHEMICAL EDUCATIONAL FOUNDATION

Welcome to the *You Be The Chemist* Challenge!

Welcome to the *You Be The Chemist* (YBTC) Challenge! The Challenge is an exciting academic competition that will expand your knowledge of the science of chemistry.

The information in this Study Guide will expose you to the fascinating world of chemistry as it applies to your everyday life and demonstrate how chemicals will continue to shape your world. Read and review the information in the Study Guide, including all relevant examples to further your understanding of the material. Challenge competition questions are based on an understanding of the Study Guide material, often requiring you to apply the information to both familiar and unfamiliar situations.

How do I use this Study Guide?

The information in the Study Guide is supplied to help you succeed at every level of the Challenge. The Study Guide is divided into three categories – basic, intermediate, and advanced, which correspond to the three levels of the Challenge – local, state, and national. Information identified as "Basic" will be covered in Local Challenges, State Challenges, and the National Challenge. Information identified as "Intermediate" will be covered in State Challenges and the National Challenge. Information identified as "Advanced" will be covered in the National Challenge.[1] These divisions will be noted on the Study Materials page of the CEF Web site. If you are unsure of what material to study for a particular competition, ask your teacher or Local Challenge Organizer or contact the Chemical Educational Foundation at challenge@chemed.org.

We encourage you to explore all the material provided in the Study Guide regardless of whether you advance to State or National Challenge competitions. The more you explore, the more likely you are to find answers to the questions you have about the world around you.

The tips below will help you get the most out of the information provided and allow you to do so with ease.

1. First, read over the Table of Contents. This will introduce you to the concepts that will be covered.
2. Review the Objectives listed at the beginning of each section, so you are familiar with the topics you will be learning. Once you have completed your review of the entire section, go back and review the Objectives once again. Can you complete the tasks listed in the Objectives?
3. When you begin to read the sections, focus on the definitions of the **bold terms.** Then, follow up by reading the bulleted information.
4. **Diagrams and illustrations:** Use these images to gain a better understanding of the concepts.
5. Some of the more specific material is labeled as a **Quick Fact** (in circles). These are not essential to your understanding of the concepts, but they give applications and further details to help you understand the material even better. Although the **Quick Facts** are not essential to comprehend the basic concepts, they are still valid competition material. Be sure to study them too!

[1]The categories of Study Guide material and/or specific concepts covered in Challenge competitions may be adjusted at the discretion of CEF and of Challenge Organizers (with the approval of CEF). Students will be notified by their Local Challenge Organizer in the event of such changes.

6. **History boxes:** These boxes provide a variety of background knowledge about the Study Guide information. Whether highlighting a famous scientist who made an important discovery related to a concept or explaining previous beliefs and how those beliefs have changed as we have learned more about our world, the History boxes help to connect science of the past to science today.

7. **Element boxes:** These boxes provide details about specific elements that relate to the topics being discussed.

8. The circles labeled **Think About It** present questions related to the Study Guide material in a particular section. The answers to some **Think About It** questions may be obvious after reading the Study Guide material, but some answers may not even be known to scientists. These questions are placed in the Study Guide to make you think! Don't worry if the answer to a **Think About It** question is not obvious to you. Use these questions to explore more about chemistry and find out what questions scientists have or have not been able to answer. (Search the internet, chemistry books, or ask a scientist or teacher to see if you can find an answer.) Keep in mind that if an answer is not obvious, and you must explore other sources to find an answer, such a question will not be asked during a Challenge competition.

Once you are finished with a section, do a quick review to make sure you learned all the concepts introduced in that section. It might help to make flashcards too. If you find that you still do not understand something, pull out a science textbook and look in the index for the information or conduct a search on the internet (be sure to find a reliable source). If you find that the explanation you've found is inadequate or unclear, ask your science teacher for help.

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SECTION I:
SCIENCE – A WAY OF THINKING

OBJECTIVES

- Identify key components of the scientific method.
- Distinguish between different types of variables in a scientific study.
- Distinguish between theories and laws.

WHAT IS SCIENCE?

Science is the combined effort people make to better understand the world in which we live. It is a way of thinking about questions that has helped us gain knowledge about our world. You have probably thought like a scientist to solve problems and answer questions. In many cases, you probably began with a thought like "I wonder why this happens." Through science, we may be able to find an answer or at least understand more about something.

THE SCIENTIFIC METHOD

The **scientific method** is a process of thought and analysis that guides scientists as they ask and answer questions about the world. Different sources describe the stages of the scientific method in different ways, demonstrating that there is more than one way to explore science. Some sources describe the scientific method in four stages, and others describe it in five stages. However it is presented, any scientific method incorporates the same basic concepts as described in the sections below.

OBSERVATION

Begin by making an observation. Explore the world around you and collect information with your senses (smell, sight, sound, touch, and taste). Observations lead scientists to ask questions and design experiments to try to answer questions. The question a scientist asks explains the *purpose* of his/her experiment.

EXAMPLE:

You pour hot water into your ice cube tray and place the tray in the freezer. When you check the freezer a bit later, you notice *the hot water seemed to form ice cubes quickly*. As a result, you wonder whether *warm water or cold water freezes more quickly*.

RESEARCH

Conduct research. Someone may have already done some work to answer your questions, so look through books and other resources to find out what people already know. Many scientists spend hours reading papers on past research to prepare for their own research. Scientists should always study prior research to gain a better understanding of an observation or question they plan to explore before conducting new research. As they read, scientists are careful to look for gaps or errors in prior research.

HYPOTHESIS

Form a hypothesis. A **hypothesis** is a statement about an observed phenomenon that may explain or make a generalization about that phenomenon. Simply stated, a hypothesis can be considered an “educated guess” that could reveal something about your observation. A good hypothesis can be tested – it will guide your experiment.

EXAMPLE:

Hypothesis – *Warm water forms ice cubes in less time than cold water*.

***Notice that you can test this hypothesis by conducting an experiment to measure the time it takes for ice cubes to form, starting from water at different temperatures.**

EXPERIMENT

Design and conduct an experiment to explore your hypothesis. Be sure to design the experiment so that you are only testing one thing at a time (see section on Designing an Experiment). If you do not design the experiment carefully, your results may be confusing or your conclusions may be wrong.

ANALYSIS

Collect and analyze data. Data are pieces of information collected before, during, or after an experiment. Data must be analyzed to determine the results of the experiment. To understand the data, scientists often use mathematics or create graphs and charts.

When collecting data, try not to disturb the experiment. (If you bump a scale or balance when taking a measurement, you may get the wrong number.) Be sure to perform an experiment a few times to make sure you get the same results and can trust your data. Record your data somewhere safe, like in a laboratory notebook.

Quick Fact

Although people have begun to use data as a singular noun meaning “information,” data is actually a collective term. It is the plural form of “datum.”

CONCLUSION

Draw a conclusion. Once you have analyzed the data, you can make a statement that explains your observation and answers your question. The scientist can now make a statement about that observation with a level of confidence. The conclusion, however, is only as good as the data, so scientists must be able to rely on their data.

- If the data support the hypothesis, then the hypothesis is considered correct or *valid*.
- If the data do not support the hypothesis, then the hypothesis is considered incorrect or *invalid*. If the hypothesis is invalid, it must be *rejected or modified*, and the scientist must start from the beginning. Remember, there is nothing wrong with starting over. Scientists learn from both valid and invalid hypotheses.

COMMUNICATION

Communicate your findings. Tell people what you’ve learned by writing a report, giving a speech, or publishing an article. Scientific findings are rarely kept secret. (Remember, you used other scientists’ findings when you reviewed prior research.) By communicating your results, you are helping other scientists explain observations they have made.

DESIGNING AN EXPERIMENT

Properly designing an experiment can be a challenge. If you are not careful, you may end up testing too many things at once, which can prevent you from really learning anything.

When you design an experiment, you must identify the variables and the control.

VARIABLES

Scientists can test how changes to one item can affect something else. These changes are called **variables.** Two major types of variables are:

- **Independent variable:** the variable you are going to change in the experiment.
- **Dependent variable**: the variable you observe when the independent variable is changed.

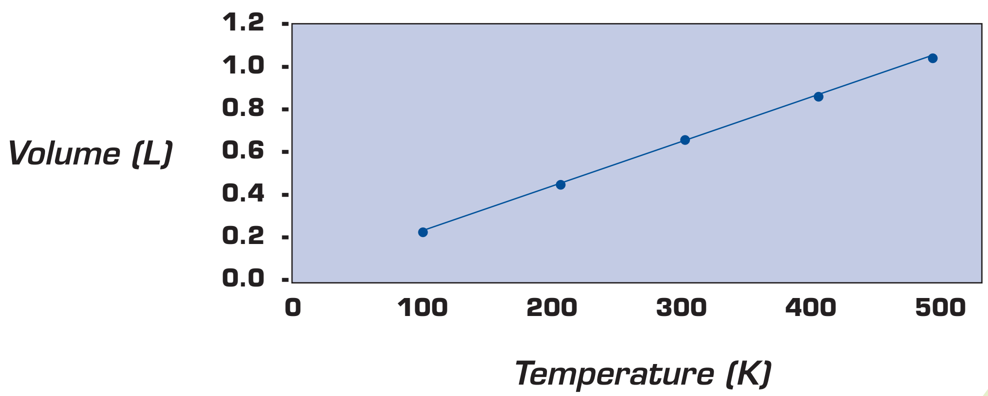
EXAMPLE:

A student wants to explore the relationship between the temperature of a gas and how much space it occupies. Each time the temperature is changed during the experiment, the student measures the new volume of the gas.

The temperature is the independent variable. The volume is the dependent variable.

The data you collect are often plotted on a graph. The independent variable is plotted on the x-axis (horizontal), and the dependent variable is plotted on the y-axis (vertical). The graph below shows the relationship between the temperature and volume of a gas:

VOLUME VS. TEMPERATURE FOR A GAS



CONTROL

A **control** is something that is not being changed throughout the experiment. Controls are used to make comparisons.

EXAMPLE:

A student would like to know what type of paint, if used on the outside of a box, will heat up the inside of the box the fastest when placed in the sun.

***The control in this experiment would be a box with no paint on the outside.* Following the experiment, you would compare the results from the control box to all of the other boxes with paint.**

Similarly, **control variables** are potential variables that are not changed during an experiment. There may be multiple control variables.

From the example above, *a control variable would be the type of box used in the experiment*. If one box was made of wood and another was made of metal, the type of box would not be controlled, and the experiment would be testing both the type of box and the type of paint on the box.

THEORIES AND LAWS

Even after a scientist makes a conclusion and communicates it, other steps may be taken. Over time, scientists can conduct similar experiments to collect more data. The combined results of these experiments can lead to the development of a theory or law.

THEORY: an explanation based on observations and supported by data.

- A theory is not a fact and may change as more information becomes available.
- Most theories have been tested over and over, producing the same results.
- Theories generally explain "why" something occurs and can be used to predict future observations.
- Examples of theories include the Theory of Relativity, the kinetic theory of gases, and the Big Bang Theory.

LAW: a description of a natural phenomenon shown to occur again and again under certain conditions.

- Laws are usually accepted as true and universal because observations and data have always shown them to be true.
- Scientific laws can be challenged and possibly disproven. Though disproving a scientific law is very rare.
- Laws generally state that a particular phenomenon will always occur if certain conditions are present.
- Many scientific laws can be expressed in terms of a single mathematical equation.
- Examples of laws include the Law of Conservation of Energy, Newton's laws of motion, and the Law of Gravity.

SECTION II: MEASUREMENT

OBJECTIVES

- Demonstrate the differences between accuracy and precision.
- Identify common physical properties.
- Explain the difference between mass and weight.
- Work with metric prefixes.
- Perform simple conversions between metric units.
- Write large and small numbers in scientific notation.

Scientists learn by making observations and taking measurements. Some observations are simple, like identifying an object's color or texture, but if scientists want to know more about a substance, they may need to take measurements.

CERTAINTY IN MEASUREMENT

Two goals of measurement:

- **Accuracy**: to get as close as possible to the true measurement (true value) of something.
- **Precision**: to be able to perform the same measurement and get the same result over and over.

EXAMPLE:

A scientist measures a ball to determine its mass. The scientist finds that the ball has a mass of 15 grams.

If 15 grams is the true value of the mass of the ball, then it is said to be an accurate measurement. If the scientist measured the ball 3 times and got a mass of 14 grams each time, the measurement is precise but not accurate.

NOTES

TYPES OF PHYSICAL MEASUREMENTS

Scientists perform measurements to learn many different things, such as how long an object is, how much space it takes up, and more. Some common types of measurements used by scientists are:

MASS: a measure of the amount of matter in a substance, or simply stated, the amount of "stuff" in a substance (see Classification of Matter section).

- Mass is usually measured with a balance or an electronic scale. To determine the mass of an object, the object is compared to another object with a known mass.
- The unit of measure that scientists use to measure mass is the kilogram (kg) or gram (g).

Quick Fact

Mass vs. Weight:
Most of us know how much our body weighs, but this measure is different from our body's mass. Weight is actually a unit of force that depends on other forces like gravity. The gravity on the earth gives us a certain weight. Other places, like the surface of the moon, have a much lower gravity, so a person's weight on the moon would be much lower (one-sixth of his/her weight on the earth). However, a person's mass does not depend on gravity and therefore, doesn't change with location. It is the same on the earth and on the moon.

WEIGHT ON OTHER PLANETS

The table below shows what a person who weighs 100 pounds (and has a mass of 45.4 kg) on Earth would weigh on other planets:

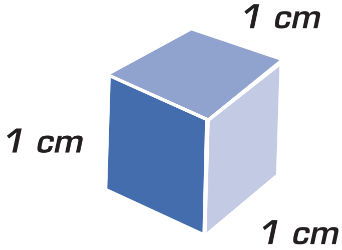
Planet	*Mercury*	*Venus*	*Mars*	*Jupiter*	*Saturn*	*Uranus*	*Neptune*
Weight (in pounds)	38	91	38	254	93	81	120

LENGTH: a measure of how long an object is or the distance an object spans.

- Length may be measured using a meter stick, a ruler, or other more complicated instruments, depending on the object. Some objects, such as DNA molecules, are so small that they have to be measured with light.
- Scientists use the meter (m) as the standard unit for measuring length.

VOLUME: the amount of space an object occupies.

- Volume can be measured in many different ways (see Laboratory Equipment section).
- Scientists measure volume in units of liters (L) or cubic centimeters (cm^3 or cc). One cubic centimeter is equal to one milliliter (mL).



TEMPERATURE: a measure of the average energy of particles of a substance or how fast the particles are moving around.

- Temperature is generally measured using a thermometer.
- Scientists typically measure temperature in units of Celsius or Kelvin.
- The range from freezing point to boiling point on the Celsius scale is 0°C to 100°C. The Kelvin temperature scale is actually an extended Celsius scale that includes absolute zero (0 K). To convert a degree Celsius temperature into Kelvin, just add 273.
- Degrees Fahrenheit, the temperature scale used commonly in the United States, ranges from a freezing point of 32°F to a boiling point of 212°F. Degrees Fahrenheit can also be calculated from degrees Celsius.

Quick Fact

Most people think of temperature as a measurement of hot and cold. To scientists, temperature measures the kinetic energy (energy of motion) of molecules in a material. Molecules with high kinetic energy have a high temperature. They are said to be hot. Molecules with low kinetic energy have low temperatures and are said to be cold.

FORMULAS:

$K = °C + 273$ — **K = Kelvin: the metric scale for absolute temperature**

$°F = (1.8 \times °C) + 32$ — **°F = degrees Fahrenheit: the engineering scale for temperature**

$°C = \dfrac{°F - 32}{1.8}$ — **°C = degrees Celsius/Centigrade: the metric scale for temperature**

DENSITY: describes the amount of matter contained by a given volume, or simply stated, the amount of "stuff" that is crammed into a certain amount of space. An object that is very dense has a lot of stuff crammed into a small amount of space.

- Density gives the relationship between two measurements: mass and volume.

$$D = \frac{m}{v}$$

- The units of measurement that chemists usually use to measure density are grams per milliliter (g/mL) or grams per cubic centimeter (g/cc).

Quick Fact

Density changes with temperature and is often reported with the temperature at which it was measured. For example, at 4°C, pure water has a density of 1.0 g/mL.

In general, when temperature increases, the atoms in a substance move farther apart, and the substance becomes less dense. When temperature decreases, the substance becomes denser. Water, however, is an exception. The density of an ice cube at 0°C is 0.917 g/mL.

EXAMPLE:

Gold has a density of 19.4 g/cc. This means that 19.4 grams of gold occupy 1 cubic centimeter or that 1 cubic centimeter of gold has a mass of 19.4 grams.

On the other hand, liquid water has a density of only about 1 g/mL, which means that 1 milliliter of water has a mass of 1 gram.

Quick Fact

A substance with a density greater than pure water will sink. A substance with a density less than pure water will float.

PRESSURE: the amount of force exerted on an area. A **force** is the amount of push or pull on an object.

$$p = \frac{F}{A}$$

- Many forces affect people and objects on the earth. They include gravity, magnetism, and the force exerted on a surface by the weight of the air above it (air pressure). These same forces act on objects under water.
- Pressure is usually measured with a barometer or a manometer.
- Common units of measure that scientists use for pressure are Pascals (Pa), atmospheres (atm), torrs (torr), or millimeters of mercury (mmHg).

ENERGY: a measure of the ability to do work or generate heat.

- Energy cannot be measured directly because it is defined as the work that one system does (or can do) on another. Therefore, only *change* in energy can be measured.
- Some common energy units that scientists use to express measures of energy are Joules (J), calories (cal), and electron volts (eV).

Quick Fact

Atmospheric or air pressure is exerted on a surface by the weight of the air above that surface.

Quick Fact

Pressure on the earth changes with altitude. At sea level, the air pressure is about 1 atm or 14.7 pounds per square inch. This means that every square inch of our bodies has almost fifteen pounds (the weight of a heavy bowling ball) pushing on it. At higher altitudes, air pressure drops, so people living in high places will have less pressure on them.

Quick Fact

A calorie is not a physical object; it is a unit of energy. Most people are familiar with nutritional calories (food calories). A food with a lot of calories is able to supply a lot of energy, while a low-calorie food cannot. The more work a person does (like marathon runners), the more calories they need in their food.

Food calories and scientific calories are measured differently. Each calorie listed on a Nutrition Facts food label is actually a kilocalorie (kcal) or 1,000 scientific calories! That's a lot of energy!

MEASUREMENT

The table below summarizes the measurements and corresponding units described on the previous pages.

Measure	What It Measures	Scientific Units	Common Units
Mass	Amount of matter in an object	Kilograms (kg), grams (g)	Pounds, stones
Length	How long an object is or the distance between objects	Meters (m)	Feet, inches, miles
Volume	Amount of space an object occupies	Liters (L), cubic meters (m^3), cubic centimeters (cm^3 or cc)	Gallons, pints, quarts
Temperature	Average energy of particles	Celsius (°C), Kelvin (K)	Fahrenheit (°F)
Density	Ratio of mass to volume for an object	Kilograms per cubic meter (kg/m^3), grams per cubic centimeter (g/cc), grams per milliliter (g/mL)	Pounds per cubic inch
Pressure	The amount of force per unit area	Pascals (Pa), atmospheres (atm), torrs (torr), millimeters of mercury (mmHg)	Pounds per square inch
Energy	The ability to perform work or generate heat	Joules (J), calories (cal), electron volts (eV)	British Thermal Unit (BTU)

In general, physical measurements can be classified in two different groups:

- **Extrinsic properties**: properties that change based on the amount of substance present. Mass is an example of an extrinsic property. The larger an object, the more mass it has.
- **Intrinsic properties**: properties that do not change based on the amount present. Density is an example of an intrinsic property. Pure water in a bucket and pure water in a large pool will have the same density.

UNITS OF MEASUREMENT

In general, scientists use a system of measurement known as the metric system. The metric system was the first standardized system of measurement. It was developed in France in the 1790s. Before this, people relied on many different systems of measurement, such as the “fps” system (foot, pound, second).

In 1960, the metric system was revised, simplified, and renamed The Système International d’Unites (International System of Units) or SI system. This system is used in almost every country around the world, except for the United States, which uses the English System (inches, quarts, etc.). The SI system is the standard system used by scientists worldwide, including those in the United States and allows scientists to easily convert or switch between large and small numbers. The SI system is still commonly called the metric system.

In the SI system, each unit of measure has a base unit. The seven base units of the SI system are:

Measure	Base Unit
Length	Meter (m)
Mass	Kilogram (kg)
Time	Second (s)
Temperature	Kelvin (K)
Amount of substance	Mole (mol)
Electric current	Ampere (A)
Luminous intensity	Candela (cd)

Of these base units, the first three (meter, kilogram, second) are called the *primary units*.

When taking measurements, some things may be very large and other things may be very small. To work with large and small numbers, scientists use metric prefixes. The table below lists some common prefixes and the quantity they represent:

Prefix	Symbol	Numerical Value
Tera-	T	10^{12} (1,000,000,000,000)
Giga-	G	10^{9} (1,000,000,000)
Mega-	M	10^{6} (1,000,000)
Kilo-	k	10^{3} (1,000)
Hecto-	h	10^{2} (100)
Deca-	da	10^{1} (10)
(no prefix)	--	10^{0} (1)
Deci-	d	10^{-1} (0.1)
Centi-	c	10^{-2} (0.01)
Milli-	m	10^{-3} (0.001)
Micro-	µ	10^{-6} (0.000001)
Nano-	n	10^{-9} (0.000000001)
Pico-	p	10^{-12} (0.000000000001)

Prefixes are added to base units to make the base unit larger or smaller in value. For example, one kilometer is 1,000 meters. One millimeter is 0.001 meters.

- **Multipliers**: prefixes greater than one, such as deca-, kilo-, and giga-.
- **Fractions**: prefixes less than one, such as deci-, milli-, and nano-.

Additional prefixes have been added over the years as new scientific instruments have allowed scientists to measure even smaller and larger amounts, such as zepto (10^{-21}) and yotta (10^{24}).

EXAMPLE:

If someone were to line up 1,000 meter sticks end to end, the resulting measurement would be 1,000 meters or 1 kilometer. The kilometer unit is useful for describing long measurements like the distance between Chicago and New York City (approximately 1,300 kilometers).

Something very small, such as the width of a human hair, is only about 110 millionths of a meter. A micrometer is 1 millionth of a meter, so a human hair is about 110 micrometers (or 110 microns).

Other SI units have been derived from the seven base units. The table below lists some common derived units:

Measurement	Derived Unit	Base Units	Derived Unit Symbol
Volume	Liter	10^{-3} x m^3	L
Force	Newton	kg x m/s^2	N
Energy, work	Joule	N x m	J
Pressure	Pascal	N/m^2	Pa
Power	Watt	J/s	W
Electric potential	Volt	W/A	V
Resistance	Ohm	V/A	Ω
Frequency	Hertz	1/s	Hz

CONVERTING METRIC UNITS

Because scientists measure things that can be very large or very small, they need to be able to convert between these units of measure quickly. The metric system makes conversion simple because prefixes are based on groups of ten.

EXAMPLE:

Runners completing an 8 kilometer race, would have run how many millimeters?

$$8\ \cancel{km} \times \frac{1{,}000{,}000\ mm}{1\ \cancel{km}} = 8{,}000{,}000\ mm$$

Quick Fact

Here's a trick to converting metric units:

To change from one prefix to another, subtract the exponent for the first prefix from the exponent for the second prefix. Use the number you get to move the decimal point that number of places to the right or left, as appropriate. Finally, fill in with zeros, if necessary.

For example, to change centimeters (10^{-2}) to millimeters (10^{-3}), the difference between the exponents is one. Therefore, you would move the decimal one place to the right.

1.0 cm = 10.0 mm
1 centimeter = 10 millimeters

SCIENTIFIC NOTATION

Because scientists often work with very large and very small numbers, they need an easy way to write these numbers. *Scientific notation* is the method used to write very large and very small numbers.

STEPS FOR LARGE NUMBERS

The speed of light in a vacuum is roughly 300,000,000 m/s. Instead of writing all those zeros, scientists use scientific notation.

1. Write the number as a multiplication problem.

   300,000,000 *m/s* = 3 x 100,000,000 *m/s*

2. Write the last number above (one followed by eight zeros) as an exponent using the number 10. For example, $10^2 = 100$ or $10^3 = 1{,}000$.

3. Rewrite the number using steps 1 and 2.

   300,000,000 *m/s* = 3 x 100,000,000 *m/s* = $\boldsymbol{3 \times 10^8}$ ***m/s***

Other numbers can be written in scientific notation. For example, the number of feet in a mile (5,280) would be:

5,280 *ft* = 5.28 x 1,000 *ft* = $\boldsymbol{5.28 \times 10^3}$ ***ft***

With numbers like the one above, scientists write one number before the decimal point and all other numbers after the decimal point.

Quick Fact

Scientific notation can be as easy as counting. Move the decimal in one direction or the other. Then count the number of places moved to figure out the correct exponent.

Using the feet per mile example – 5,280 feet:

1. Move the decimal to the left, so it appears immediately after the first number – 5.28.

2. Since you moved the decimal three places, the exponent is written as 10^3.

3. Therefore, 5,280 feet is written as 5.28×10^3.

STEPS FOR SMALL NUMBERS

Small numbers can be expressed in scientific notation too. For example, certain sensors can detect radiation on the far infra-red portion of the electromagnetic spectrum with wavelengths of about 0.00000110 meters.

Using the decimal method from the Quick Fact above …

1. Move the decimal to the right, so it appears immediately after the first number.

   0.00000110 → 000001.10

2. Count the number of places you moved the decimal to determine the exponent for the 10.

   6 places → $\mathbf{10^6}$

3. Make the exponent a negative number. (The negative exponent indicates the number is very small.)

   $\mathbf{10^6 \rightarrow 10^{-6}}$

4. Write the number as a multiplication problem.

   $\mathbf{000001.10 \times 10^{-6} = 1.10 \times 10^{-6}}$

SECTION III: CLASSIFICATION OF MATTER

OBJECTIVES

- Classify matter as either a pure substance or a mixture.
- Classify matter into types of mixtures, compounds, and elements.
- Identify the common phases of matter and their characteristics.
- Differentiate between physical and chemical changes.
- Distinguish between phase changes.
- Explain energy as it relates to physical and chemical changes.

Simply stated, chemistry is the study of matter and changes in matter. What is matter? **Matter** is defined as anything that has mass and takes up space.

Matter is characterized by its physical and chemical properties. **Physical properties** are characteristics that can be observed with our senses or measured with instruments. Physical properties include color, shape, boiling point, melting point, and density. **Chemical properties** are observed when substances react with each other, undergoing a chemical reaction (see section on Chemical Changes). Chemical properties include toxicity and flammability.

TYPES OF MATTER

Matter can be divided into two main categories:

1. **Pure substance:** something that has a structure and properties that do not change.
   - Pure substances cannot be separated into other substances by physical means. Water and carbon dioxide gas are examples of pure substances.
   - Pure substances can be classified as either elements or compounds.
   - Pure substances have the same composition (makeup) throughout.
2. **Mixture:** two or more substances that are physically combined.
   - Different parts of the mixture have different properties. Tossed salad is a mixture because it is made of different parts, such as lettuce, carrots, and dressing.
   - When two substances are combined to form a mixture, no chemical reaction occurs. Each part of the mixture keeps its original properties. This allows scientists to separate mixtures back into the original parts (see section on Physical and Chemical Separations).
   - A mixture can be classified as either homogeneous or heterogeneous.
   - Mixtures have random compositions.

Quick Fact

The words homogeneous and heterogeneous might be easier to understand by exploring the meanings of the words themselves. The Greek prefix “homo” means “same,” and the prefix “hetero” means “different.” The suffix “genous” means “produced by or composed of.”

PURE SUBSTANCES

Element: a pure substance that cannot be broken down further by ordinary chemical means. An element is made of only one type of atom, distinguished by its atomic number (see Atomic Structure section).

- **Allotropes**: different forms of the same element that occur in the same physical state of matter.
  - Two solid allotropes of carbon that you may recognize are graphite (in pencils) and diamond.
  - Oxygen has two gaseous allotropes – O_2 (diatomic oxygen) and O_3 (ozone).

O

OXYGEN

Atomic #8

Characteristics:

- Makes up two-thirds of the human body by weight, mostly in the form of water
- Is the most common element in the earth's crust and the 3^{rd} most abundant element in the sun
- Is essential for fermentation and for oxidation processes, such as respiration, combustion, and rusting
- Has two gaseous allotropes:
  1. *Diatomic oxygen* (O_2) – a colorless and odorless gas that makes up 21% of air by volume.
  2. *Ozone* (O_3) – formed by electrical discharges (lightning) or ultra-violet (UV) radiation acting on oxygen in the atmosphere. Ozone forms a thin protective layer in the atmosphere that helps block UV radiation from reaching the earth's surface.

Most elements combine with oxygen to form compounds with names ending in "-ate" or "-ite," such as nitrites ($-NO_2$) and sulfates ($-SO_4$).

Compound: a pure substance made up of two or more elements bound together in a defined ratio.

- Water is a compound of hydrogen and oxygen atoms bound together in a 2:1 ratio (2 hydrogen atoms to 1 oxygen atom – H_2O). Other familiar compounds include table salt (NaCl), glucose ($C_6H_{12}O_6$), and battery acid (H_2SO_4).
- There are approximately one million different known compounds. Chemists develop and study new compounds every day.
- A **binary compound** is composed of two elements (NaCl, CO_2). A **tertiary compound** is composed of three elements ($C_6H_{12}O_6$).

HISTORY: **JOHN DALTON**

John Dalton (1766-1844) developed the atomic theory of matter. This theory became the basis for all future models of the atom. Dalton's four-part atomic theory states that:

- All matter is composed of indivisible particles called atoms.
- All atoms of a given element are identical. Atoms of different elements have different properties.
- Chemical reactions involve the combination of atoms, not the destruction of them.
- When elements react to form compounds, they react in defined, whole-number ratios.

Dalton's studies on gases led to the development of the law of partial pressures. This law is also known as Dalton's Law.

- **Dalton's Law**: the total pressure of a mixture of gases equals the sum of the pressures of the gases in the mixture, with each gas acting independently.

$$P_{total} = P_1 + P_2 + P_3 + \ldots$$

MIXTURES

Homogeneous mixture: a type of mixture that is considered to be the same throughout. Apple juice is a homogeneous mixture because the juice at the top of your glass is the same as the middle and the bottom. Any sip of the juice should taste the same.

- Homogeneous mixtures made of metals are called *alloys* (see section on Alloys).
- In liquid or gas form, homogeneous mixtures are usually called *solutions*.
  - **Solution**: a homogeneous mixture in which one or more substances (solutes) are dissolved in another substance (solvent). Solutions consist of elements or compounds mixed together at a very fine level. For example, dissolving salt into water creates a salt-water solution.
  - **Solvent**: the substance of greater quantity in the solution. In saltwater, the solvent is water.
  - **Solute**: the substance of less quantity in the solution. In saltwater, the solute is salt.

Heterogeneous mixture: a type of mixture in which the composition is not uniform. It is not finely mixed and does not have the same makeup throughout. For example, beach sand is a heterogeneous mixture. If you look closely, you can see different colors from the different substances that make up the sand. No two handfuls of sand from a beach will be exactly the same.

Quick Fact

Sodas are homogeneous mixtures made up of various solutes, including sugar and carbon dioxide gas, dissolved in water.

Quick Fact

If both the solute and solvent exist in equal quantities (such as in a 50% ethanol, 50% water solution), the substance that is more often used as a solvent is designated as the solvent (in this case, water).

Water is often referred to as a "universal solvent" because it dissolves more substances than any other liquid.

Some mixtures are not classified simply as homogeneous or heterogeneous.

Colloid: a mixture in which very small particles (of about one micrometer in size or less) are spread evenly through another. Because of the tiny size of the particles, some colloids have the appearance of solutions.

- A colloidal mixture is somewhere between homogeneous and heterogeneous.
- Colloids consist of fine particles of one substance mixed into another.
- **Aerosols**: colloidal suspensions of liquid or fine solid particles in a gas.
  - Smoke is an aerosol with solid particles.
  - Fog, mists, clouds, and sprays are aerosols with liquid particles.
- **Sols**: colloidal suspensions of fine solid particles in a liquid or another solid.
  - Paints, muddy river water, and sewage are liquid sols.
  - Pearls, opals, and pigmented plastics are solid sols.
- **Gels**: consist of liquids spread throughout a solid, such as jelly, butter, and cheese.
- **Foams**: consist of gases finely spread throughout liquids or solids.
  - Whipped cream and soda foam are liquid foams.
  - Marshmallows and Styrofoam™ are solid foams.
- **Bubble**: a single ball of gas in a liquid.
- **Emulsions**: consist of liquids spread throughout other liquids, such as oil and vinegar salad dressing, hand cream, and mayonnaise.
  - Emulsions often have a cloudy appearance because the boundary between components (boundary between water and oil) scatters the light that passes through the sample.
  - Emulsions are often unstable. Homemade oil and vinegar salad dressing is an unstable emulsion that will quickly separate unless shaken continuously. This occurrence is called **coalescence** – when small droplets recombine to form bigger ones.

Quick Fact

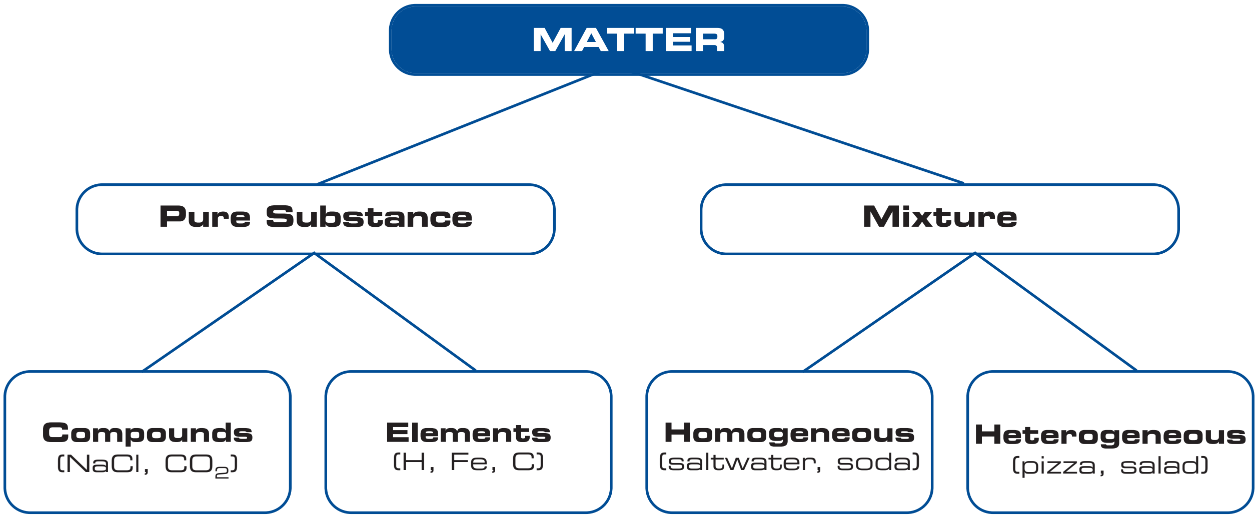
The Tyndall effect is the scattering of light that occurs when particles spread throughout liquid and gas colloids are large enough to scatter the light but small enough that they do not settle. It is the reason headlight beams are visible on a foggy night.

Quick Fact

Sometimes, it is difficult to determine whether a mixture is heterogeneous or homogeneous. In the ocean, the amount of salt in the water may be different at different depths. Scientists would not say the ocean is completely homogeneous. A water sample taken from 50 feet below the surface is not exactly the same as water taken from 200 feet below the surface. Can you think of other substances that may be difficult to classify as either heterogeneous or homogeneous mixtures?

What about milk … is it homogeneous or heterogeneous?

The following flow chart shows the different types of matter:



NOTES

STATES OF MATTER

Matter can be classified based on its state. The most familiar states of matter are solids, liquids, and gases. For example, ice, water, and steam are all H_2O in different states. As heat is added or removed from one state of matter, the matter may undergo a phase change.

SOLIDS: have a definite volume and a definite shape. A brick, a table, and a pencil are examples of solids.

- Solids are generally denser than liquids and gases. The molecules that make up a solid are usually locked into place.
- **Crystalline solids**: solids made up of atoms or molecules that are organized in specific repeating patterns.

  - Crystalline solids have a definite melting point, so they change rather quickly from a solid to a liquid when heated above their melting point.
  - Diamonds, ice, and table salt are crystalline solids.
- **Amorphous solids**: solids with atoms or molecules that are locked into place but without a specific form or neat repeating structure.
  - Most amorphous solids do not have a definite melting point. Instead, they soften and become more flexible when heated.
  - Glass is an example of an amorphous solid. It is made up of silicon dioxide (SiO_2) but does not have a repeating structure.
  - Wax and many polymers, including polystyrene and most plastics, are amorphous solids (see section on Polymers).

LIQUIDS: have a definite volume but no definite shape. Water and milk are liquids.

- A liquid will take the shape of its container, but it cannot change its overall volume.
- The molecules that make up liquids are in constant random motion.
- Even though liquid molecules move around freely, they still experience weak attractive forces. These forces cause the molecules to remain fairly close to one another. Multiple weak bonds or forces can result in strong interactions.

Quick Fact

Sugar (sucrose) can exist as an amorphous solid, such as when it forms cotton candy, or it can exist as a crystalline solid, such as the small crystals that are generally found in people's kitchen sugar bowls.

Think About It...

Knowing that rock candy is crystalline, and cotton candy is amorphous, how would you classify chocolate candy bars and lollipops?

Chocolate bars are actually crystalline solids made up of crystalline cocoa butter, while lollipops and hard candies are actually glassy amorphous solids! Now, what do you think about other candies – Tootsie Rolls®, saltwater taffies, or M&Ms®?

- **Viscosity** is a property of liquids that describes the "thickness" of the material. It is a measure of the liquid's resistance to flow. The less viscous a liquid is (the lower its viscosity), the more easily it flows.
  - The ability of a material to flow is related to the strength of the bonds between molecules of the liquid. The stronger the bonds, the more difficult it is for the liquid to flow (see Chemical Bonds section).
  - Water has a low viscosity, while honey has a high viscosity.
- **Surface tension** is a property of liquids that describes the tendency of liquid molecules to be attracted to one another at the surface. The strong attraction of molecules at the surface of the liquid creates a surface "film," which makes moving an object through the surface of a liquid more difficult than moving the object when it is completely submerged in the liquid.
  - Surface tension causes liquids to keep a low surface area, For example, soap bubbles and rain droplets form a spherical shape rather than spreading out flat.
  - Water has a very high surface tension (~0.073 N/m) because of strong hydrogen bonding (see Chemical Bonds section). Oil has a much lower surface tension than water because its molecules are weakly attracted to each other. If you pour a spoonful of vegetable oil on a plate and then pour a spoonful of water on another plate, the water does not spread as far as the oil because its molecules are strongly bound together.

Quick Fact

Even though a sewing needle is denser than water, when the needle is placed carefully on its side, it will remain on the water's surface because of surface tension.

Quick Fact

Popular belief is that the water strider bug can "walk" on water because of surface tension. However, new research has shown that, in actuality, the water strider's legs are covered with microscopic hairs that trap tiny air bubbles, allowing the bug to "float" across the water's surface.

GASES: have no definite volume and no definite shape. The air around you is a mixture of gases.

- If a gas is put into a container, it will occupy the entire container and take the shape of the container.
- Gas molecules move about quickly in random directions because they do not have strong bonds or attractions between them.
- Gas molecules will not remain close to one another for any length of time. Because of the extra space between the molecules, gases are easily compressed.
- **Effusion** is the process by which gas molecules flow through small holes in a container.
  - **Molecular mass** is the mass of one molecule of a substance.
  - According to Graham's law, gases with a lower molecular mass effuse more quickly than gases with a higher molecular mass, just like it is easier for a smaller person to squeeze through a tight space than it is for a larger person.

Quick Fact

A balloon filled with gas will deflate over time as a result of effusion.

EXAMPLE:

Two balloons are filled to the same size with different gases – one with oxygen gas and the other with hydrogen gas. When left alone, the hydrogen gas balloon deflates faster because hydrogen gas molecules have a smaller molecular mass and can escape (effuse) more quickly than the oxygen molecules.

- **Diffusion** is the movement of molecules from an area of high concentration to an area of low concentration as a result of the molecules' kinetic energy (see section on Energy Changes).
  - Diffusion continues as the molecules move at random until all of the molecules are spread evenly throughout the area.
  - The molecules of fluid substances can undergo diffusion.

EXAMPLE:

If you add a drop of red food coloring to a glass of water, the food coloring will spread through the water, until the water is evenly colored red.

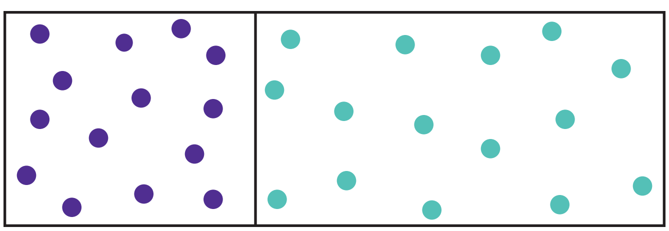
Likewise, if you walk to school and smell the odor of skunk spray, you most likely won't smell the odor on your way home from school. As a result of diffusion (and air movement – wind), the odor becomes weaker over time as the skunk spray spreads through the air.

Quick Fact

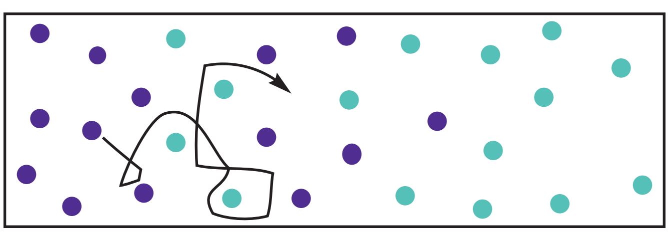
A **fluid** is any substance made up of molecules that flow or move freely. A fluid easily changes shape when a force is applied.

Liquids and gases are fluids; plasmas are fluids as well.

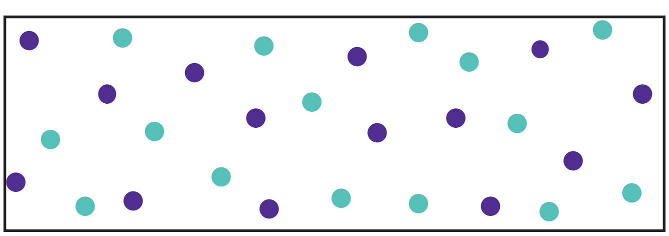
The process of diffusion is illustrated below:



The molecules of two fluids are separated by a barrier.



With the barrier removed, the molecules begin to diffuse to the other side. The arrow shows the random motion of a molecule from a higher concentration to a lower concentration.



The two fluids have reached equilibrium – they are evenly spread throughout the container.

– Certain factors can affect the rate of diffusion, such as temperature, concentration, and the size of the molecules.

- In general, molecules will diffuse more quickly at higher temperatures. Therefore, a drop of food coloring will diffuse more quickly in a cup of hot water than in a cup of cold water.
- Smaller molecules diffuse more quickly than larger molecules.
- Molecules in a highly concentrated area will diffuse quickly to areas that are less concentrated.

Quick Fact

Osmosis is the diffusion of water across a semi-permeable membrane (a membrane that allows some ions or molecules to pass through but not others).

Osmosis is an important biological process because it allows water to pass in and out of cells.

Along with the three familiar states of matter mentioned above, there is another less familiar state, called plasma.

PLASMA: an ionized gas. Some electrons in plasma atoms are free, instead of being bound to the atom or molecule (see Atomic Structure section).

- Because the positive and negative charges move somewhat freely, plasma can conduct electricity and respond strongly to electric or magnetic fields.
- Plasma occurs under high-energy situations. For example, the extreme heat associated with lightning or the interior of stars produces plasma.
- Plasma usually takes the form of neutral gas-like clouds.

PHYSICAL CHANGES

Matter often changes, and these changes can be either physical or chemical.

Physical change: a change in a substance's form, without changing its chemical makeup.
The chemical formula of the substance remains the same before and after the change.

- Cutting a piece of paper into smaller pieces is an example of a physical change. The paper is in smaller pieces, but the chemical makeup of the paper has not changed.

PHASE CHANGES – TEMPERATURE

During a phase change, matter changes from one state to another as a result of changes in temperature or pressure.

Think About It...
Have you ever left a glass of ice water sitting out on a warm day, and noticed that the ice got smaller or completely disappeared? The ice didn't actually disappear, so what happened?

Melting: a phase change in which a solid becomes a liquid.

- **Melting point**: the temperature at which a substance begins changing its state from a solid to a liquid.

EXAMPLE:

At very low temperatures, water molecules have very low energy, so the water becomes a solid. The most common solid form of water is called ice.

As temperature increases, energy enters the solid ice. The energy releases the molecules from ice crystals, so they can move around freely, while still remaining in contact. When this occurs, ice melts to become liquid water.

- At one atmosphere (1 atm) of pressure, the melting point of ice is 0°C or 32°F.
  - The term "atmosphere" means the air pressure at sea level and is used as a unit of measurement for pressure.

Freezing: a phase change from a liquid to a solid.

- **Freezing point**: the temperature at which a liquid begins to form a solid.
- At 1 atm of pressure, the freezing point of water is 0°C. As you can see, the melting point and the freezing point of a substance are generally the same.

EXAMPLE:

Removing energy from liquid water causes the water molecules to return to ice crystals.

Quick Fact
The freezing point of a liquid can be lowered by adding a non-volatile solute. For example, the freezing point of water can be lowered by adding salt. As more salt is added, the freezing point becomes lower, making it more difficult for the liquid to freeze. Roads are salted or sanded in cold weather to prevent the formation of ice.

Evaporation: a phase change from a liquid to a gas.

- **Boiling point**: the temperature at which a liquid begins to form a gas.
- Once a liquid begins to boil, its temperature will remain constant until all of the liquid changes to a gas.

EXAMPLE:

When liquid water is heated, the molecules gather enough energy to escape to the gas phase. The gas phase of water is known as steam or water vapor.

- Boiling point varies with atmospheric pressure.
  - At high altitudes (such as on top of a mountain) where atmospheric pressure is lower, a material will boil at a lower temperature.
  - The "normal boiling point" (nbp) is the temperature at which the vapor pressure of a liquid equals 1 atm.

Quick Fact

The boiling point of a liquid can be increased by adding a non-volatile solute, such as pouring salt into water. As more salt is added, the boiling point becomes higher. Therefore, people living high on a mountain often like to put salt in their water so that it boils at a high enough temperature for proper cooking.

Condensation: a phase change from a gas to a liquid.

- **Dew point:** the temperature at which a gas begins to condense into a liquid.

EXAMPLE:

When energy is removed from water vapor, the vapor cools and condenses into liquid water.

If the temperature drops in the evening after a hot summer day, the water vapor in the air cools. As the cooler air nears the surface of the earth, it may come in contact with a surface. As the water vapor touches the surface, it cools even more and condenses, leaving droplets of liquid water on the surface, such as grass.

If the temperature is cold enough, you may see frost on the grass instead of dew.

Quick Fact

Condensation is the reason the windows inside your house get foggy on a cold winter day. Tiny water particles in the air come in close contact with the cold window and form water droplets on the surface of the window. Condensation can also occur on the outside of your windows on a hot summer day if the air inside your house is cooled by air-conditioning.

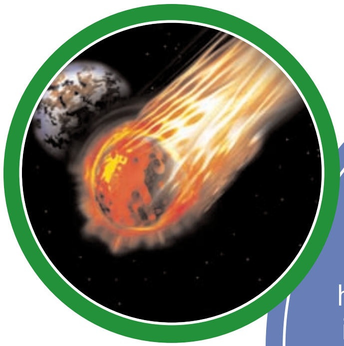
Sublimation: a phase change in which molecules go directly from a solid to a gas.

- Sublimation is difficult to observe with water because it happens so slowly. However, you may have noticed that ice cubes in the freezer get smaller over time. Molecules from the ice cube (solid phase) eventually escape to the gas phase.
- Sublimation can be easily observed with solid carbon dioxide (known as dry ice) or with iodine crystals.

Quick Fact

A comet is a small, icy body that orbits the Sun. Comets consist of a solid nucleus made of ice and dust, surrounded by a cloudy atmosphere called the coma and one or two tails.

Some comets become visible to humans on earth for several weeks as they pass close to the Sun. As a comet nears the inner solar system, heat from the sun causes some of the ice on the surface of the nucleus, to sublime, forming the coma. Radiation from the sun pushes dust particles away from the coma. These particles form a tail called the dust tail. We can see comets because the gas and dust in their comas and tails reflect sunlight.



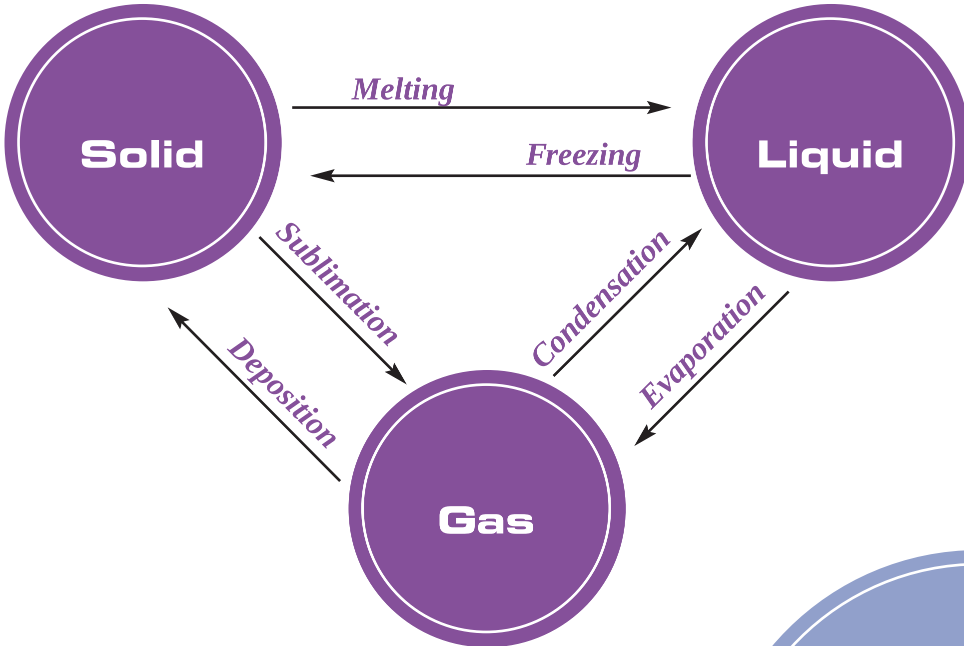
Deposition: a change from the gas phase directly to the solid phase.

EXAMPLE:

When ice crystals form on the inside of glass windows in cold weather, deposition has occurred.

NOTES

The following image illustrates the relationship between each of the phase changes:



PHASE CHANGES – PRESSURE

Phase changes occur because a substance has been given energy or energy has been taken away. One way to give a substance energy is through temperature. Another way is through pressure.

EXAMPLE:

Placing a heavy object on a block of ice causes the ice to melt.

Think About It...

Some people believe that the pressure exerted by the blade of an ice skate causes the ice to melt, allowing a person to glide across the ice. People have struggled to design an experiment to test this belief.

At 1 atm of pressure and normal room temperature, carbon exists as a solid known as graphite. Graphite is the gray substance found in the tip of a pencil. Under conditions of very high temperature and pressure, carbon will exist as a solid with a different structure – a diamond. Diamond and graphite are both in the solid phase but have distinctly sructured atoms.

- Therefore, graphite and diamond are two allotropes of carbon.

Quick Fact

The preferred state of carbon is the graphite form. Diamonds are actually unstable and over time will convert to graphite. This change takes about 1 million years, so while diamonds aren't necessarily forever, they do last a very long time.

CHEMICAL CHANGES

Chemical change: a change in which the atoms in the substance are rearranged, and bonds between the atoms are broken or formed. When a chemical change occurs, the resulting substance is different from the starting substance.

- Chemical changes involve more than just a change in form. A chemical change involves a reaction in which the structure or composition of the material changes.
- Chemical changes are almost always accompanied by energy changes.
- Chemical changes also involve the restructuring of how electrons hold the various atoms together.
- In some cases, chemical changes can be seen by a color change. In other cases, it can be very difficult to determine if a chemical change has taken place.

EXAMPLE:

A chemical change occurs when iron rusts. The iron reacts with oxygen (in the presence of water) in the air, forming iron oxide, the reddish-brown substance commonly called rust. The rust has different properties than the iron.

Other examples of chemical changes are the burning of fuel and the baking of a cake.

PHYSICAL AND CHEMICAL SEPARATIONS

It is important for scientists to be able to separate mixtures into their original parts. To do so, they will use a separation process.

Separation process: a process that divides a mixture of substances into two or more different parts.

- A separation process uses the different properties of the mixture's parts to accomplish the separation.

EXAMPLE:

A mixture of rocks and sand could be separated by using a screen, which would allow the fine sand particles to fall through but not the large rocks.

Pieces of iron could be separated from plastic in a recycling center using a magnet.

NOTES

PHYSICAL SEPARATIONS

Physical separations use the physical properties of the parts of a mixture to separate those parts without changing their properties. There are many physical separation processes (see Laboratory Separations section).

Filtration is a means of separating a mixture based on the different sizes of the parts. Filtration relies on the size difference between the particles making up one substance and the particles making up another substance.

- In a laboratory setting, filtration is often done by using filter paper to separate particles in a liquid. The filter paper comes in many grades. These grades represent the size of the tiny holes or pores in the paper. Different pore sizes allow the scientist to trap increasingly smaller particles within the mixture.
  - The filter paper is folded into a cone and placed in a funnel.
  - A solution is poured through the filter paper into a container (usually a flask).
  - The larger particles in the solution are trapped by the paper, while the liquid flows through the paper and collects in the flask below.
  - The collected liquid is called the **filtrate**. It is free of the suspended solid particles.

EXAMPLE:

One example of filtration (without the use of filter paper) occurs commonly in people's kitchens. After cooking pasta in a pot of water on a stove, the mixture is poured through a colander. The colander traps the pasta but lets the water pass through.

Quick Fact

Different size solids can be separated from one another by screening or sieving. A mixture is passed through a screen that traps the larger particles and allows the smaller ones to pass through. For example, a sieve is a useful tool for separating rocks from sand by trapping the rocks and letting the sand go through.

Filtration can be used to separate mixtures on a molecular level as well. For example "reverse osmosis" is a filtration technique that can be used to remove salt from sea water.

NOTES

CHEMICAL SEPARATIONS

Chemical separations require some type of chemical reaction to take place.

Precipitation: a means of separating a component in a solution by reacting it with another substance to form a solid.

- The solid that forms from the solution after the chemical reaction occurs is called a **precipitate.**
- If the solid particles are very small, they may remain in suspension. These particles can then be separated by means of filtration.

EXAMPLE:

One example of a chemical separation through precipitation involves water purification. Sometimes, lead (Pb^{2+}) has a way of getting into the water supply and needs to be removed because it is hazardous to your health. Chemists accomplish this by adding chloride (Cl^-) or sulfate (SO_4^{2-}). These substances will combine with the lead to make a new solid compound.

$$Pb^{2+} + SO_4^{2-} \rightarrow PbSO_{4(s)}$$

$$Pb^{2+} + 2Cl^- \rightarrow PbCl_{2(s)}$$

Once the solid compound is formed, the lead, which is part of a solid compound, can be removed from the water by filtration. As you can tell, chemical separations can be quite helpful!

Quick Fact

The water that moves through the pipes of your home often contains minerals and other substances. These substances may react to form solid precipitates that settle in the pipes, creating clogs.

ENERGY CHANGES

Physical and chemical changes involve energy (see Energy section).

Energy: the capacity to do work or produce heat.

- Energy is usually measured in "calories" or "Joules" (see Measurement section).
- Energy may take a variety of forms. The most common form is heat. Energy is also present in light, sound, electricity, and chemical bonds.
- Matter changes whenever energy is added or taken away. Although the energy is added or taken away, it is not lost. It simply changes form.

In every physical and chemical change, the total amount of energy stays the same. This principle is called the law of conservation of energy or the first law of thermodynamics.

Law of conservation of energy (**first law of thermodynamics**): energy easily changes from one form to another, but it cannot be created or destroyed.

EXAMPLE:

The energy stored in the tip of a match is in the form of stored chemical energy. When you light the match, the chemical energy becomes light energy and heat energy.

PHYSICAL ENERGY CHANGES

Phase changes generally involve an exchange of energy between the substance and its surroundings. The total energy of a substance is the sum of its kinetic and potential energy.

Kinetic energy (KE): energy associated with motion.

- If you know the mass and velocity (speed) of an object, you can determine its kinetic energy using the equation below:

$$\text{KE} = \frac{(mass) \times (velocity)^2}{2} \quad \text{or} \quad \text{KE} = \frac{1}{2} mv^2$$

- The faster an object is moving, the more kinetic energy it has.

Potential energy: stored energy.

- A ball at the top of a hill has potential energy. If it began to roll down the hill from the force of gravity, it would gain kinetic energy, while losing potential energy.

- If you know the mass, gravity, and height of an object, you can determine its potential energy using the equation below:

$$\text{PE} = \text{mgh}$$

CHEMICAL ENERGY CHANGES

Like physical changes, chemical reactions involve the gain or loss of energy. Many chemical reactions involve energy in the form of heat.

Heat: the amount of energy that moves from one object to another because of a difference in temperature.

- **Exothermic reaction**: a chemical reaction that gives off heat.
  - Any burning reaction is an exothermic reaction because it releases heat. These reactions include lighting a match and burning coal.
- **Endothermic reaction**: a chemical reaction that requires heat to be added.
  - When a substance becomes cold after a chemical reaction, the reaction was endothermic.

SECTION IV:
ATOMIC STRUCTURE

OBJECTIVES

- Define atoms and describe the parts of an atom.
- Identify and describe ions and isotopes.
- Locate elements on the periodic table.
- Utilize the periodic table to describe properties of elements.

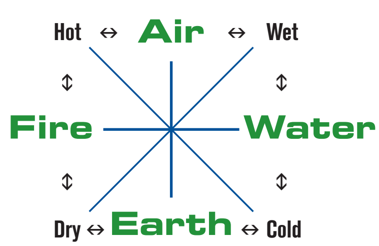
ATOMS

All matter is composed of basic elements that cannot be broken down further. These elements are the basic building blocks of matter that consist of one type of atom.

Atom: the fundamental unit of an element. The smallest particle of an element that retains the element's properties.

HISTORY: **ELEMENTS AND ATOMS**

In 440 BC, the Greek philosopher Empedocles taught that all matter was made of four simple substances: earth, air, water, and fire. About 100 years later, Aristotle related these substances to "blendings" of properties, such as coldness, hotness, dryness, and moistness. These beliefs are examples of theories that were later proven to be incorrect.

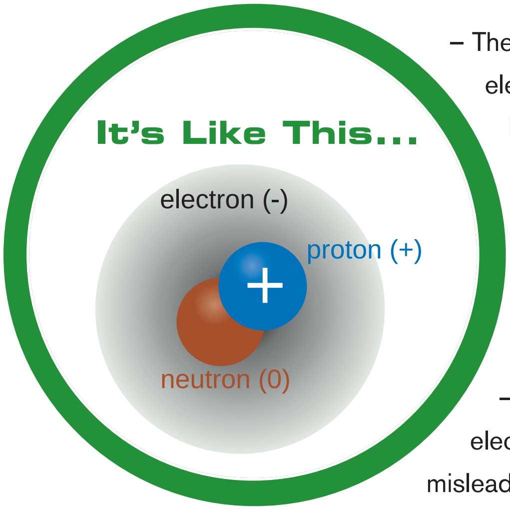


Before Aristotle popularized his ideas on matter, an ancient Greek philosopher Leucippus taught his students that while sand is made of fine particles, from a distance the beach looks continuous like the sea. Considering these teachings, one of Leucippus's students, Democritus, came up with the idea that all matter was made up of smaller indivisible building blocks. Democritus called these fundamental units of matter "atomos," meaning "cannot be divided." Unfortunately, since Aristotle rejected this belief, the atomic concept was not explored further until about 2,000 years later.

In 1661, modern chemistry began to emerge when Robert Boyle wrote what may be considered the first chemistry textbook, *The Sceptical Chymist*. His book rejected Aristotle's views and gave the first modern definition of elements.

Today, scientists know that atoms are made up of even smaller parts - protons, neutrons, and electrons.

- **Protons** and **neutrons** are located in the center of the atom. Together they make up the *nucleus* of the atom.
- **Electrons** occupy the space outside the nucleus.



– The proton and neutron both have a mass that is 1,836 times heavier than the electron. This difference in mass would be similar to comparing the mass of a house cat to the mass of a school bus.

– The proton has a positive (+) charge. The neutron has no electric charge. The electron has a negative (-) charge.

– Atoms are normally *electrically neutral*, which means they have the same number of protons and electrons.

– Electrons are located somewhere outside the nucleus. Many books show the electron as a small particle spinning around the nucleus, but this is a little misleading. Electrons actually move around the nucleus in cloud-like regions.

HISTORY: ERNEST RUTHERFORD

Ernest Rutherford (1871-1937) developed a nuclear theory of the atom. His theory is known as the Rutherford atomic model.

- **Rutherford atomic model:** an atomic model that describes the atom as having a tiny, dense, positively-charged core called a nucleus. Nearly all of the atom's mass is in the nucleus. The light negative parts, called electrons, travel around the nucleus. This movement of electrons is similar to planets revolving around the sun.

The Rutherford atomic model is also called the nuclear atom or the planetary model of the atom.

Rutherford also conducted studies on radioactivity, coming up with the terms alpha, beta, and gamma (see Radioactivity and Nuclear Reactions section).

With scientific advances, scientists have been able to study atoms using a particle accelerator or "atom smasher"! Two types of particle accelerators are cyclotrons and synchrotrons.

- In these accelerators, particles can move at such high speeds that the atom can break apart when it hits something.
- Particle accelerators have led to the discovery of more than 100 different subatomic particles, including quarks. **Quarks** are super-tiny particles that make up protons and neutrons. A quark has a fractional charge of either $^1/_3$ or $^2/_3$ the charge of an electron.

When two or more atoms interact with one another, they may form a larger unit called a molecule.

Molecule: the smallest particle of an element or compound that keeps the chemical properties of the element or compound.

- A molecule consists of atoms held together by chemical bonds.

Quick Fact

A water molecule is made up of two hydrogen atoms and one oxygen atom, giving the molecule the chemical formula H_2O.

A single molecule is extremely small and cannot be seen by the naked eye. While you can see a tiny drop of water, that drop contains lots and lots … and lots of water molecules!

ELEMENTS

A chemical element is distinguished by its atomic number.
Atomic number: the number of protons in an atom's nucleus.

Mass number (**atomic mass number):** the number of protons plus the number of neutrons in an atom's nucleus.

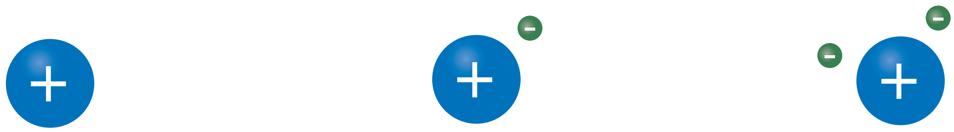
- The mass number for each isotope of an element is different (see section on Isotopes).

IONS

The number of protons in an atom never changes, but some atoms can either gain or lose electrons.

- Atoms with no positive or negative charge are electrically neutral.
- **Ion**: an atom or molecule that has lost or gained one or more of its outer electrons. As a result, ions have a positive or negative electrical charge.
  - – **Anions** are negatively-charged ions. When an atom gains electrons, the atom becomes negatively charged because the number of negatively-charged electrons exceeds the number of positively-charged protons.
  - – **Cations** are positively-charged ions. When an atom loses electrons, the atom becomes positively-charged.
- The atomic number and atomic mass of each ion stays the same because the number of protons and neutrons does not change.

The image below shows three forms of hydrogen: the electrically neutral atom and two ions.



A hydrogen cation (H^+) The hydrogen atom (H) A hydrogen anion (H^-)

*In the atom images above, the electron cloud, as shown in the image on page 36, has been removed for easier visualization.

ISOTOPES

Atoms of the same element can have different numbers of neutrons, but again, the number of protons never changes.

Isotopes: atoms of the same element with different numbers of neutrons.

H

HYDROGEN

Atomic #1

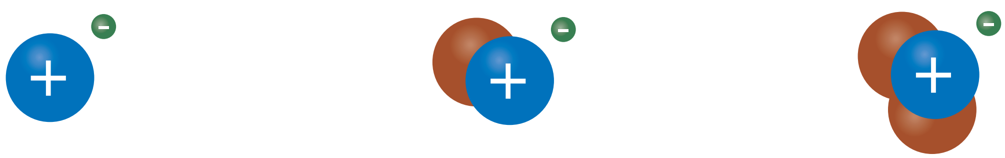
Hydrogen gets its name from the Greek words “hydro” (water) and “genus” (forming) because hydrogen is a component of water.

Characteristics:

- Is the lightest and most abundant element in the universe
  - Is estimated to make up more than 90% of all atoms (approximately 75% of the mass of the universe)
  - Is the primary component of Jupiter and the other giant gas planets
- Is highly flammable and used as a fuel by stars, causing them to shine brightly
- Is colorless and odorless in its gaseous form
- Has three common isotopes:
  1. *Protium*, the most common isotope, has one proton and no neutrons.
  2. *Deuterium* has one proton and one neutron.
  3. *Tritium* has one proton and two neutrons and is radioactive. (Tritium is produced in nuclear reactors and was used in the hydrogen bomb.)

EXAMPLE:

The most common isotope of hydrogen has no neutrons. A hydrogen isotope with one neutron is called deuterium. A hydrogen isotope with two neutrons is called tritium. The image below illustrates these isotopes of hydrogen:*



Hydrogen	Deuterium	Tritium
Atomic number = 1	Atomic number = 1	Atomic number = 1
Atomic mass number = 1	Atomic mass number = 2	Atomic mass number = 3

*In the atom images above, the electron cloud, as shown in the image on page 36, has been removed for easier visualization.

HISTORY: **HENRY CAVENDISH**

Henry Cavendish (1731-1810) was an English chemist and physicist, who first recognized hydrogen gas as a distinct substance, different from air.

He also demonstrated that burning hydrogen produced water.

- Using Cavendish's research, Antoine Lavoisier proposed that when hydrogen was burned, it was actually combining with something. He soon determined that hydrogen was combining with oxygen to make water.

Quick Fact

Atoms of the most abundant form of carbon have 6 protons and 6 neutrons in their nucleus. Therefore, this isotope has a mass number of 12 and is called carbon-12. About 1 of every 100 carbon atoms has 7 neutrons and a mass number of 13 (carbon-13). The even less common carbon-14 isotope is radioactive and has 8 neutrons.

Relative atomic mass (atomic weight): the "weighted" average mass of all an element's isotopes.

- The word "weighted" is important because the atomic weight of an element takes into account how often each isotope of an element is found in nature.

EXAMPLE:

Chlorine has two isotopes: chlorine-35 and chlorine-37. Chlorine is found in nature as chlorine-35 about 75 percent of the time. Chlorine is found in nature as chlorine-37 about 25 percent of the time. Therefore, the relative atomic mass of chlorine is:

$$\left(\frac{75\%}{100\%} \times 35\right) + \left(\frac{25\%}{100\%} \times 37\right) = 35.5 \textit{ amu (atomic mass units)}$$

HISTORY: **SCALE OF ATOMIC WEIGHTS**

John Dalton was the first scientist to propose that atoms had weight and thus, created a scale of atomic weights for elements. He chose the lightest element, hydrogen, as his reference. He assigned hydrogen a value of one and gave all other elements higher values depending on how much heavier their atoms were compared to hydrogen.

Soon after, Swedish chemist Jöns Jakob Berzelius proposed that the standard reference be oxygen instead of hydrogen. Hydrogen is so light that it was difficult to analyze. Berzelius believed comparing atoms of other elements to a heavier standard made more sense. His resulting table of atomic weights was quite similar to the one used today.

Until 1960, atomic weights were expressed on a scale in which oxygen was assumed to have 16 mass units. In 1961, a new unified scale was developed based on a value of 12 atomic mass units (amu) for the carbon-12 isotope.

PERIODIC TABLE

In 1869, a Russian chemistry professor named Dmitri Mendeleev developed a method for organizing the elements. He arranged the known elements in order of atomic weight. Then, he grouped them into rows and columns based on their chemical and physical properties. The result is what we call the **periodic table** (review the Periodic Table handout).*

HISTORY: MENDELEEV'S PERIODIC TABLE

The periodic table created by Mendeleev in 1869 is very similar to the one we see today, except his table had some missing elements. These missing elements were unknown at that time.

When Mendeleev created his table, he had no idea what atoms were made of or why they behaved in certain ways. His table is a remarkable piece of work because he created it before anyone knew about the structure of atoms. Even more amazing, Mendeleev was able to predict the existence and properties of several new elements based on the gaps in his table.

So, what's the big difference between Mendeleev's table and today's table? Today's table is organized by increasing atomic number, not increasing atomic weight.

- In 1914, Henry Moseley discovered that the atomic numbers of elements could be determined through experimentation. Before this discovery, atomic numbers were simply the order of elements based on increasing atomic weight.
- Once atomic numbers were determined to be significant on their own, the periodic table was reorganized.

ARRANGEMENT OF ELEMENTS ON THE PERIODIC TABLE

Groups (families): the vertical columns on the periodic table.

- Members of each group have the same number of electrons in their outer electron shell (see section on Electron Configuration).
- Because most reactions involve only the outer electrons, members of the same group participate in the same types of reactions.
- Members of the same group usually have very similar properties.

Periods: the horizontal rows on the periodic table.

- Members of the same period have the same number of electron shells but differ in how the shells are filled.
- To make the table easier to read, the elements in the sixth period between 57-71 (the Lanthanide Series) and in the seventh period between 89-103 (the Actinide Series) have been removed from the table and are shown in separate rows below the table.

*The Periodic Table handout is available on CEF's Web site. Ask your teacher or Local Challenge Organizer for details.

Elements on the periodic table are often identified as either metallic or non-metallic.

- Metallic elements are found mainly on the left side of the periodic table. These elements normally give up electrons.
- Non-metallic elements are on the right side. These elements normally take electrons in a chemical reaction.
- Semi-metals (metalloids) are located along the zigzag dividing line between the metals and non-metals. They have some characteristics of both metals and non-metals.

Quick Fact

Hydrogen is at the top of group 1, but it is not considered an alkali metal. Instead, hydrogen exists naturally as a diatomic gas. Therefore, hydrogen can have characteristics of two groups - one and seven.

ELEMENTAL GROUPS

Group 1A – Alkali metals (with the exception of hydrogen)

- Are soft low-density metals that have low melting points and tarnish easily
- Are extremely reactive and are rarely found in elemental form (as pure elements) in nature; are known to react strongly with water
- Have only one electron in their outer shell (see section on Electron Configuration)
- Have a strong tendency to give an electron away to achieve the very stable, filled outer shell state
- Form very strong bonds with the halogen elements that need only one electron to complete their outer shell
- Include lithium, sodium, and potassium

Na

SODIUM

Atomic #11

Sodium is the 6th most abundant element on the earth. It is found in many minerals, the most common of which is halite (rock salt or NaCl).

- Every gallon of sea water contains about 0.25 pounds of NaCl.
- Everyday table salt is mainly made up of the compound NaCl.

Characteristics:

- Is a soft, malleable, silvery, and highly reactive metal
- Is an essential element for living systems
  - Helps to regulate the balance of water in the body
  - Prolonged sweating results in sodium loss from the body

The most important sodium compound since ancient days has been table salt, which was typically used to preserve food. Another common sodium compound is sodium bicarbonate ($NaHCO_3$). It is commonly called baking soda. Sodium bicarbonate is used in baking, in antacids to neutralize excess stomach acid, and in fire extinguishers.

K

POTASSIUM

Atomic #19

Potassium gets its name from the English word “potash.”

Characteristics:

- Is a soft silvery metal in elemental form
- Has a lower density than water
- Is an essential element for human health, helping to keep a normal water balance between the cells and body fluids
- Reacts with water to produce KOH, hydrogen gas, and heat, which usually ignites the hydrogen

Potassium can be obtained by eating vegetables and fruits. Foods high in potassium include bananas, cantaloupe, and oranges.

Potassium is required for plant growth. It is found in most soils and is commonly used in fertilizers.

Just as the carbon-14 isotope is used to “carbon date” organic materials, potassium-40 is used to date rocks (see section on Radioactivity.)

Quick Fact

Potassium is an electrolyte, a substance that conducts electricity in the human body. Other electrolytes in the body include sodium, chloride, calcium, and magnesium.

Group 2A – Alkaline earth metals

- Are harder and denser than alkali metals
- Are gray- or silver-colored metals with high melting points
- Are very reactive metals (although less reactive than alkali metals)
- Are found in the earth's crust, but not in elemental form because of their reactivity
  - Are instead found in many rocks on the earth
- Have two electrons in their outer shell, which they tend to give away
- Include beryllium and magnesium

Quick Fact

Alkaline earth metals are called alkaline because they often form solutions with a pH greater than 7. Solutions with a pH level greater than 7 are defined as 'basic' or 'alkaline' solutions (see Acids, Bases, & pH section).

Mg

MAGNESIUM

Atomic #12

Magnesium gets its name from the Latin word "magnesia" after an ore found in the city of Magnesia, Thessaly, which is in Greece.

Characteristics:

- Is a grayish-white metal
- Is an essential element for human health
  - Helps to transmit nerve impulses and cause muscles to contract
  - Is found in our bones
- Is the lightest industrial metal, with a density that is about the same as man-made plastics

When magnesium metal is placed in a flame, it produces a bright white light as it combines with oxygen in the air to form magnesium oxide. Because of this property, magnesium is commonly used in flares.

Magnesium is a component of Epsom salts (hydrated $MgSO_4$). Epsom salts are used to soothe aches and pains.

Ca

CALCIUM

Atomic #20

Calcium gets its name from the Latin word "calx" meaning lime because it is found naturally in limestone as calcium carbonate.

Characteristics:

- Is the 5th most abundant element in the earth's crust but is not found in its elemental form
- Forms many natural compounds, including calcium carbonate, an important mineral for bone strength
- Is an essential component of leaves, bones and teeth, and shells

Calcium is a component of mortar, plaster, and cement. The Romans used it extensively for construction, and literature dating back to 975 AD mentions that plaster of paris ($CaSO_4$) is useful for making casts to set broken bones.

Groups 3-12B – Transition metals

- Have useful mechanical properties, such as good thermal and electrical conductivity
- Have high densities and very high melting points
- Cannot be divided neatly into individual groups
  - All of the transition elements have similar properties
  - Have low to moderate reactivity

Quick Fact

Transition metals often form colorful compounds. They also often form compounds in more than one oxidation state.

- ***Oxidation state:*** the charge that develops on an atom as a result of a loss or gain of electrons.

Most elements can only use electrons from their outermost orbital to bond with other elements (see sections on Electron Configuration and Bonding Basics). Transition metals can use the *two* outermost orbitals to bond with other elements. This chemical trait enables them to bond with many elements, and unlike other groups, transition metals do not always use the same number of outer electrons during chemical reactions. For example, iron sometimes gives away two electrons and at other times, gives away three.

Groups 10 and 11 contain the **precious metals** – silver, gold, palladium, and platinum.

Group 12 metals have low melting points. For example, mercury is a liquid under normal conditions.

Quick Fact

The transition metals iron, cobalt, and nickel are the only elements known to produce a magnetic field.

Fe

IRON

Atomic #26

Iron is found in large quantities in the sun and in many types of stars.

Characteristics:

- Is the 4th most common element, forming about 5.6% of the earth's crust; the core of the earth is believed to be mostly molten iron
- Is a strongly magnetic element
- Is an essential element for many living organisms, including humans.

Pure iron metal oxidizes in moist air to form rust. Iron ore appears as a reddish-brown solid.

Alloying iron with carbon creates steel (adding other metals, such as nickel and chromium, changes the properties of the steel, giving it greater strength, resistance to corrosion, less brittleness, and other favorable characteristics).

Zn

ZINC

Atomic #30

Zinc was known to humans in impure forms (in the form of alloys and compounds) since at least 500 BC. However, zinc metal was not known or used until much later.

Characteristics:

- Bluish-white lustrous metal
- Used in dry cell batteries
- Used to form alloys with copper, nickel, and aluminum

The largest industrial use of zinc is in the **galvanization** of iron to prevent corrosion. A layer of zinc is deposited on the iron. In the presence of air, the zinc oxidizes and protects the iron by forming a coat of $Zn_2(OH)_2CO_3$, to prevent further corrosion.

Zinc oxide (ZnO) is used as a pigment in paints and is a component in some cosmetics and ointments.

Like other elements in its family, zinc is found most often as a sulfide compound. Zinc sulfide (ZnS) is used in fluorescent lights, x-rays and TV screens.

Ag

SILVER

Atomic #47

The chemical symbol for silver comes from the Latin word for silver, "argentum." Silver has been used since ancient times. Today, it is often used to make coins.

Characteristics:

- Is a very ductile and malleable metal
- Is harder than gold
- Has the highest thermal and electrical conductivity of all metals

Silver is stable in pure air and water but undergoes a chemical reaction when exposed to air containing sulfur compounds. Silver combines with sulfur to form silver sulfide, a black substance known as *tarnish*.

Silver can be used to make high-quality mirrors. Silver has the ability to reflect almost 100% of the light that hits it. However, silver loses much of this ability when it becomes tarnished.

Sterling silver is an alloy of 92.5% silver and 7.5% copper. It is used in jewelry and silverware.

Au

GOLD

Atomic #79

Gold is used as a money standard in many countries. It is formed into bars and ingots for accounting and storage purposes.

Characteristics:

- Is the most ductile and malleable metal
- Is normally yellow in color but may look black, purple, or red when finely divided
- Is an excellent conductor of heat and electricity
- Reflects infra-red radiation well (see section on Types of Electromagnetic Radiation)
  - May be formed into a foil to help shield spacecrafts and skyscrapers from the sun's heat

Pure gold is very soft, so it generally needs to be alloyed with other metals to make it stronger. The term *carat* relates to the purity of gold. Gold that is 100% pure is called "24 carat" gold.

Hg

MERCURY

Atomic #80

Mercury is commonly known as "quicksilver." It gets its chemical symbol from the Latinized Greek name "hydragyrum," meaning liquid silver.

Characteristics:

- Is the only metal element that is liquid at room temperature (with room temperature between 65-75°F)
- Is found in the red mineral cinnabar (mercury (II) sulfide)
- Forms alloys called amalgams with gold, silver, and tin
- Does not conduct heat as well as other metals but does conduct electricity

Mercury was once commonly used in thermometers, barometers, and laboratory vacuum pumps. However, concerns about its toxicity have limited its use. Mercury exposure can cause damage to the central nervous system and the mouth, gums, and teeth.

Group 17A – Halogens

- Are non-metal elements
- Found most often as diatomic molecules - F_2, Cl_2
- Are highly reactive and only need a single electron to complete their outer shells
  - – Often bond with alkali metals
- Include fluorine, chlorine, and bromine

Quick Fact

The name halogen comes from the Greek words “hals” meaning “salt” and “gennan” meaning “to form or generate."

Halogens react with metals to form salts, such as sodium chloride (NaCl) and calcium chloride ($CaCl_2$).

F

FLUORINE

Atomic #9

Fluorine was first successfully isolated in 1886 by French chemist Ferdinand Frederic Henri Moissan, after nearly 75 years of unsuccessful attempts by numerous scientists.

Characteristics:

- Is the most electronegative element (see section on Electronegativity)
- Is the most reactive element, reacting with nearly all organic and inorganic substances
- Is a pale yellow-green color and is higly corrosive in gaseous form

In the late 1600s, minerals containing fluorine were used to etch glass. Eventually, scientists determined that the substance attacking the glass was hydrogen fluoride (HF). Early work with HF and fluorine was filled with accidents because of its extreme reacitvity.

When HF is dissolved in water, it is known as hydrofluoric acid, a very

Quick Fact

Fluorine is so highly reactive that it forms compounds with krypton, xenon, and radon – elements that were considered to be unreactive for many years.

Quick Fact

Fluorine’s reactivity also makes it difficult to store because it attacks glass and causes most metals to burst into flames.

Cl

CHLORINE

Atomic #17

Chlorine gets its name from the Greek word "chloros" meaning "pale green."

Characteristics:

- Is a very reactive halogen that combines directly with almost all elements
- Is not found in its elemental free state in nature but is found abundantly in the form of NaCl
- Is used commercially as a bleaching agent and disinfectant
- Is used to "chlorinate" water to make it safe to drink

As a gas, chlorine has a yellowish-green color. It is extremely irritating to the respiratory system and was used for chemical warfare during World War I.

NOTES

Group 18A – Noble gases (inert gases)

- Are all colorless, odorless gases under normal conditions and have low boiling points
- Have almost no reactivity
- Have complete outer electron shells, creating a very stable state (see section on Electron Configuration)
- Do not tend to form chemical bonds and are unlikely to gain or lose electrons
- Are commonly used in lighting, such as incandescent light bulbs and neon lights
- Include helium, neon, and argon

Quick Fact

Noble gases glow with bright colors in a vacuum discharge tube when an ordinary current and voltage are applied.

Restaurant signs that use these bright colored tubes are called "neon" signs because the most commonly used element is neon, which creates the red-orange color, you often see. Argon produces a pale purple color and other colors are created by mixing the gases together or mixing them with other elements.

He

HELIUM

Atomic #2

Helium gets its name from the Greek root "helios" meaning "sun." Helium was discovered in the sun's spectrum.

Characteristics:

- Is the lightest of the noble gases
- Is the second most abundant element in the universe
- Used for inflating lighter-than-air balloons because it is lighter than oxygen
  - Unlike hydrogen, helium is not combustible
  - French physicist, Jacques Charles is credited for being the first to use helium in a passenger balloon
- Does not combine easily with other elements to form compounds because it is a noble gas

Ar

ARGON

Atomic #18

Argon gets its name from the Greek word "argos" meaning "inactive."

Characteristics:

- Is colorless and odorless as a gas and a liquid
- Makes up slightly less than 1% of the earth's atmosphere by volume
- Is used in incandescent light bulbs because it prevents oxygen from corroding the hot wire filament inside the bulb

Argon-40 is the most abundant isotope of argon. Because argon-40 is produced by the decay of potassium-40, it is used to determine the age of rocks and minerals. This radioactive dating process is known as potassium-argon dating. By comparing the proportion of K-40 to Ar-40 in a sample, the date that the rock or mineral formed can be determined. Geologists have used this method to date rocks as old as 4 billion years. This dating method is also useful for dating many of the artifacts that are often found near ancient human burials.

LANTHANIDE AND ACTINIDE SERIES

The two rows at the bottom of the periodic table contain the **inner transition metals** (**rare earth metals)** and are not considered to be a part of any of the 18 groups.

Lanthanide series (or lanthanoid series): the inner transition metals from period 6 named after the first element in the series, Lanthanum.

- Are shiny, silvery-white metals
- Are chemically similar to each other with their properties differing only slightly with atomic number
  - Have high melting and boiling points
  - React violently with most nonmetals and burn easily in air

Actinide series (or actinoid series): the inner transition metals from period 7 named after the first element in the series, Actinium.

- Are strongly electropositive, shiny, hard metals that tarnish in air
- Are all radioactive elements and are generally used in the nuclear energy field

Studies on actinide properties have been difficult because of their radioactive instability.

Quick Fact

Most lanthanides emit colored light when they are bombarded by a beam of electrons, so they are used as phosphors in fluorescent light bulbs and television sets. They are also often used in lasers and sunglass lenses.

Quick Fact

Only thorium and uranium occur naturally in the earth's crust. However, neptunium and plutonium have appeared naturally in extremely small amounts in uranium ores as a result of decay.

ELECTRON CONFIGURATION

An element's chemical properties are determined by protons in the nucleus and the number of electrons moving in specific energy levels around the nucleus.

Electrons are found in atomic *orbitals*, a space that can hold up to two electrons. (Be careful not to confuse orbitals with orbits.)

- The **Heisenberg uncertainty principle** states that we can never determine exactly where the electron is located.
- **Electron density** is the probability of finding the electron in a particular part of the orbital.

Atomic orbitals are grouped into different *subshells* at different distances from the nucleus. Each shell is identified by its principal quantum number "n," which stands for the shell's energy level.

- For the lowest energy shell, n = 1. This shell is closest to the nucleus.
- As "n" increases, the shells are farther from the nucleus, higher in energy, and can hold more electrons.

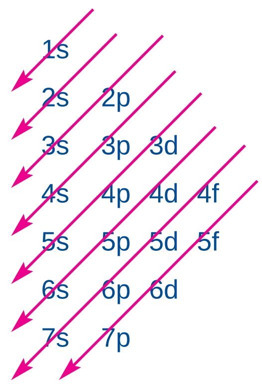
A subshell is composed of one or more orbitals that, in turn, hold a different number of electrons.

- *s subshells* have one orbital, which can hold up to two electrons.
- *p subshells* have three orbitals, each of which can hold two electrons, for a total of six electrons.
- *d subshells* have five orbitals, each of which can hold two electrons, for a possible total of ten electrons.
- *f subshells* have seven orbitals, each of which can hold two electrons, for a total of up to fourteen electrons.

The orbital and electron capacity for the four subshells is shown in the table below:

Shell	Number of Orbitals	Maximum Number of Electrons
s	1	2
p	3	6
d	5	10
f	7	14

Subshells are filled in a specific order, as shown in the diagram below:



As the diagram shows, the order for filling in the sublevels is: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, etc.

Start at the top (1s), with the beginning of each arrow, and follow it all the way to the point. As you go, fill in the subshells that the arrow passes through.

Scientists use a particular format to write an atom's electron configuration. The electron configuration below represents the element helium.

$$1s^2$$

- The first number refers to the principle quantum number "n," which stands for the energy level. In this case, the "1" represents the first electron shell, telling us that the electrons of helium occupy the first energy level of the atom.
- The letter refers to the type of orbital. In this case, the "s" tells us that the two electrons of the helium atom occupy an "s" subshell.
- The superscript number refers to the total number of electrons in that subshell. In this case, the "2" tells us that there are two electrons in the "s" subshell at the "1" energy level.

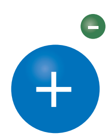
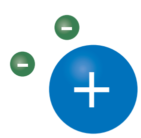
SECTION V:
CHEMICAL BONDS

OBJECTIVES

- Explain the relationship of Coulomb's law and electronegativity to chemical bonding.
- Identify the three primary types of bonds.
- Use Lewis Dot Structures to illustrate bonding.

COULOMB'S LAW

Opposites attract! The charges in an atom or molecule attract if they are different (one positive and one negative) but push each other apart if they are the same (both positive or both negative).

Correct		The electron (-) in its orbit is attracted to the proton (+) at the center of an atom.
Correct		These two electrons repel each other.
X **Incorrect**		The two electrons are attracted to a proton, but the electrons are not positioned in the best way. How should they be positioned?

*In the atom images above, the electron cloud has been removed for easier visualization.

ELECTRONEGATIVITY

Electronegativity is a measure of how strongly the nucleus of an atom attracts electrons in a chemical bond. In other words, electronegativity measures how greedy an atom is for electrons.

- Electronegativity increases from left to right across the periodic table.
  - The number of protons in an atom increases from left to right across the periodic table. Protons have a positive charge, so atoms with many protons have a high charge. Atoms with a high charge will have the strongest pull.

EXAMPLE:

The electronegativity (from least to greatest) for the second row of the periodic table is: Li, Be, B, C, N, O, F.

Notice that neon (Ne) is not listed. Neon is already full of energy level 2 electrons.

EXAMPLE:

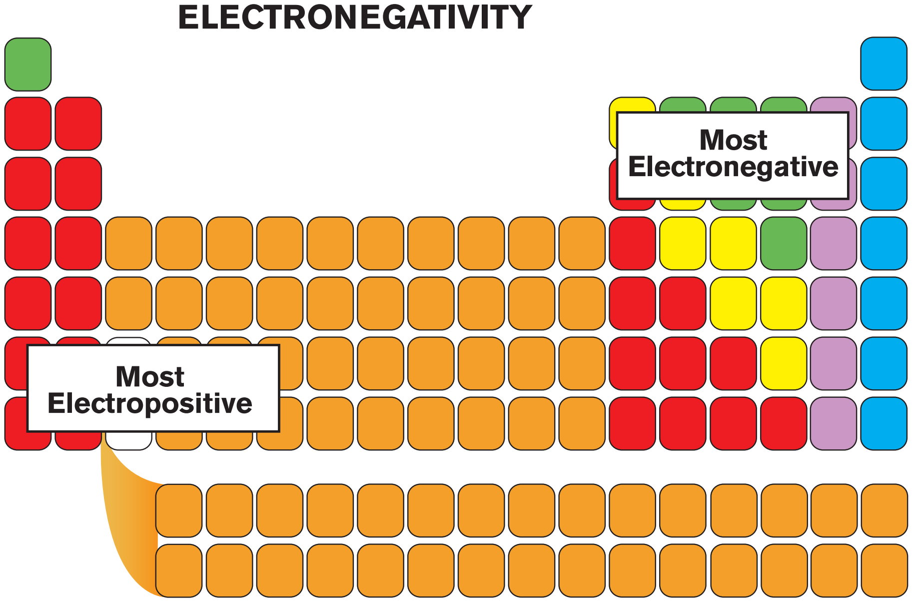
Compare sodium to its neighbor, magnesium. According to Coulomb's law, magnesium has a greater pull on the electrons because magnesium has more protons to do the pulling. Therefore, the electronegativity of magnesium is higher than that of sodium.

Quick Fact

Electronegative elements are found in the upper right of the periodic table (excluding the noble gases). Fluorine is the most electronegative element. Francium is the least electronegative, and therefore, is the most **electropositive** element, which means it gives up electrons most easily.

- Electronegativity decreases from top to bottom down the periodic table.
- When bonding, strong electronegative elements will attract electropositive elements.

ELECTRONEGATIVITY



- **Ionization energy:** the amount of energy needed to remove the outermost electron from an atom. The most electronegative elements have the highest ionization energies.

- Other properties of elements are related to electronegativity. Going from left to right across the periodic table:
  - Acid properties increase (more acidic).
  - Base properties decrease (less basic).
  - Atomic radii decrease.
  - Metallic properties decrease.

Quick Fact

The trend going from top to bottom down the periodic table is as follows:

Acid properties decrease (less acidic).

Base properties increase (more basic).

Atomic radii increase.

Metallic properties increase.

BONDING BASICS

Bonding involves only the outermost electrons of an atom, known as the *valence electrons*.

EXAMPLE:

Beryllium (Be) contains 4 electrons: 2 inner-shell electrons (level 1) and two second-shell electrons (level 2). Normally, only the level 2 electrons are involved in bonding. These level 2 electrons are Beryllium's valence electrons.

Let's explore how electronegativity controls three types of bonding – ionic, covalent, and metallic.

IONIC BONDING

Ionic bonds occur between atoms of elements located on opposite sides of the periodic table when one atom gives electrons, and another atom takes them.

EXAMPLE:

When sodium (Na) and chlorine (Cl) combine to make sodium chloride (NaCl), the chlorine atoms want the extra electrons from the sodium atoms. Chlorine is on the electronegative side (the greedy side) of the periodic table. Sodium is on the electropositive side and gives away electrons to the chlorine atoms.

Step 1: $Na \rightarrow Na^+ + \text{electron}$ **(production of Na cation plus release of electron)**

Step 2: $\text{electron} + Cl \rightarrow Cl^-$ **(released Na electron reacts with Cl to produce a Cl anion)**

Combined: $Na + \cancel{\text{electron}} + Cl \rightarrow Na^+ + Cl^- + \cancel{\text{electron}}$

Notice that the electron produced in Step 1 is used in Step 2, so it is cancelled out in the Combined step (see section on Ions for a review of cations and anions).

EXAMPLE:

What would happen if magnesium (Mg) atoms were bonding with Cl atoms instead?

Step 1: Mg → Mg^{+2} + 2 electrons (production of Mg cation plus release of electrons)

Step 2: 2 electrons + 2 Cl → 2 Cl^{-} (released Mg electrons react with Cl to produce Cl anions)

Combined: Mg + ~~2 electrons~~ + 2 Cl → Mg^{+2} + 2Cl^{-} + ~~2 electrons~~

Notice that twice as many Cl atoms are needed to take in the electrons released by the Mg. Therefore, the formula is $MgCl_2$.

Quick Facts

The periodic table can be used to predict ionic compounds. Remember: all atoms want electron configurations like the noble gases.

In the $MgCl_2$ example, Mg wants to be like Ne. Mg can only do this by losing two electrons. Chlorine wants to be like Ar, which only requires one electron. Thus, two chlorine atoms are required to complete the bond.

Here's the trick:

- Count two boxes backward from Mg to get to Ne. Give the 2 to the Cl.
- Count one step forward for Cl to get to Ar. Give that 1 to the Mg.
- The result is Mg_1Cl_2. Since we don't show the number one in formulas, we write $MgCl_2$.

COVALENT BONDING

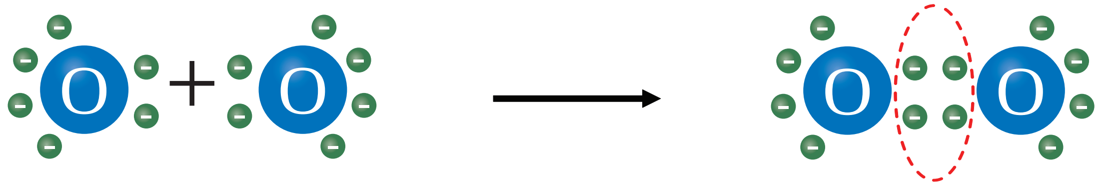
Covalent bonds occur when the electrons are *shared* between two adjacent atoms.

- Covalent bonds are especially stable.
- In a covalent bond, one atom does not actually lose an electron that is then gained by another atom. Instead, the atoms share the electrons.

EXAMPLE:

The oxygen you breathe is the gaseous molecule O_2. Because this molecule is made of two oxygen atoms, each of the atoms equally wants to look like the nearest noble element, neon. The two oxygen atoms agree to share each other's electrons.

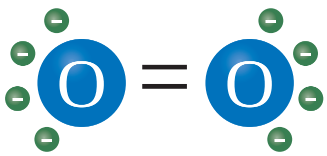
Remember that only the six outermost (level 2) electrons participate in the bond, so the reaction can be shown as:



The four electrons inside the dotted oval, above and to the right, are shared, so each oxygen atom has access to eight electrons. Therefore, both atoms look like the electron configuration of neon.

A covalent bond contains two electrons, which means there are actually two covalent bonds in an O_2 molecule (4 electrons shared, divided by 2 electrons in each bond = 2 bonds).

- Chemists often show two electrons as a line. Thus, O_2 is shown as:



- The bond that forms between the oxygen molecules to make O_2 is called a double covalent bond. Other chemicals may contain single covalent bonds or triple covalent bonds.
  - **Single bond:** a covalent bond sharing only one pair of electrons.
  - **Double bond:** a covalent bond sharing two pairs of electrons.
  - **Triple bond:** a covalent bond sharing three pairs of electrons.

Quick Fact

Structures (such as the one at left) that show atomic centers and either lines or dots for the outer shell electrons are called Lewis Dot Structures.

This name was given in honor of Gilbert N. Lewis for his contributions to bonding theory.

Quick Fact

Think of compounds held together by covalent bonds as being married. Ionic bonds might just be dating, choosing to date a different partner because of some environmental change, such as the addition of a solvent.

Atoms do not have to be identical to form a covalent bond. They must simply be near each other and have similar electronegativities. They generally occur between non-metal elements.

METALLIC BONDING

Metallic bonds form when the atoms in a metal contribute their electrons to a "sea" of shared electrons that spans the entire object.

- Metallic bonds are collective by nature. A single metallic bond does not exist.
- In a metal, the outermost electrons are shared among all the atoms in the solid.
  - The creation of an electron "sea" only occurs if there is nowhere else for the electrons to go.
  - Metallic bonds tend to occur when the Coulombic forces attracting the electrons are weak in comparison to the electron energy, allowing the electrons to be easily lost by the atoms. Each atom gives up its outermost electrons, forming a "sea" of electrons.
- Elements along the left side of the periodic table often form metallic bonds.
- Metallic bonds also form in elements that have electrons with high energy. These elements' atoms do not give up electrons to other substances so easily.
  - Row 4 elements, such as Sc, Ti, and V, have high energy.
  - A few of these metallic elements are actually quite easy to maintain in pure form, such as copper, silver, and gold.

Quick Fact

Silver and gold are precious metals because they do not react with most other elements, keeping their shiny characteristics.

Many of the properties of metals are a result of the high mobility of electrons in a metallic bond and the ability of those electrons to span the entire object.

- **Luster:** the ability of a metal to reflect light, resulting in a shiny appearance.
  - The large numbers of free moving electrons in a metal absorb and re-emit light.
- **Electrical conductivity:** a measure of the rate at which electricity can travel through a material.
  - Metals have good electrical conductivity because their electrons can move easily throughout the metal.
- **Thermal conductivity:** the measure of the rate at which thermal or heat energy can travel through a material.
  - Metals also have good thermal conductivity because as heat is applied to a part of the metal, the electrons become excited and travel to the other side of the metal. When the excited electrons travel, they carry the heat with them. The electrons are much better at carrying the heat energy than the nuclei of the atoms.

Think About It...

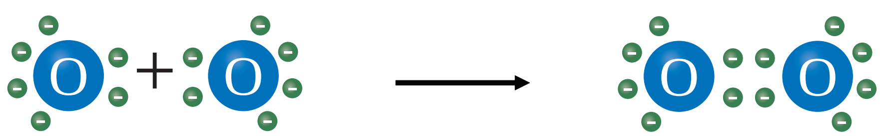
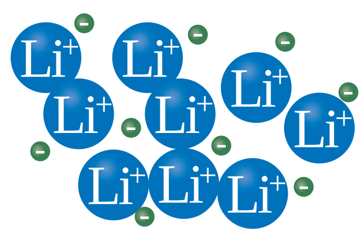
In the movie *A Christmas Story*, the character Flick is dared to put his tongue on a metal flagpole in freezing temperatures. He takes the dare, and his tongue gets stuck to the metal. Why would this NOT have happened if the flagpole had been made of wood or plastic?

- **Malleability:** the ability of metal to be flattened, shaped, or formed, without breaking, when pressure is applied, such as the ability of a metal to be hammered into a thin sheet.
  - – The mobility of electrons allow metal atoms to slide past one another when stress is applied without experiencing strong repulsive forces that would cause other materials to shatter.
- **Ductility:** the ability of metal to be stretched into a thin wire or thread without breaking.
  - – Like malleability, the mobility of electrons in a metallic bond allow the atoms to slide past one another as the metal is pulled and reshaped.

BONDING REVIEW

Below is a review of the three primary types of bonds – ionic, covalent, and metallic.

- Ionic bonding is essentially the result of a greedy atom combining with a giving atom.
- Covalent bonding is the result of atoms that need electrons but none of the atoms being greedy or giving.
- Metallic bonding is the result of massively shared electrons.

Ionic Bonding	
Covalent Bonding	
Metallic Bonding	

SECTION VI:
CHEMICAL REACTIONS

OBJECTIVES

- Identify the reactants and products of a chemical reaction.
- Describe and identify examples of types of chemical reactions.
- Explain and identify reversible chemical reactions.
- Identify exothermic and endothermic reactions.
- Understand rates of chemical reactions and the effects of catalysts.

A **chemical reaction** occurs when two or more substances interact, producing a change in substances.

- **Reactants**: the starting materials for a chemical reaction.
- **Products**: the substance or substances produced from a chemical reaction.
- **Byproduct**: a secondary product created at the same time as the desired product(s).

EXAMPLE:

A simple chemical reaction between hydrogen and oxygen is shown below:

$$2H_2 + O_2 \rightarrow 2H_2O$$

- **The hydrogen (H_2) and oxygen (O_2) are the reactants.**
- **Heat energy initiates the reaction.**
- **The resulting water (H_2O) is the product.**

TYPES OF CHEMICAL REACTIONS

In the chemical reactions described below, the letters A, B, C, and D are used to represent chemicals.

SYNTHESIS REACTION: a chemical reaction in which two reactants (A + B) combine to form a product (AB).

$$A + B \rightarrow AB$$

EXAMPLE:

$$2H_2 + O_2 \rightarrow 2H_2O$$

This simple reaction can power an automobile, such as the hydrogen-powered car.

Think About It...

For this reaction to power a hydrogen car, there is plenty of oxygen in the air, but how will we get the hydrogen?

DECOMPOSITION REACTION (DEGRADATION REACTION): a chemical reaction in which a compound (AB) breaks apart into two or more products (A + B).

$$AB \rightarrow A + B$$

EXAMPLE:

$$2HgO \underset{\Delta}{\rightarrow} 2Hg + O_2$$

In this decomposition reaction, mercury oxide (HgO) splits into mercury metal and oxygen gas.

The small triangle under the arrow means the reaction needs heat to take place. (Think of the little triangle as a little flame!)

Quick Fact

If you wanted to say what was occurring in the reaction at left using the English language, you would say: "Two mercury oxide bits decompose to give us two mercury atoms plus one molecular oxygen."

DISPLACEMENT REACTION: a chemical reaction in which a reactant (A) takes the place of some part of a compound (BC) to make a new compound AC and releasing a separate product (C).

$$A + BC \rightarrow AB + C$$

EXAMPLE:

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

This displacement reaction happens when you combine a piece of solid (s) magnesium metal with some liquid (aq) hydrochloric acid. When these reactants combine, they produce two products, a liquid solution called aqueous magnesium chloride and hydrogen gas (g).

Think About It...

How would we say the displacement reaction (at left) in English?

REARRANGEMENT REACTION (DOUBLE DISPLACEMENT REACTION): a chemical reaction in which parts of two compound reactants (AB + CD) are rearranged to form two or more different compound products (AD + BC).

$$AB + CD \rightarrow AD + BC$$

EXAMPLE:

$$Fe_2O_3 + 6HCl \rightarrow 2FeCl_3 + 3H_2O$$

In this double displacement reaction, iron oxide combines with hydrochloric acid to produce iron chloride and water.

Quick Fact

In the molecular world, the reaction may also look like:
AB + CD → DA + CB
since atoms do not care whether they are on the left or right side.

However, when scientists write these reactions, cations or the more electropositve element is generally written first (thus AD and BC).

CHAIN REACTIONS

A **chain reaction** is a series of chemical reactions in which the products of one reaction initiate further chemical reactions of the same kind. A product in the first step becomes a reactant in the second step, a product from the second step becomes the reactant for a third reaction, and so on.

- **Initiation reaction**: the chemical reaction that starts a chain reaction. The result sets up a sequence of repeating reactions.
- **Propagating reactions**: reactions that produce products that cause another reaction.
- **Termination**: the reaction or reactions that consume the substances needed to continue the reactions. At this point, the starting materials are exhausted.

Chain reactions are found in gas explosions, combustion, the formation of smog, and nuclear reactions.

Quick Fact

While chain reactions often occur rapidly, some may occur slowly, such as when edible oils oxidize.

With some chain reactions, the rate of the reaction continues to increase as the number of reacting particles increases, eventually resulting in an explosion.

EXAMPLE:

$$Cl_2 + \text{light} \rightarrow Cl\cdot + Cl\cdot$$

$$Cl\cdot + H_2 \rightarrow HCl + H\cdot$$

$$H\cdot + Cl_2 \rightarrow HCl + Cl\cdot$$

$$Cl\cdot + H_2 \rightarrow HCl + H\cdot$$

When chlorine and hydrogen interact (in the presence of light energy), a chain reaction occurs.

- **The light absorbed by a chlorine molecule breaks the molecule into special chlorine atoms, called chlorine free radicals (Cl•).**
- **The chlorine radicals then react rapidly with hydrogen molecules to produce hydrogen chloride and hydrogen free radicals (H•).**
- **The hydrogen radicals react with chlorine molecules, producing hydrogen chloride and chlorine radicals.**
- **Then, the chlorine radicals react further with hydrogen to continue the chain until some other reaction uses up the free radicals of chlorine or hydrogen.**

REVERSIBLE REACTIONS AND EQUILIBRIUM

Reversible reactions are reactions that can go forward (from reactants to products) or backward (from products to reactants), depending on the conditions of the experiment.

- Reversible reactions are usually represented in a chemical equation by a double arrow:

$$A + B \leftrightarrows C + D$$

EXAMPLE:

$$2NaCl + CaCO_3 \leftrightarrows Na_2CO_3 + CaCl_2$$

The forward reaction can be seen in saltwater lakes:

$$2NaCl + CaCO_3 \rightarrow Na_2CO_3 + CaCl_2$$

Sodium chloride (salt) in the water reacts with calcium carbonate (limestone rocks) to produce sodium carbonate and calcium chloride. Calcium chloride is the salty residue seen on rocks near salt-water lakes.

The reverse reaction can be produced in the laboratory:

$$Na_2CO_3 + CaCl_2 \rightarrow 2NaCl + CaCO_3$$

Sodium carbonate reacts with calcium chloride to produce sodium chloride and calcium carbonate.

- In a reversible reaction, both reactants and products may be present at the same time in a state of dynamic equilibrium.
  - **Equilibrium**: the state of a chemical reaction at which the forward and reverse reactions occur at equal rates, so the concentration of the reactants and products does not change.
  - Equilibrium describes how far a reaction goes, for instance, how much product a reaction can produce (unless we manipulate it!).

Quick Fact

Chemists and chemical engineers often try to manipulate equilibrium, causing a particular reaction to make more product than usual (essentially tricking the reaction).

ENERGY OF CHEMICAL REACTIONS

Chemists often initiate chemical reactions to produce energy. Other times, chemists add energy to a reaction to get a desired product.

- **Exothermic reactions:** chemical reactions that give off energy.
  - Exothermic reactions may occur spontaneously and often release energy in the form of heat, light, or sound.

$$A + B \rightarrow C + D + \text{energy}$$

EXAMPLE:

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + \text{energy}$$

The exothermic reaction above shows how methane and oxygen produce carbon dioxide, water, and heat energy.

Quick Fact

A common example of an exothermic reaction is burning wood in a wood stove. Wood combines with the oxygen in the air to produce carbon dioxide, water, light, and heat.

- **Endothermic reactions:** chemical reactions that require or absorb energy.

$$A + B + \text{energy} \rightarrow C + D$$

EXAMPLE:

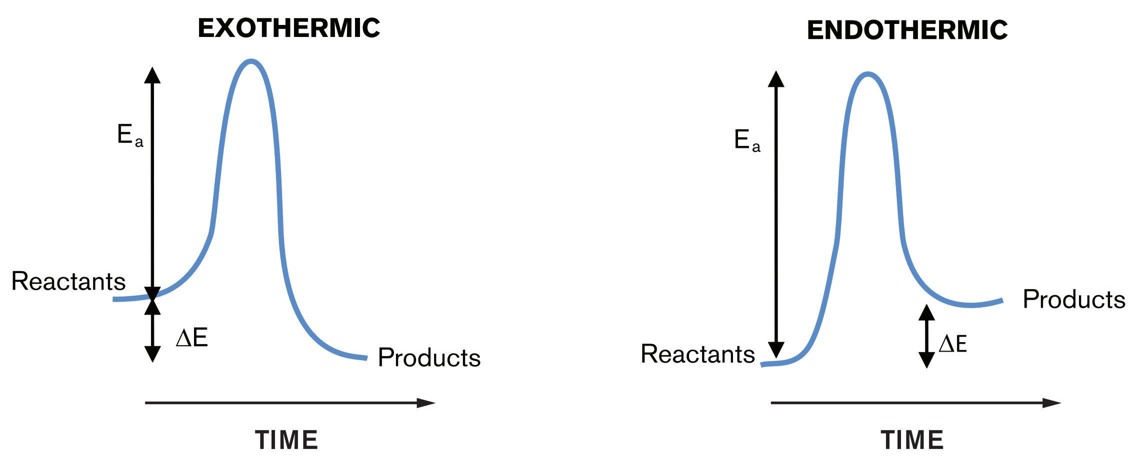
$$\text{energy} + 2Al_2O_3 \rightarrow 4Al + 3O_2$$

The endothermic reaction above shows that energy is added to bauxite (aluminum oxide) to produce aluminum metal and oxygen gas.

Quick Fact

A common example of an endothermic reaction is the process of photosynthesis. During photosynthesis, plants use the energy from the sun to convert carbon dioxide and water into glucose and oxygen.

- **Energy of activation:** the amount of energy needed to cause a chemical reaction.
  - Energy of activation (activation energy) is represented by E_a.
  - As the diagrams on the following page illustrate, the activation energy is the energy required to get the reaction "over the hill," to make it occur.
  - The symbol, ΔE, represents the change in energy from the starting energy of the reactants to the final energy of the products.



– In the exothermic reaction above, the reactants start at a higher energy level and end at a lower energy level. The difference, ΔE, is released from the reaction.

– In the endothermic reaction above, the products end up at a higher energy level than the reactants, which reflects the energy that had to be put into the reaction.

RATES OF CHEMICAL REACTIONS

The *rate*, or speed, of a chemical reaction is commonly affected by temperature and the concentration of the reactants and products.

- An increase in temperature usually increases the rate of the reaction.
- An increase in the concentration of the reactants increases the rate of the reaction.
- An increase in the concentration of the products decreases the rate of the reaction.

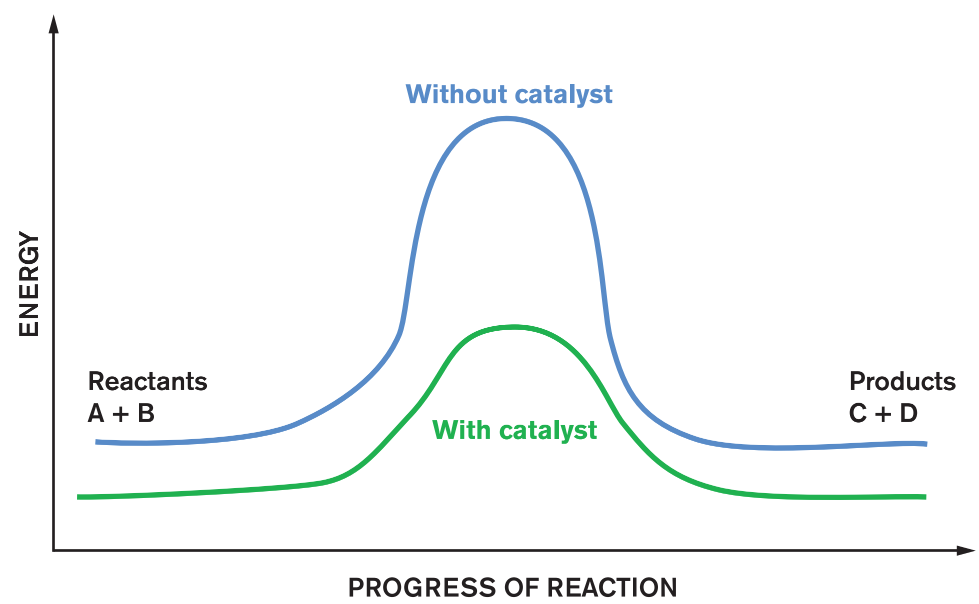
A **catalyst** is a substance that is not consumed in the reaction but helps to change the rate of the reaction.

- The catalyst usually alters the pathway the reaction takes.
- A faster pathway usually involves a lower energy barrier, which allows the reaction to occur at a faster rate.

Quick Fact

Humans need catalysts! Your body burns fuel (in the form of food), just like a car's engine burns fuel. However, your body doesn't require the amount of heat that a car needs to burn fuel because you have special catalysts in your body called enzymes. Enzymes allow us to burn fuel at our normal body temperature.

The diagram below illustrates the effect of a catalyst on the rate of a reaction. With a catalyst added, the “energy hill" a reaction has to climb is much lower.



NOTES

SECTION VII:
BALANCING CHEMICAL EQUATIONS

OBJECTIVES

- Explain the law of conservation of matter.
- Apply the law of conservation of matter to correctly balance equations.

When a chemical reaction occurs, it can be described by a chemical equation.

- A chemical equation shows the chemicals that react (the reactants) on the left side of the equation, and the chemicals the reaction produces (the products) on the right side.
- The chemicals are represented by their chemical symbols.
- Unlike mathematical equations, the two sides are separated by an arrow to show that the reactants form the products.

CONSERVATION OF MATTER

Law of conservation of matter (law of conservation of mass): matter cannot be created or destroyed, although it may be rearranged. According to this law, the mass of the reactants must equal the mass of the products.

Because all matter is made of atoms, the law implies the conservation of atoms as well, meaning that atoms are not lost (with the exception of nuclear reactions).

What does conservation of atoms mean? The number of atoms of each element on the reactants side (left side of the arrow) must equal the number of atoms of each element on the products side (right side of the arrow). When the atoms on both sides are equal, the equation is balanced, demonstrating conservation of atoms.

- If a hydrogen atom goes into a reaction, it has to appear somewhere in the products of the reaction.
- Likewise, if three hydrogen atoms appear on the reactant side of a chemical equation, three must appear on the product side.

EXAMPLE:

When methane interacts with oxygen in the air, the following reaction occurs:

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

As written above, the reaction correctly indicates that methane and oxygen combine to form carbon dioxide and water. However, this reaction violates conservation of matter because there are more oxygen atoms on the right and more hydrogen atoms on the left.

The correctly balanced reaction looks like this:

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

To determine the number of atoms in a chemical formula:

- Multiply the number in front of the chemical formula by the subscript number in the chemical formula.
- The number one is never written. CH_4 stands for $1C_1H_4$.
- To make sure the equation in the example above is balanced correctly:

  – Write the number of each type of atom on the reactant side.

  – Write the number of each type of atom on the product side.

  – Compare the numbers.

Quick Fact

Balancing chemical equations is like putting a puzzle together. You may not be able to tell which pieces fit where, so you may have to try a few different ways before you find a good fit. With chemical equations, you may not be able to see which numbers will work to balance the equation, so you have to try it out and see!

EXAMPLE:

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

Reactant Side of Equation	Product Side of Equation
C: $1 \times 1 = 1$	C: $1 \times 1 = 1$
H: $1 \times 4 = 4$	H: $2 \times 2 = 4$
O: $2 \times 2 = 4$	O: $(1 \times 2) + (2 \times 1) = 4$

HISTORY: **ANTOINE LAVOISIER**

Antoine Lavoisier (1743-1794) proposed the first version of the law of conservation of matter. This law states that during an ordinary chemical change, there is no noticeable increase or decrease in the quantity of matter.

Lavoisier is known as the Father of Modern Chemistry because he changed chemistry from a qualitative to a quantitative science.

He recognized and named oxygen. He also discovered the role oxygen plays in combustion.

GUIDELINES FOR BALANCING A CHEMICAL EQUATION

1. Write out the unbalanced equation and look to see which elements are not balanced (not equal).

$$C_2H_6 + O_2 \rightarrow CO_2 + H_2O$$

- There are 2 carbon atoms on the left side and only 1 carbon atom on the right side.
- There are 6 hydrogen atoms on the left side and only 2 hydrogen atoms on the right side.
- There are 2 oxygen atoms on the left side and 3 oxygen atoms on the right side.

2. Balance the equation. You will do this by multiplying the different atoms and molecules on each side by different amounts.

- Multiply CO_2 (on the right side) by 2. This is shown by placing a 2 in front of CO_2. The number in front of the molecule or atom (in this case, 2) is called a **coefficient**. As mentioned before, when no coefficient (or no subscript) is written, it is assumed to be one.

$$C_2H_6 + O_2 \rightarrow 2CO_2 + H_2O$$

- Be sure to multiply all atoms by the coefficient. Therefore, $2CO_2$ means there are 2 carbon atoms and 4 oxygen atoms. ***Do not change the subscripts***.
- Add a coefficient of 3 in front of H_2O on the right side, making it $3H_2O$.

$$C_2H_6 + O_2 \rightarrow 2CO_2 + 3H_2O$$

- Now, there are 2 carbon atoms on each side and 6 hydrogen atoms on each side, but there is an uneven number of oxygen atoms (two on the left side and seven on the right side).
- Add a coefficient of 3.5 in front of O_2 on the left side, making it $3.5O_2$.

$$C_2H_6 + 3.5O_2 \rightarrow 2CO_2 + 3H_2O$$

This equation is balanced, but we're not quite done.

3. A balanced equation should contain no decimals. In the equation above, the oxygen on the right is written as having a half molecule. Because there is no such thing as half an oxygen molecule, we eliminate that problem by multiplying everything by two.

$$2C_2H_6 + 7O_2 \rightarrow 4CO_2 + 6H_2O$$

This equation is the properly balanced equation for the reaction.

SECTION VIII:
CHEMICALS BY MASS

OBJECTIVES

- Define relative atomic mass.
- Identify the quantity of a mole and how it relates to the mass of atoms.

A single atom barely weighs anything. A hydrogen atom has a mass of only about 0.000 000 000 000 000 000 000 002 grams.

Even though atoms are very tiny, scientists still need to know the mass of an atom. In 1961, scientists adopted carbon-12 as the standard atom for comparing the mass of all other atoms.

- The mass of carbon-12 was fixed at 12 atomic mass units (amu) because it has 6 protons and 6 neutrons.
- All other masses were determined in relation to the standard carbon atom.
  - **Relative atomic mass (A_r):** the weighted average mass of all an element's isotopes compared with one-twelfth the mass of one atom of carbon-12 (see section on Isotopes).

$$A_r = \frac{\textit{average mass of isotopes of an element}}{\frac{1}{12} \times \textit{mass of 1 atom of carbon-12}} \quad \textit{or} \quad A_r = \frac{\textit{average mass of isotopes of an element}}{\textit{1 amu}}$$

  - The relative atomic mass or atomic weight of each element is listed on most periodic tables (see the Periodic Table of Atomic Weights handout).*
  - For most elements, the relative atomic mass is close to a whole number, so the whole number is often used for calculations.

Quick Fact

A mass spectrometer is used to determine the relative masses of atoms (as compared to carbon). The instrument separates ions to find out the percentage of each isotope that is present. These percentages are used to calculate the relative atomic mass of an element.

The mass spectrometer was invented by British scientist Francis Aston in 1919.

*The Periodic Table of Atomic Weights handout is available on the Study Materials section of CEF's Web site. Ask your teacher of Local Challenge Organizer for details.

THE MOLE

Along with the mass of tiny atoms, chemists often need to know the amount of a substance that has been used or formed during a chemical reaction. Atoms cannot be counted in the same way we count larger substances. So how did chemists solve this problem?

- The solution has been to weigh a large sample and then associate that weight with the number of atoms. The number chemists use is 6.0221415×10^{23}.
  - This number, called Avogadro's number, is usually estimated at 6.02×10^{23}.
  - **Mole**: an amount of substance containing 6.02×10^{23} particles.
  - Therefore, a mole of carbon is 6.02×10^{23} atoms of carbon.
- The relative atomic mass of an element in grams always contains 6.02×10^{23} atoms.

Quick Fact

The number written at left is a HUGE number. The number could be written out as 602,214,150,000,000,000,000,000, although, we do not actually know what those zeroes are after the 5.

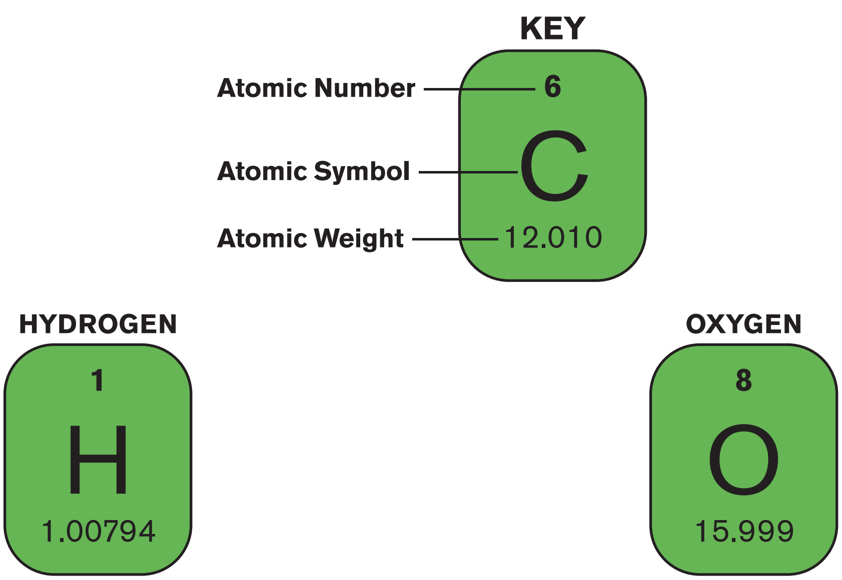
Quick Fact

Chemists sometimes have special parties called Mole Day parties. Their parties begin precisely at 6:02 a.m. on October 23 (10/23).

EXAMPLE:

- **A mole of hydrogen - 6.02×10^{23} atoms of hydrogen - has a mass of 1.00 grams. Likewise, 1.00 grams of hydrogen contains 6.02×10^{23} atoms.**
- **A mole of carbon – 6.02×10^{23} atoms of carbon – has a mass of 12.0 grams. Likewise, 12.0 grams of carbon contain 6.02×10^{23} atoms because carbon is 12 times heavier than hydrogen.**

Periodic tables often show an element's relative atomic mass (atomic weight) in the following format:



HISTORY: **AMEDEO AVOGADRO**

Amedeo Avogadro (1776-1856) was an Italian chemist known for his contributions to the theory of molarity and molecular weight. He also formulated Avogadro's Law.

- **Avogadro's Law** states that for equal volumes of gases at the same temperature and pressure, the gases will have equal numbers of molecules.
- Therefore, the relative molecular weights of any two gases are the same as the ratio of the densities of the two gases (under the same conditions of temperature and pressure).

Think About It...

If you have a mole of people, how many people are there?

Have there ever been that many people on the earth – at one point in time or over the entire history of human life on the earth?

NOTES

SECTION IX:
CHEMICALS BY VOLUME – SOLUTIONS

OBJECTIVES

- Describe the parts of a solution.
- Explain concentration and how to determine the concentration of a combined solution.
- Define saturation and supersaturation.
- Distinguish between polar and non-polar solvents.
- Describe important physical properties of solutions.

Many chemicals are combined in uniform mixtures called solutions. A **solution** is a homogeneous mixture of one or more substances (the solutes) dissolved in another substance (the solvent).

- Solutions consist of elements or compounds mixed together at a very fine level.
- **Solvent:** the substance of greater quantity in the solution. Solvents are sometimes referred to as the "inert ingredient."
- **Solute:** the substance of less quantity in the solution. Solutes are sometimes referred to as the "active ingredient."

EXAMPLE:

Dissolving salt into water creates a salt-water solution. In saltwater, the solvent is water. The solute is salt.

- **Concentration:** the amount of solute in the solution.
  - The concentration of a salt-water solution would tell you how much salt is in the solution.
  - To concentrate a solution, you would add more solute or remove some of the solvent in a given volume. By contrast, to *dilute* a solution, you must add more solvent or reduce the amount of solute in a given volume.
  - Two common ways to express concentration are percent composition by mass and molarity.
    1. **Percent Composition by Mass (%):** the mass of the solute divided by the mass of the solution (the mass of the solute plus the mass of the solvent), multiplied by 100.

EXAMPLE:

The percent composition by mass of a 100 g salt-water solution that contains 20 g of salt is:

$$\frac{20\ g\ salt}{100\ g\ solution} \times 100 = 20\%\ salt\text{-}water\ solution$$

Quick Fact

When you mix two solutions (of the same substances) with equal volumes, the concentration of the combined solution is the average of the two starting concentrations.

For example, if you mix equal amounts of a 20% salt-water solution with a 40% salt-water solution, the combined solution would have a 30% concentration.

2. **Molarity (M):** the number of moles of solute per liter of solution (which is not necessarily the same as the volume of solvent!).

EXAMPLE:
The molarity of a 4 liter solution containing 1 mole of salt is:

$$\frac{\textit{1 mol salt}}{\textit{4 L solution}} = \textit{0.25 M solution}$$

Quick Fact

In general, as the temperature of a solvent increases, the more solute it can dissolve. However, some compounds exhibit reverse solubility. Reverse solubility means that as a solvent gets warmer, less solute can be dissolved.

For example, oxygen gas dissolves better in cold water than in hot. As a result, certain fish must live in cold water in order to "breathe."

- **Saturation:** the point at which no more of a solute can be dissolved into a solvent.
  - This point changes significantly with different environmental factors, such as temperature, pressure, and contamination.
- **Supersaturated solution:** a solution that contains more dissolved solute than a saturated solution and is usually not stable.
  - A supersaturated solution is usually formed by cooling a saturated solution.
  - Crystals may form rapidly when the edge of the container is scratched or if a tiny seed crystal is added.

Solvents can be classified into polar and non-polar solvents.

- **Polar solvents:** solvents made up of molecules that have an uneven distribution of electrons, creating a negative and positive side. Generally, polar solutes will only dissolve in polar solvents.
  - Water, acetone, and acetic acid are polar solvents.
- **Non-polar solvents:** solvents made up of molecules that have an even distribution of electrons with the charges on the molecules neutralized. Non-polar solutes generally only dissolve in non-polar solvents.
  - Oil, benzene, and chloroform are non-polar solvents.

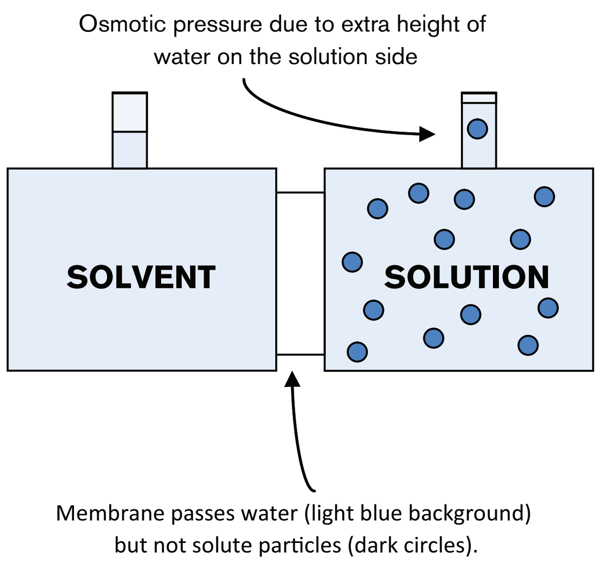
To figure out whether a solvent will dissolve a certain solute, remember: "Like dissolves like." This means that the more alike two molecules are, the more likely they are to mix.

- **Hydrophobic:** means "water fearing." Molecules that are different from water in terms of their polarity do not mix with it. They repel water and are said to "fear" it.
  - Oils, fats, wax, and alkanes are hydrophobic.
- **Hydrophilic:** means "water loving." Polar molecules, like water molecules, tend to mix well with water. They are said to "love" water.
  - Ammonia (NH_3) and glucose are hydrophilic.

PHYSICAL PROPERTIES OF SOLUTIONS

Certain physical properties of solutions make them very useful to chemists. Some of these important physical properties are described in the table below:

Physical Property of Solutions	Description of Property
Freezing point depression:	Solutions freeze at a lower temperature than the pure solvent. Antifreeze for automobiles (a solution made of water and ethylene glycol) freezes at a lower temperature than the pure solvent, water.
Boiling point elevation:	Solutions boil at a higher temperature than the pure solvent. Adding salt to water causes the solution to boil at a higher temperature, which decreases the amount of time it takes to cook food. This property of solutions is important at high elevations. Without adding salt, the water may boil at such a low temperature that the food might not be cooked properly.
Osmotic pressure:	Certain membranes permit the passage of a solvent, but not the solute. When such membranes are used to separate a solution and a solvent, more solvent flows into the solution. Eventually, the solution's weight builds up enough pressure to stop the flow.



SECTION X:
ACIDS, BASES, AND pH

OBJECTIVES

- Define and identify common acids and bases.
- Explain pH and describe substances as acidic or basic based on the pH scale.
- Identify and describe common indicators.

Quick Fact

Small changes on the pH scale actually mean large changes in acidity. A change in just one unit (pH 6.0 to pH 5.0) indicates a tenfold increase in acidity. If pH decreases by 3 points, the acidity has increased by 1,000.

The pH scale is used to measure the acidity of a solution.

- Acids release hydrogen ions [H^+], so the acid content of a solution is based on the concentration of hydrogen ions.
- Most substances range from 0 to 14 on the pH scale.
  - The smaller the number on the pH scale, the more acidic the substance is.
  - The higher the number on the pH scale, the more basic a substance is.
  - Pure water has a neutral pH of 7.0.
  - Some very strong acids may have a pH less than 0, and some very strong bases may have a pH greater than 14. For example, concentrated hydrochloric acid (HCl) has a pH less than zero.

Substance	Approximate pH	Approximate pH Indicator Paper Color
Battery acid	0.8	
Stomach acid	1.0-2.0	
Lemon juice, cola	2.3-2.5	
Vinegar	2.9	
Apple juice, orange juice	3.3-3.8	
Coffee	5.0-5.5	
Milk	6.5	
Pure water	7.0	
Human blood	7.4	
Sea water	8.0	
Baking soda solution	8.5-9.0	
Milk of magnesia	10.5	
Household ammonia	11.5-12.0	
Bleach	12.5	
Liquid drain cleaner	13.5-14.0	

ACIDS

Acids are solutions (or chemical compounds dissolved in water) that have an excess of hydrogen ions (H^+).

- Acids are able to give up H^+ ions to bases.
- Acids can conduct electricity and are corrosive in nature, with the ability to dissolve some metals.
  - When an acid reacts with a metal, it produces a metal salt and hydrogen.

EXAMPLE:
When magnesium comes in contact with hydrochloric acid, the acid dissolves the metal, producing magnesium chloride (a salt) and hydrogen gas.

$$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$$

Quick Fact

The word "acid" comes from the Latin term "acidus," which means sour. Acids generally have a sour taste.

Though remember, you should never taste a substance to determine what it is!

Quick Fact

Clean rain usually has a pH of 5.6, which is slightly acidic because of the carbon dioxide that is naturally present in the atmosphere. Rain measuring less than 5 on the pH scale is abnormally acidic and therefore, called *acid rain*.

BASES

Bases are solutions (or chemical compounds dissolved in water) that have an excess of hydroxide ions (OH^-) and accept H^+ ions from acids.

- Bases are able to donate OH^- ions to acids.
- Bases feel slippery to the touch and are often used to make soaps.
  - However, strong bases, such as drain cleaner, can be dangerous to your skin.
- Although the term "alkali" is often used as a synonym for base, they are not the same thing.
  - Alkalis are basic, ionic salts of an alkali metal or an alkaline earth metal. Therefore, *all alkalis are bases, but not all bases are alkalis*.

EXAMPLE:
Calcium carbonate and soda lye are bases that are also alkali salts.
Ammonia is a base, but not an alkali.

Quick Fact

Bases typically have a bitter taste and, like acids, can conduct electricity.

STRENGTH OF ACIDS AND BASES

Acids and bases may be strong or weak depending on how well an acid or base produces ions in water.

- A strong acid produces many hydrogen ions, whereas a weak acid produces fewer hydrogen ions. As a result, litmus paper registers slightly different colors depending on the strength of the acid (see section on Indicators).
- The strength of an acid may be affected by the size of the anion (see section on Ions) produced when the hydrogen is released into water.
  - Larger anions are more stable and are more easily separated from the hydrogen ion.
  - Larger anions are more acidic. (Anion size increases from top to bottom down the periodic table.)
- Strong bases act in a manner similar to strong acids, producing hydroxide ions instead of hydrogen ions.

An **amphoteric** substance behaves as either an acid or a base and is able to neutralize both bases and acids.

- Water and ammonia are common amphoteric substances. $Al(OH)_3$ and $Zn(OH)_2$ are also amphoteric substances.

INDICATORS

Indicators are substances that change color at a specific pH. They provide a way to determine the acidity of a solution. Some common indicators are:

- **Litmus paper:** an indicator that turns red in an acidic solution or blue in a basic solution.
- **Phenolphthalein solution:** an indicator that changes from clear to pink for a pH greater than 9.
- **Bromothymol Blue (BTB):** an indicator that turns yellow in acidic solutions and blue in basic solutions.

Quick Fact

Certain foods and flowers act as indicators. Cherries and beets appear red in acidic solutions but turn blue or purple in basic solutions. The flowers of hydrangea plants are blue in acidic soil but pink or white in basic soil.

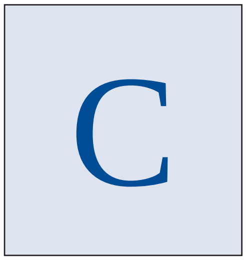
SECTION XI: ORGANIC CHEMISTRY

OBJECTIVES

- Define organic chemistry and identify organic compounds.
- Identify and describe hydrocarbons.
- Name organic compounds according to their structure.
- Explain hydrocarbon resources, such as coal and crude oil.
- Identify common chemicals in the human body and their functions.

Organic chemistry is the study of carbon-based or carbon-containing compounds.

- **Organic compounds:** a class of chemical compounds that contain the element carbon.
  - They often contain hydrogen but may also contain other elements (such as nitrogen or oxygen).
  - They are made by living things but can also be synthesized from inorganic materials in a lab.
  - They are usually held together by covalent bonds.
  - The few carbon-containing compounds not classified as organic include carbon dioxide, carbides, carbonates, and cyanides.



CARBON

Atomic #6

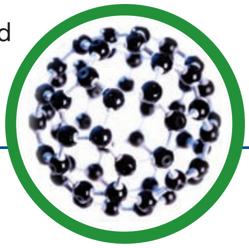
Carbon gets its name from the Latin word "carbo," meaning coal.

Characteristics:

- Is an essential building block for all organic compounds and forms more compounds than any other element
- Is a main constituent of DNA, proteins, and carbohydrates
- Is a main component of widely used fuels (coal, oil, and natural gas)

Elemental carbon is found in nature in four different solid forms:

- *Amorphous carbon*: found in soot, charcoal, and coal
- *Graphite*: (soft black layered form) used as a lubricant and in pencils
- *Diamond*: (crystalline form) one of the hardest known materials
- *Buckminsterfullerene*: (C_{60}, a sphere-shaped molecule) often called a "buckyball"



HISTORY: **ORGANIC CHEMISTRY**

"Organic chemistry" was named after the word "organism" becasue all organic compounds were originally obtained from living organisms or the remains of those organisms. Prior to 1828, scientists believed that organic compounds could only be created in living matter, while inorganic compounds were created from non-living matter.

However, in 1828, German chemist Friedrich Wöhler was able to synthesize urea (an organic substance found in the urine of most animals, including humans) from an inorganic compound. Wöhler's discovery changed the view of organic chemistry to focus on the internal structure of atoms in matter.

- **Inorganic compounds:** a class of chemical compounds produced either by natural processes or by humans in a laboratory.
  - They include minerals (such as salts and silicates), metals and their alloys (such as iron, copper, and brass), and pure water.

Quick Fact

The difference between inorganic and organic compounds has become less important in recent years. This change is mainly the result of the discovery of new compounds, the majority of which have been synthesized in labs rather than coming from "natural," organic sources.

NAMING ORGANIC COMPOUNDS

Hydrocarbons are the simplest organic compounds.

- Hydrocarbons are made only of carbon (C) and hydrogen (H) atoms.
- They are given different names depending on how many carbon and hydrogen atoms are in the chain.

Three types of hydrocarbons are alkanes, alkenes, and alkynes, which get their name based on the type of chemical bond they contain.

Alkanes: hydrocarbons that contain only single bonds.

- The name of each alkane begins with a different prefix, depending on how many carbon atoms are present.
- The names of alkanes all end with "-ane."

EXAMPLE:

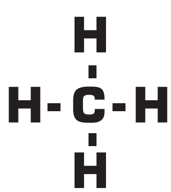
Methane is an alkane with only one carbon atom. Ethane is an alkane with two carbon atoms.

The table below lists the naming prefixes according to the number of carbon atoms they contain:

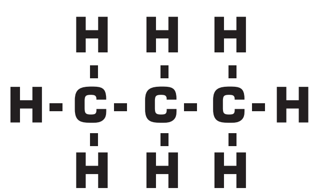
Number of Carbons	Prefix
1	meth-
2	eth-
3	prop-
4	but-
5	pent-
6	hex-
7	hept-
8	oct-
9	non-
10	dec-

EXAMPLE:

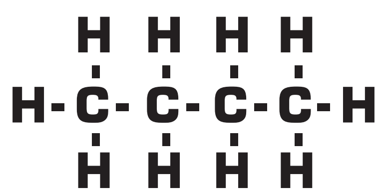
The figures below show the structure of certain alkanes and their names. Notice that the prefix for the alkane matches the number of carbon (C) atoms in the figure.



Methane
(Natural Gas)



Propane



Butane

Alkenes: hydrocarbons that contain one or more double bonds between the carbon atoms in each molecule.

- At least two carbon atoms must be present in an alkene
- To name an alkene, replace the "-ane" ending of an alkane with "-ene."

EXAMPLE:

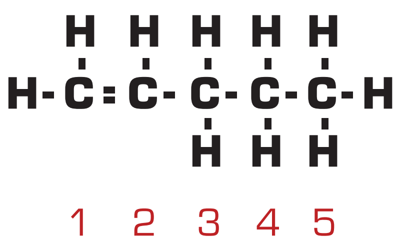
Ethane (CH_3 – CH_3) becomes ethene (CH_2 = CH_2). Notice that the line between the molecules changes from a single line (single bond) in ethane to a double line (double bond) in ethene.

- If there are more than three carbon atoms in the hydrocarbon, each carbon atom gets a number. This number is used to identify the position of the double bond.
  - The carbon atom on the end closest to the double bond is given the number 1. The carbon atom next to carbon 1 gets the number 2, and so on down the chain.

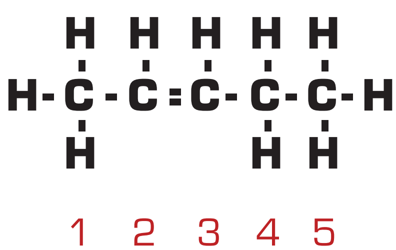
- To name an alkene, write the number of the carbon atom that is in the double bond and closest to the end of the chain first, to indicate the position of the double bond.
  - If the double bond is located at carbon 1, then this number does not have to be included in the name.
- The hydrocarbon chain is always numbered in the direction that gives the double bond the lowest possible number.

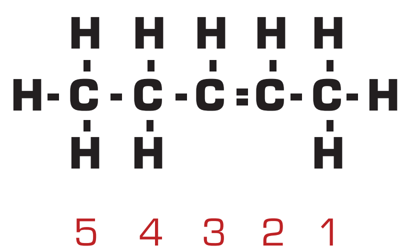
EXAMPLE:

The figures below show the structure of certain alkenes and their names. Notice that the number in front of the name matches the number closest to the end where the double bond is located. The prefix for the alkene matches the number of carbon (C) atoms in the figure.



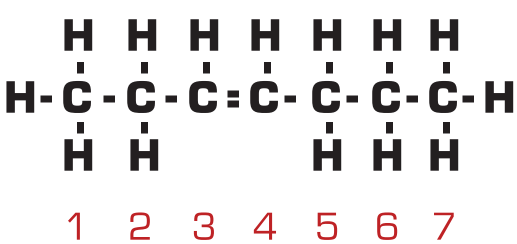
1-pentene (or simply "pentene")





2-pentene

Notice that the two figures above have the same name.



3-heptene

Alkynes: hydrocarbons that contain one or more triple bonds.

- To name alkynes, replace the "-ane" ending of an alkane with "-yne."

EXAMPLE:

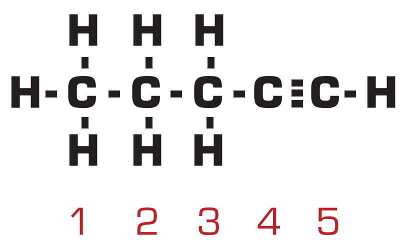
Ethane ($CH_3 - CH_3$) becomes ethyne ($HC \equiv CH$). Notice that the line between the molecules changes from a single line (single bond) in ethane to a triple line (triple bond) in ethyne.

Likewise, propyne is $HC \equiv CCH_3$.

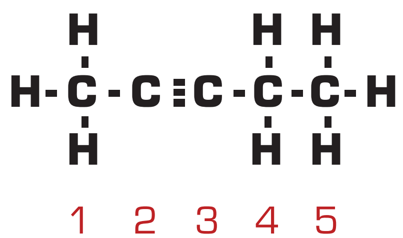
- Like alkenes, alkyne molecules that contain more than three carbon atoms have a number to indicate the position of the triple bond.
  - The carbon atoms are numbered in the direction that gives the triple bond the lowest possible number.

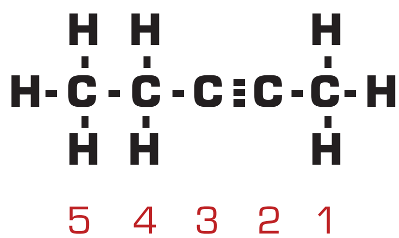
EXAMPLE:

The figures below show the structures of certain alkynes and their names. Notice that the number in front of the name matches the number closest to the end where the double bond is located. The prefix for the alkyne matches the number of carbon (C) atoms in the figure.



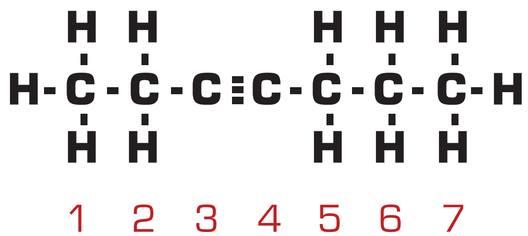
1-pentyne (or simply "pentyne")





2-pentyne

Notice that the two figures above have the same name.



3-heptyne

HYDROCARBON RESOURCES

Fossil fuels are fuels that come from the remains of living organisms that died millions of years ago. Fossil fuels are used to give power to our homes, schools, cars, and more. Some important fossil fuels are coal, oil, and natural gas.

Coal: a black to brownish-black sedimentary rock.

- Coal contains carbon, hydrogen, and oxygen combined with smaller amounts of nitrogen and sulfur.
- It was originally formed from ancient plants.
  - These plants died, decomposed, and were buried under layers of water and dirt millions of years ago.
  - The heat and pressure from the layers of dirt and rocks that built up over the trapped dead plant matter converted the plant remains into coal.
- Coal is considered to be a nonrenewable energy source because it takes millions of years to create.
- Most of the coal used in the U.S. and worldwide is burned by electric power plants to produce electricity (see Energy section).
- Coal is rich in hydrocarbons. The simplest hydrocarbons, alkanes, are easily combusted and provide energy.
  - **Combustion:** an exothermic reaction between a substance (the fuel) and a gas (the oxidizer) to release heat. Combustion is commonly called burning.
  - Combustion normally occurs in oxygen gas but can also take place in other gases.

EXAMPLE:

The general formula for the combustion of an alkane is seen in the combustion of methane below:

$$\text{Alkane (methane)} + \text{Oxygen} \rightarrow \text{Carbon dioxide} + \text{Water} + \text{energy}$$

$$CH_4\,(g) + 2O_2\,(g) \rightarrow CO_2\,(g) + 2H_2O\,(g)$$

- Just like methane, when coal is burned, it releases energy and water in the form of steam. The energy and steam give power to generators, creating electricity.

Crude oil: a mixture made of mostly hydrocarbon compounds.

- Crude oil can be found in many areas of the world. It has a major global influence both as a fuel source and in many industrial applications.
- Over 90% of crude oil is used to produce some type of fuel.
- To use crude oil, it has to be refined, and the components of the oil must be separated through a process called fractional distillation.
  - **Fractional distillation:** a method of separating a mixture into its components based on their relative boiling points.
  - Most of the compounds found in crude oil are alkanes, so they have similar chemical properties.
  - Fractional distillation can separate the hydrocarbons using heat because the longer the carbon chain of a hydrocarbon component, the higher its boiling point. The shorter the carbon chain, the lower its boiling point is.
  - The lower the boiling point, the higher up the fractionating column the component travels. The higher the boiling point, the less distance the component can travel up the fractionating column (see image on the following page).

NOTES

FRACTIONAL DISTILLATION

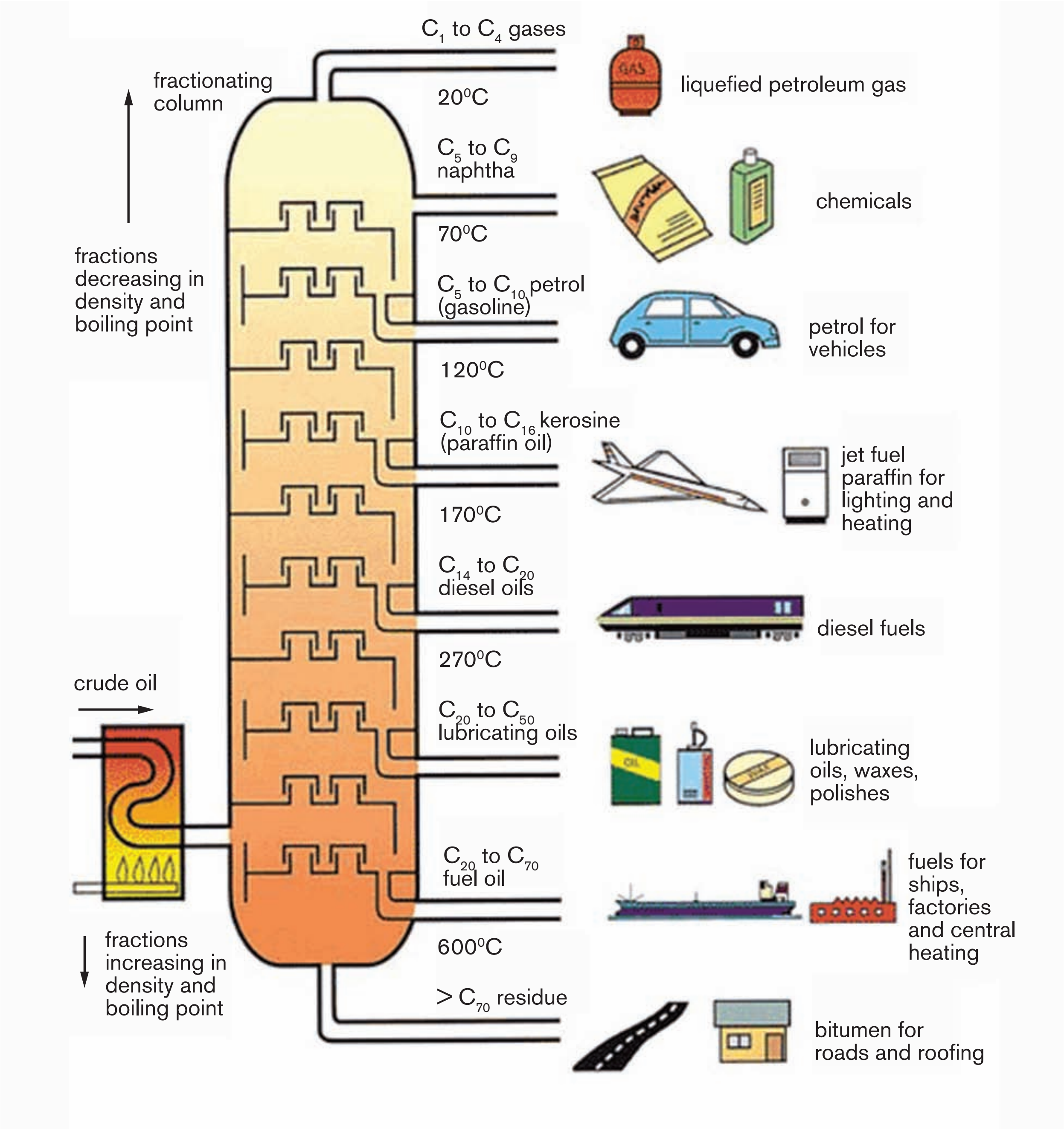


Image source: "Coryton Refining – Refinery Processes." *The Institute of Petroleum*. Energy Institute. September 16, 2008 www.energyinst.org.uk/education/coryton/page7.htm

CHEMISTRY IN THE HUMAN BODY

The human body is composed of chemical compounds, such as water and "organic" chemicals, containing the element carbon.

Organic compounds in the human body have varying roles that are essential to our survival. These organic compounds include proteins, carbohydrates, lipids, nucleic acids, and vitamins.

Quick Fact

About 60% of an adult human's body is made of water. Newborn babies are about 75-78% water. When a baby reaches one year old, the percentage of water in the body drops to about 65%.

PROTEINS

Proteins are complex organic compounds consisting of carbon, oxygen, hydrogen, nitrogen, and occasionally sulfur. Proteins are involved in almost all cell functions, helping the body to grow and to repair damage.

- Proteins are found in body organs, muscles, ligaments, skin, and hair.
- Some proteins provide structural support, others are involved in body movement, and others protect the body from germs.
- Proteins in the body include:
  - *Enzymes* to catalyze chemical reactions in the body.
  - Many *hormones*, such as insulin, to help regulate body functions.
- The body uses protein to make **hemoglobin**, a component of red blood cells that carries oxygen through the body.
- Proteins are made of long chains of amino acids, and therefore, are classified as polymers (see section on Polymers).
- **Amino acids:** the building blocks of proteins.
  - There are only about 20 amino acids, but they bond together in different combinations to create the millions of proteins found in living organisms.
  - The body can make most of the amino acids it needs to create proteins, but there are some it cannot make.
    - The nine amino acids that the human body cannot make are called **essential amino acids** and must be obtained from food sources.
    - While both plant and animal foods contain protein, animal foods, such as meat, fish, poultry, dairy, and eggs, are considered *high quality* or *complete proteins* because they contain the essential amino acids.

Quick Fact

The nine essential amino acids include phenylalanine, tryptophan, isoleucine, leucine, valine, methionine, lysine, and threonine. Histidine is also categorized as an essential amino acid. However, because studies have shown the dietary requirements for histidine to be very low, sources often exclude histidine and refer to only eight essential amino acids.

CARBOHYDRATES

Carbohydrates are the most abundant group of organic compounds found in living organisms. These chemicals are broken down during chemical processes known as metabolism to generate energy. They serve as the main source of energy in the human body.

All carbohydrates are made up of sugars (or saccharides).

- **Monosaccharides:** simple sugars that cannot be broken down further.
  - Monosaccharides result from the digestion of more complex carbohydrates.
  - They are small enough to be absorbed through the walls of the digestive system and transported into the blood.
  - Most monosaccharides are either:
    - *Hexoses* (sugars containing 6 carbons) like glucose, fructose, and galactose.
    - *Pentoses* (sugars containing 5 carbons) like ribose and deoxyribose.
  - Glucose, galactose, and fructose are known as *structural isomers* becasue they have the same chemical formula ($C_6H_{12}O_6$), but the atoms are arranged in a completely different order.
  - Ribose and deoxyribose sugars are important in the formation of the nucleic acids, DNA and RNA.

Quick Fact

The monosaccharide, glucose, passes in and out of cells where it is used for energy purposes. Abnormal levels of this chemical in the blood can lead to diabetes.

Quick Fact

The names for most sugar compounds end in the letters "-ose."

- **Disaccharides:** sugars composed of two monosaccharides.
  - Lactose (milk sugar) is a disaccharide made from one molecule of glucose and one molecule of galactose.
  - Sucrose (cane sugar), commonly found in kitchens, is made from one molecule of glucose and one molecule of fructose.
  - Maltose (malt sugar), is made of two molecules of glucose and is formed from the digestion of starch.
- **Polysaccharides:** relatively complex carbohydrates composed of many monosaccharide units joined together.
  - They are polymers that can have a linear or branched stucture.
  - Most natural carbohydrates occur in the polysaccharide form.
  - The polysaccharide *starch* is a naturally abundant nutrient carbohydrate found in corn, potatoes, wheat, and rice.

Quick Fact

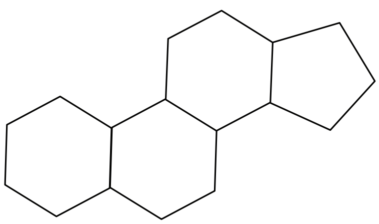
Chitin is a polysaccharide found in the exoskeletons of crustaceans, spiders, and insects.

LIPIDS

Lipids are a type of organic compound that are insoluble in water and play an important role in insulation (reducing the rate of heat transfer to keep the body warm) and regulation of body functions.

Lipids are often referred to as fats, but fats are only one type of lipid.

- **Trigylcerols (fats and oils):** organic compounds that are part of the lipid family, made up of a glycerol (a type of alcohol) and three fatty acids.
  - Fats have structural and metabolic functions in the human body. They are said to be the body's energy reserve because excess fat is stored in the body and is metabolized for energy when needed, such as for low-intensity exercise or in times when food is not available.
  - Fats are generally solid at room temperature, and oils are generally liquid at room temperature.
  - Triglycerols are also an important part of the human diet. Edible fats include butter and margarine. Edible oils include peanut oil and olive oil.
- **Steroids:** important biologically-active compounds in the lipid family that have a basic structure composed of three 6-membered rings and one 5-membered ring (as shown in the image below).

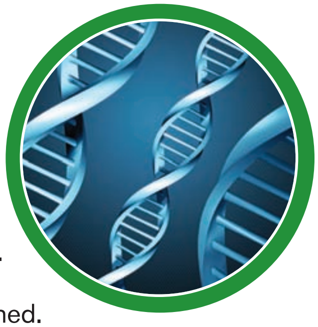


  - The most common steroid is cholesterol.
    - **Cholesterol:** a major component of cell membranes that is produced in a variety of body tissues, such as the liver, and obtained through food.
    - In the human body, cholesterol is concentrated in the brain and spinal cord.
    - Cholesterol is necessary for the production of vitamin D, bile acid (for stomach digestion), estrogen, and testosterone.

NUCLEIC ACIDS

Nucleic acids are large molecules that carry genetic information and direct cellular functions.

The two general types of nucleic acids are deoxyribonucleic acid (DNA) and ribonucleic acid (RNA).



- **DNA:** a nucleic acid that contains the genetic instructions for the biological development of all cellular life forms (and most viruses).
  - DNA has a double helix structure made up of two "polynucleotide" chains.
  - Each chain has a sugar-phosphate backbone with a series of bases attached.
  - The four different types of bases that are present in DNA are adenine, thymine, cytosine, and guanine.

- **RNA:** a nucleic acid that transmits genetic information from DNA to proteins made by the cell.
  - RNA is generally a single-stranded chain.
  - Like DNA, RNA contains the bases adenine, cytosine, and guanine. However, instead of thymine, RNA's fourth nucleotide is the base uracil.

VITAMINS

Vitamins are fat- or water-soluble organic substances, obtained naturally from plant and animal foods. They are essential for the normal growth and activity of the body.

Some common vitamins and their characteristics are listed in the table below:

Vitamin	Characteristics/Functions
Vitamin B	• The B vitamins are eight water-soluble vitamins critical for metabolic reactions in the body and for fighting infections. • The B vitamins include B1 (thiamine), B2 (riboflavin), and B3 (niacin). • The B vitamins often work together to deliver health benefits to the body, such as promoting cell growth and enhancing the function of the immune and nervous systems.
Vitamin C (ascorbic acid)	• The majority of plants and animals are able to synthesize their own vitamin C. Humans cannot synthesize vitamin C, so they must obtain the vitamin through their diet. • Vitamin C is commonly found in the juices of fresh fruits and vegetables. • Lack of vitamin C in the diet may cause scurvy, a disease characterized by extreme weakness and bleeding under the skin.
Vitamin D	• Vitamin D is a fat-soluble vitamin found mostly in milk, milk products, green leafy vegetables, seeds, and nuts. • Vitamin D is important for the development of bone tissue. • In humans, vitamin D_3 is produced by skin exposed to sunlight, specifically UVB radiation (see section on Types of Electromagnetic Radiation).

HISTORY: **DOROTHY HODGKIN**

Dorothy Crowfoot Hodgkin (1910–1994) used x-ray crystallography to determine the structures of penicillin, vitamin B12, and insulin. Her research has allowed scientists to develop treatments for diseases and deficiencies, such as anemia and diabetes. In 1964, Hodgkin won the Nobel Prize in chemistry, becoming the third woman to receive that honor.

SECTION XII:
ENERGY

OBJECTIVES

- Define the law of conservation of energy.
- Identify and describe types of energy.
- Explain current energy consumption.
- Identify different types of electromagnetic waves and explain their uses.

Energy is defined as the capacity to do work or produce heat. It may take a variety of forms, such as light, sound, electricity, chemical bonds, mechanical motion, and the most common form – heat.

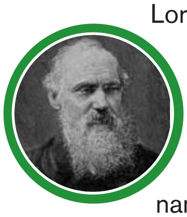
Energy can be converted from one form to another.

- **Law of conservation of energy (first law of thermodynamics):** while energy is easily changed from one form to another, it can be neither created nor destroyed.

EXAMPLE:

Heat from burning coal can be used to boil water, creating steam. The energy of steam can then be used to turn a rotary engine and produce mechanical energy. The engine may give power to a generator that produces electricity, and the electricity may be turned into light in a lamp, sound in a CD player, or heat in an electric toaster.

HISTORY: LORD WILLIAM THOMSON KELVIN



Lord William Thomson Kelvin (1824-1907) was a Scottish physicist and mathematician known for his work on energy. He changed the view of heat as being a fluid to an understanding based on the energy of the motion of molecules.

Thomson first defined the absolute temperature scale in 1847, which was subsequently named after him.

In 1851, he published ideas leading to the second law of thermodynamics.

HEAT

Heat is a form of energy that flows from one substance to another because of a difference in temperature. Heat will flow from a material at a higher temperature to one that is at a lower temperature.

Heat is commonly transferred (moved from one object to another) in one of three ways:

- **Conduction:** the transfer of heat through solid matter without any noticeable movement of the matter.
  - Energy is transferred from atom to atom in solid matter.
  - Conduction causes the metal handle of a pot on a stove to become warm as heat moves through the metal atoms that make up the pot and excite the atoms in the handle.
- **Convection:** the transfer of heat by movement within a fluid.
  - Convection occurs because of temperature differences within the fluid or between the fluid and its container.
  - As a result of convection, a cup of tea gets cooler as hot steam rises into the air.
- **Radiation:** the transfer of heat (as electromagnetic waves) through an empty space or clear material without heating the empty space or clear material.
  - The most common form of radiation is solar radiation. In solar radiation, the rays from the sun heat up the earth (see section on Energy).

HISTORY: **KARL SCHEELE**

Karl Scheele (1742-1786) was a Swedish chemist who published a book called *Chemical Treatise on Air and Fire*. In this book, he distinguished heat transfer by thermal radiation from heat transfer by convection or conduction.

He discovered many chemical substances, such as barium, chlorine, manganese, molybdenum, tungsten, and most notably, discovered oxygen independently of Joseph Priestley.

Scheele is also credited with recognizing the effect of light on silver compounds, thus providing the basis of photography.

Although heat is a form of energy, heat is measured differently than energy.

- **Calorie:** the amount of heat that is needed to increase the temperature of a gram of liquid water one degree Celsius.
- **British Thermal Unit (BTU):** the amount of heat that is needed to increase the temperature of a pound of liquid water one degree Fahrenheit. BTUs are the unit commonly used in the United States.
  - A BTU is equivalent to 252 heat calories.
  - BTUs are also used to describe cooling power. The air-cooling power of an air conditioning system describes the amount of thermal (heat) energy removed from an area.
  - Therefore, a 75,000 BTU heater and a 75,000 BTU air conditioner have the same capacity.

CURRENT ENERGY CONSUMPTION

Determining the total consumption of energy in BTUs would require us to calculate very large numbers. Therefore, large scale energy use is usually measured in Quads.

- **Quad:** one quadrillion BTUs (10^{15} times the amount of energy required to raise the temperature of one pound of water one degree Fahrenheit).

According to the 2008 energy consumption statistics provided by the U.S. Department of Energy:

- Total annual consumption of energy in the U.S. is about 100 Quads. World consumption is about 472 Quads.
- While the U.S. has only approximately 4.5% of the world's population, the country accounts for more than 21% of the world's energy consumption!

CURRENT U.S. ENERGY CONSUMPTION

Energy Source	Percentage of Total Energy Use
Petroleum (oil)	~ 37%
Natural gas	~ 24%
Coal	~ 22%
Nuclear energy	~ 8%
Renewable energy resources	~ 7%

Even if the U.S. was able to increase the use of renewable energy sources by 5 times their current use, we would still only be able to supply about one-third of our energy needs.

Fossil fuel resources, such as coal, oil, and gas, are hydrocarbons (see Organic Chemistry section.)

- When fossil fuels burn, they produce carbon dioxide. This production of CO_2 is a growing concern because of the build-up of CO_2 in the earth's atmosphere. The amount of CO_2 produced by a fuel varies directly with the percentage of carbon that it contains.
  - Natural gas contains about 75% carbon by weight.
  - Gasoline is about 85% carbon by weight.
  - Anthracite coal (a variety of mineral coal that has the highest carbon count and contains the fewest impurities) is about 90% carbon. Bituminous coal may contain between 75% and 90% carbon and is the most abundant variety of coal found in the U.S.
- Since the beginning of the Industrial Revolution (approximately 230 years ago), the concentration of CO_2 in the earth's atmosphere has increased by over 30 percent.
- There are many different estimates on how much the presence of CO_2 and other greenhouse gases in the earth's atmosphere could affect the earth's temperature.

With the large population growth in India, China, and other parts of the Far East, world energy use is expected to continue to grow at a high rate. How quickly our fossil fuel resources are depleted may depend on what limitations are put on their use in order to combat environmental concerns, such as global warming.

ELECTROMAGNETIC WAVES

Electromagnetic waves carry energy through space by radiation.

Electromagnetic waves are characterized by wavelength and frequency.

- **Wavelength:** the distance between one wave crest (top of the wave) and the next or between one wave trough (bottom of the wave) and the next.
- **Frequency:** the number of complete waves or pulses that passes a given point per second.

The product of the wavelength and frequency is a constant, c, the speed of light ($= 3 \times 10^8$ m/sec).

HISTORY: **MICHAEL FARADAY**

Michael Faraday (1791-1867) discovered electromagnetism by moving a magnet inside a wire coil to create electricity. Based on this discovery, he built the first electric motor.

Faraday developed the concept of a "field" to describe magnetic and electric forces. He also pioneered the field of electrochemistry. He introduced several words that we still use today to discuss electricity: ion, electrode, cathode, and anode.

To honor his accomplishments, a unit of electricity was named after him. The "farad" measures capacitance, an amount of electrical charge.

HISTORY: **MAX PLANCK**

Max Planck (1858-1947) was a German physicist often regarded as the father of quantum physics. In 1900, he hypothesized that energy (E) of a wave was directly proportional to its frequency (v).

$$E = hv$$

The constant, h, is known as *Planck's constant*.

Planck received the Nobel Prize in physics in 1918 for his quantum theory after it had been successfully applied to the photoelectric effect by Albert Einstein and to the atom by Niels Bohr.

TYPES OF ELECTROMAGNETIC RADIATION

Some types of electromagnetic radiation, in order of decreasing frequency (high to low) and longer wavelength (short to long) are:

TYPES OF ELECTROMAGNETIC RADIATION

Type	Wavelength
Gamma-rays	~ 10^{-10} m (meters) and below
X-rays	~ 10^{-8} m to 10^{-10}
Ultra-violet light	2 x 10^{-7} to 4 x 10^{-7} m
Visible light	4 x 10^{-7} to 7 x 10^{-7} m
Infra-red light	~ 10^{-4} m to 10^{-6} m
Radar, microwaves	~ 10 m to 10^{-3} m
Radio waves	~ 10 m and above

GAMMA-RAYS: electromagnetic waves with the smallest wavelengths and very high energy (the most energy of any other wave in the electromagnetic spectrum!). They are released by radioactive atomic nuclei and are also produced by the radioactive decay of other subatomic particles.

- They are the most penetrating electromagnetic waves, easily passing through the human body.
- Gamma-rays can kill living cells and are, therefore, used by doctors to kill cancerous cells in cancer patients.

X-RAYS: electromagnetic waves that have very short wavelengths and high energy. They are generally emitted as a result of the movement of an atom's electrons.

- When you have an X-ray photograph taken at a hospital, X-ray sensitive film is put on one side of your body, and X-rays are passed through you. Because your bones and teeth are dense and absorb more X-rays than your skin does, silhouettes of your bones or teeth are left on the X-ray film. Your skin appears transparent.
- The thickness of the earth's atmosphere prevents X-rays from passing through to the earth's surface.

Quick Fact
Metals aborb even more X-rays than bone, which is why metal fillings can be clearly seen on dental X-rays.

HISTORY: WILHELM CONRAD ROENTGEN

X-rays were first observed and documented by Wilhelm Conrad Roentgen (1845-1923), a German scientist who discovered them by accident when experimenting with vacuum tubes. His discovery sparked excitement across the world and medical uses were immediately explored.

Roentgen refused to take out any patents related to his discovery on moral grounds. He did not even want the rays to be named after him. In 1901, he received the first Nobel Prize for physics, and in November 2004, the 111th chemical element, Roentgenium, was named after him.

ULTRA-VIOLET (UV) RADIATION: electromagnetic waves that fall on the short wavelength side between 200 and 400 nanometers.

- The sun is our primary source of natural UV radiation.
- Exposure to UV light causes skin pigments to darken (sun tan) and, in excess, can lead to skin cancer.
- Physicists classify UV light into three types, based on wavelength:
  - UVA (400 to 320 nm)
  - UVB (320 to 290 nm)
  - UVC (290 to 200 nm)
- The shorter the wavelength and the lower the number, the greater the energy level of the light and the more damage it can do.

Quick Fact
The molecules in sunscreens absorb most UVB light and prevent it from reaching the skin. Sunscreens are typically classified by their "SPF" (Sun Protection Factor). The higher the SPF, the better your skin is protected. To protect your skin from damage, doctors generally recommend that you use a sunscreen with an SPF of 15 or higher. The sunscreen should be reapplied at least every two hours.

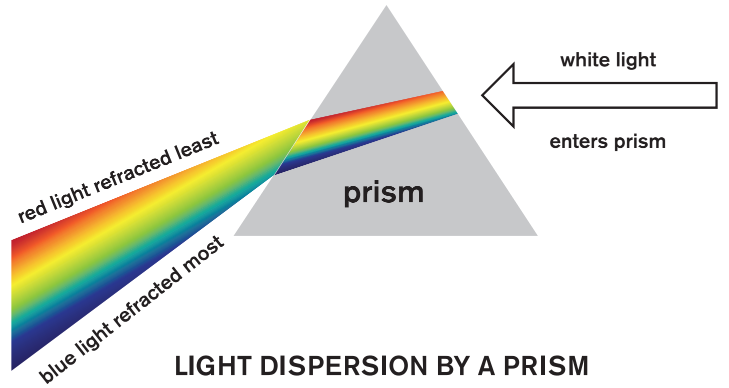
EXAMPLE:

Direct exposure to UVC for a length of time would destroy the skin. Fortunately,
UVC is completely absorbed by gases in the atmosphere before it reaches the ground. The longer wavelengths of UVB and UVA pass right through the atmosphere, even on a cloudy day. (That's why you can still get sunburned on a cloudy or hazy day.)

VISIBLE LIGHT (WHITE LIGHT):

a combination of waves that can be detected by the human eye. Wavelengths of visible light extend from about 400 nm for violet light to 700 nm for red light.

- White light may be split into different colors by passing it through a prism.
- The order of the resulting colors is: red, orange, yellow, green, blue, indigo, and violet.



LIGHT DISPERSION BY A PRISM

INFRA-RED (IR) RADIATION: electromagnetic waves that cannot be seen by the human eye but are often sensed as heat.

- IR radiation falls on the long wavelength side of the visible spectrum.
- "Near infra-red" waves have wavelengths closer to visible light and are used in TV remotes. "Far infra-red" waves are thermal and are closer to the wavelength of microwaves.
- IR radiation is used:
  - In medical imaging to measure skin temperatures.
  - In meteorological imaging to measure surface temperatures of different land and water features.
  - As an industrial heat source for some cooking and drying processes.

Quick Fact

Astronomers have found that IR radiation is quite useful for exploring areas of the universe surrounded by clouds of gas and dust. The longer wavelength of IR radiation can pass through these clouds, revealing what was invisible before.

MICROWAVES: electromagnetic waves with wavelengths between infra-red radiation and radio waves.

- Microwaves are used in satellites because they can pass through fog, light rain, and clouds. Visible light, on the other hand, is blocked by rain and fog, which you may have noticed during storms.
- Microwaves can be measured conveniently using inches or centimeters.
  - Shorter microwaves (only a few inches long) are used for remote sensing, such as in radar (like the Doppler radar used in weather forecasts).
- Microwave ovens, found in many household kitchens, get their name because they use long microwaves (about a foot in length) to heat food. The microwaves heat the food by causing the vibration of water molecules.

RADIO WAVES: electromagnetic waves with the longest wavelengths in the electromagnetic spectrum. Some radio waves are only a few millimeters and others extend over 100 meters (longer than a football field!).

- Radio waves are used for many different types of communication, such as transmitting music to radios or carrying signals to TVs and cell phones.
- Radio wave frequencies (very low energy) excite the nuclear spin transitions observed in:
  - *Nuclear Magnetic Resonance (NMR) spectroscopy:* a technique for determining the structure of organic compounds.
  - *Magnetic Resonance Imaging (MRI):* a medical imaging technique for observing the structure and function of the body.

NOTES

SECTION XIII:
RADIOACTIVITY AND NUCLEAR REACTIONS

OBJECTIVES

- Define radioactivity and radioisotopes.
- Explain half-life and use it in calculations.
- Identify common radioactive elements and describe their properties.
- Describe the difference between nuclear fusion and nuclear fission.
- Identify man-made elements and their location on the periodic table.

Elements tend to exist in more than one form, called isotopes, which differ in the number of neutrons in their nucleus. As a result, they also differ in their mass number (see section on Isotopes).

Isotopes have the same chemical properties as each other and undergo the same reactions. Because isotopes differ in atomic mass, they have different physical properties and are likely to undergo chemical reactions at different rates.

EXAMPLE:

Hydrogen is composed of three isotopes (protium, deuterium, and tritium). The lightest isotope, protium, generally undergoes chemical reactions most rapidly.

RADIOACTIVITY

Radioactivity is the spontaneous breakdown of an unstable nucleus in an atom that involves the release of energy in the form of electromagnetic radiation or particles.

- **Radioisotopes**: atoms that are radioactive.

The **half-life** of an isotope is the time it takes for one-half of the nuclei present in a sample to undergo radioactive decay.

- After one half-life, 50% of the original sample will remain.
- After two half-lives, 25% of the original sample will remain, and so on.

Quick Fact

Carbon-12 is the most common form of carbon. It was adopted in 1961 as the standard for defining all atomic weights.

Carbon-13 is non-radioactive and is frequently used for isotopic labeling studies. These studies follow how a carbon atom goes through specific reactions.

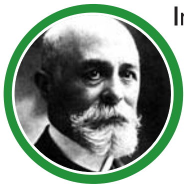
Carbon-14 is used in a process called carbon dating.

– After 5,730 years, half of the nuclei in a sample of a carbon-14 decay. (This period of time is its half-life.)

– Scientists use the predictable decay of carbon-14 to determine the age of organic materials up to 50,000 years old.

– Carbon dating is useful for studying artifacts left behind by ancient cultures.

HISTORY: **HENRI BECQUEREL**



In 1896, French physicist Henri Becquerel (1852-1908) accidentally discovered radioactivity while investigating phosphorescence in uranium salts. In 1903, he shared the Nobel Prize in physics with Pierre and Marie Curie in recognition of his discovery and their study of natural radioactivity.

The SI unit for radioactivity, the becquerel (Bq), is named after him.

HISTORY: **MARIE CURIE**

Marie Curie (1867-1934) discovered that the substance, thorium, was "radioactive," a term she created. Curie, along with her husband Pierre, discovered the radioactive elements polonium and radium.

In 1903, the Curies and Henri Becquerel were awarded the Nobel Prize in physics for the discovery and exploration of natural radioactivity. In 1911, Curie received her second Nobel Prize in chemistry for isolating radium and determining its atomic weight. She was the first woman to receive a Nobel Prize and the only woman, to this day, to receive two Nobel prizes.

Po

POLONIUM

Atomic #84

Polonium was discovered by Marie and Pierre Curie in 1898, thus the element was named for Marie Curie's birth country of Poland.

Characteristics:

- Is a very rare natural element, found in extremely small amounts in uranium ores
- Is mainly used as a source for neutrons and has specialty uses in eliminating static electricity in machinery and removing dust from photographic film

Polonium has 25 known isotopes. Its most common isotope, Po-210, has a half-life of only 138 days. The radioactive decay of Po-210 produces a lot of heat (140 watts per gram).

Rn

RADON

Atomic #86

Radon comes from the radioactive decay of the element radium.

Characteristics:

- Is radioactive; the isotope with the longest half-life is radon-222 with a half-life of only four days
- Is a colorless radioactive gas at a normal room temperature of about 70-75°F
- Glows with a yellow color when cooled to its solid state
- Is emitted naturally, in some regions, from the soil and rocks and can sometimes build up in people's homes

The World Health Organization estimates that 15% of all lung cancer cases are caused by exposure to radon. Radon test kits are available to check for radon accumulation in homes, especially basement levels.

Ra

RADIUM

Atomic #88

Radium was discovered by Marie and Pierre Curie in 1898. Its name comes from the Latin word "radius" meaning "ray."

Characteristics:

- Is a highly reactive metal
- Is a brilliant, white metal in pure form but blackens when exposed to air
- Occurs naturally in the environment from the decay of uranium and thorium

Its most stable isotope, radium-226, has a half-life of about 1,600 years.

Pure radium and some of its compounds glow in the dark. As a result, radium was used in the mid 1900s in a luminous paint on the hands and numbers of watches to make them glow in the dark until the risks of radium exposure were known.

The radioactive decay of an unstable nucleus may release several types of radiation including alpha radiation, beta radiation, and gamma radiation.

Quick Fact

Radiation may be used in medicine to treat disease and to look inside the body to diagnose medical problems. Radiation has proven useful to kill cancer cells by causing mutations (defects) in DNA, thus preventing the cancerous cells from being able to grow and divide.

ALPHA (α) RADIATION (ALPHA PARTICLES): radiation composed of helium-4 nuclei (having a nucleus that is the same as helium with two protons and two neutrons).

- Alpha radiation travels very slowly and only a very short distance through air and cannot penetrate skin or even a thin sheet of paper.
- Alpha particles are not radioactive. After losing their energy, they attract two electrons to become a helium atom.

BETA (β) RADIATION (BETA PARTICLES): radiation composed of electrons, emitted from an unstable nucleus, that are in high velocity.

- Beta radiation can travel several meters through air but is stopped by solid materials.
- Beta particles can penetrate human skin, but clothing often helps to block most beta particles.
- If the release of a beta particle does not get rid of the extra energy in an unstable nucleus, the nucleus will often release the rest of the excess energy in the form of gamma-rays.

GAMMA (γ) RADIATION (GAMMA-RAYS): radiation composed of high-energy photons in the form of electromagnetic radiation.

- Gamma radiation is able to travel many meters in air and easily penetrates most materials, including several centimeters through human tissue.
- Gamma radiation frequently accompanies the emission of alpha and beta radiation.

HISTORY: JOHANNES WILHELM ("HANS") GEIGER

Hans Geiger (1882-1945) was a German physicist known for his work on radioactivity. In 1928, with fellow physicist Walther Müller, he developed a device to measure radioactive emissions, which became known as the **Geiger Counter**. The two worked to improve the device's sensitivity, performance, and durability, thus creating a tool that is used in laboratories around the world today.

NUCLEAR ENERGY

Nuclear reactions are changes that occur in the structure of atomic nuclei. The energy that results from nuclear reactions is called nuclear energy or atomic energy. Nuclear energy is released from atoms in two different ways: nuclear fission and nuclear fusion.

NUCLEAR FISSION: a nuclear reaction that occurs when an atomic nucleus splits into two smaller parts (nuclei), usually about the same size. When this happens, energy is released.

- Uranium nuclei can be split easily by bombarding them with neutrons.
- Once a uranium nucleus is split, multiple neutrons are released, initiating a chain reaction as other uranium nuclei are split.

U

URANIUM

Atomic #92

Uranium was first identified in pitchblende ore in 1789. It was named after the planet Uranus, which had been discovered around that time.

Characteristics:

- Is the heaviest naturally occurring element on the earth, except for minute traces of neptunium and plutonium
- Is highly radioactive, toxic, and carcinogenic
- Has 16 isotopes, all of which are radioactive

Uranium's radioactivity was first detected by Henri Becquerel in 1896. Today, it is primarily used in nuclear fuels and explosives. Uranium, specifically the isotope uranium-235, is the principle element used in nuclear reactors and in certain types of atomic bombs.

Uranium compounds have been used for centuries as additives in glass, to which they impart interesting yellow and green colors and fluorescent effects.

NUCLEAR FUSION: a nuclear reaction that occurs when the nuclei of atoms join to make a larger nucleus. Again, energy is given off in this reaction.

- Nuclear fusion only occurs under very hot conditions.
- The sun and all other stars create heat and light through nuclear fusion. In the sun, hydrogen nuclei fuse to make helium.

Quick Fact

The hydrogen bomb uses nuclear fusion. Hydrogen nuclei fuse to form helium and, in the process, release huge amounts of energy and create a massive explosion.

HISTORY: **ENRICO FERMI**

Enrico Fermi (1901-1954) was an Italian physicist most noted for his work on beta decay, the development of the first nuclear reactor, and for contributions to the development of quantum theory. He worked on the Manhattan Project during World War II to produce the atomic bomb, though he warned of its power.

In 1938, Fermi won the Nobel Prize in physics for his work on radioactivity.

Pu

PLUTONIUM

Atomic #94

Plutonium was discovered in 1940 at the University of California at Berkeley and was named after the planet Pluto. The element's discovery, however, was kept classified by the government until 1946.

Characteristics:

- Is a very heavy, silvery metal in pure form
- Is a rare radioactive element; found in minute amounts (0.72%) in uranium ore
- Used mainly as fuel for nuclear reactors and nuclear bombs

Plutonium was produced in large quantities in the U.S. during World War II, as part of the Manhattan Project to create the atomic bomb.

Over one-third of the energy produced in most nuclear power plants comes from plutonium.

NOTES

MAN-MADE ELEMENTS

The elements above atomic number 92 are known as transuranic or transuranium elements and do not occur naturally on the earth. Most of these heavier elements have been made by bombarding the element uranium with neutrons or other particles in a cyclotron, a machine used to accelerate charged particles to high energies.

Many of the man-made, transuranic elements are named for important chemists or physicists:

Atomic #	Element	Symbol	Named for ...
99	Einsteinium	Es	Albert Einstein, the famous scientist who developed the Theory of Relativity
101	Mendelevium	Md	Dmitri Mendeleev, who developed the periodic table
102	Nobelium	No	Alfred Nobel, who commercialized dynamite and endowed the Nobel Prizes for physics, chemistry, medicine, literature, and peace
103	Lawrencium	Lr	Ernest O. Lawrence, inventor of the cyclotron
104	Rutherfordium	Rf	Ernest Rutherford, who helped develop the modern understanding of the atomic nucleus
106	Seaborgium	Sg	Glenn Seaborg, known for his work in the separation and purification of plutonium and for proposing the “Actinide” concept for reorganizing the periodic table
107	Bohrium	Bh	Niels Bohr, who proposed a model of atomic structure that explained the role of the electron
109	Meitnerium	Mt	Lise Meitner, co-discoverer of nuclear fission

SECTION XIV: LABORATORY EQUIPMENT

OBJECTIVES

- Identify basic laboratory equipment.
- Describe how certain laboratory equipment is used to conduct experiments.

BASIC EQUIPMENT

Chemists use many different types of equipment in the laboratory to perform their experiments, including glassware, measuring tools, and many specialized instruments. The following are some examples of common laboratory equipment:

Beaker: a wide container with a flat bottom made of glass or plastic.

- A beaker is a simple container used to mix, heat, or simply hold liquids (and sometimes solids).
- If the beaker is *graduated* (marked with or divided into degrees or measurements), it can give approximate measurements of liquid volumes.

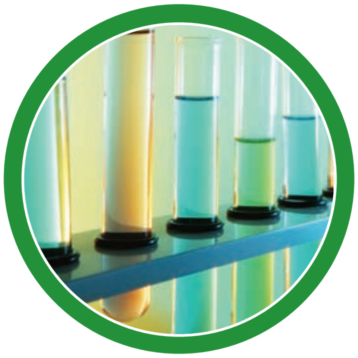
Flask: a glass container with a "neck" that widens to a flat or rounded base and is used to measure, heat, or store liquids.

- The neck of the flask allows scientists to attach a clamp to it or place a stopper in it.
- **Erlenmeyer flask:** a flask with a cone-shaped base that is used most often to hold liquids during a filtration or titration process.
  - A liquid contained in an Erlenmeyer flask will heat quickly because the wide surface area of the flask's bottom provides a greater surface for heating.

- **Florence flask (boiling flask):** a round (spherical) flask that may have a rounded or flat-bottomed base.
  - A liquid contained in a Florence flask will heat evenly because the round bottom distributes the heat around the flask.
  - Florence flasks are generally stronger than other flasks because they are often used to boil liquids for distillation processes. Therefore, they must be able to withstand extreme temperature changes.

- **Volumetric flask:** a flask with a pear-shaped base and a long neck that is generally fitted with a stopper.
  - Volumetric flasks are used to precisely measure and prepare liquids of fixed volumes. Each volumetric flask is calibrated to a specific volume, such as 50 mL, 100 mL, or 250 mL.
  - A single line around the long neck marks the point where the flask should be filled.

Test tube: a cylindrical glass tube that has a rounded, u-shaped bottom and is usually open at the top.

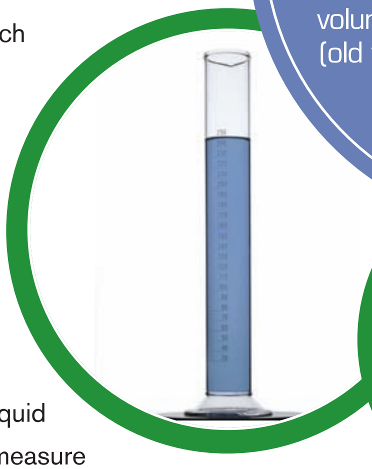


- Test tubes are used to hold or heat small amounts of a substance during laboratory experiments.
- A stopper may be placed in the opening of a test tube to temporarily store the substance it contains.

MEASURING LIQUID VOLUMES

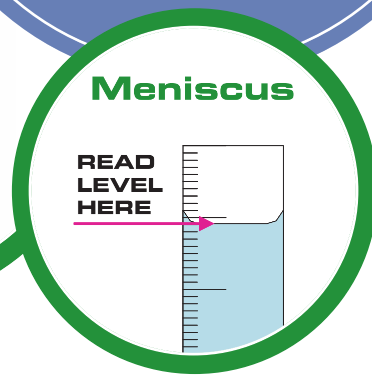
Graduated cylinder: a tall, cylindrical container with a wide, flat base that is used to measure the volume of a liquid. The top of the cylinder is open and usually has a small curved lip, so liquids can be poured easily from the cylinder.

- It was named a "graduated" cylinder because it has markings along the side to indicate how much volume is being measured.
- Graduated cylinders are made in many different sizes, generally ranging from 10 mL to 2,000 mL.
- In most graduated cylinders, the surface tension of the liquid being measured will often create a concave surface called a *meniscus.* To determine the volume of the liquid in a graduated cylinder, scientists take the measure from the bottom of the meniscus (see image at right).



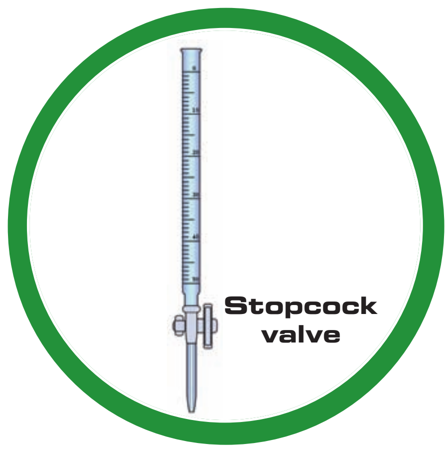
Quick Fact

Graduated cylinders are often used to measure the volume of small solids that have an irregular shape. First, the cylinder is partially filled with water. Then, the solid is placed inside, and the level of the water rises. The volume of water that was displaced (old volume minus the new volume) is equal to the volume of the solid.



Buret (or burette): a long glass cylinder used to measure and dispense a specific volume of liquid very accurately. The cylinder is open at one end and has a stopcock valve attached to the other end.

- The stopcock valve controls the amount of liquid that flows from the buret.
- Burets are generally used for titrations (see Laboratory Analysis section). During a titration, the liquid in the buret is typically dispensed into an Erlenmeyer flask containing some other substance or solution.
- Liquids in a buret should not contain bubbles because bubbles affect measurements. Therefore, once a buret is filled with a liquid, you should open the stopcock valve to allow a few milliliters of the liquid to flow from the buret into a beaker or flask. After closing the stopcock valve, gently tap the sides of the buret to release any remaining bubbles from the liquid.



MEASURING MASS

Mass is measured using devices known as balances. Balances measure mass by comparing an object of unknown mass to an object of known mass, so the measurement is not affected by gravity.

Triple-beam balance: a device used to measure the mass of an object. An object is placed on a platform connected to three beams marked with certain measurements. Each beam has a different size weight attached to it. The weights are moved across the beams until the pointer at the opposite end shows that the weights and the object on the pan are in balance. The locations of the weights on the beam indicate the mass of the object.

Many science laboratories use instruments called analytical balances or *single-pan electronic balances* because of their ability to measure with great precision.

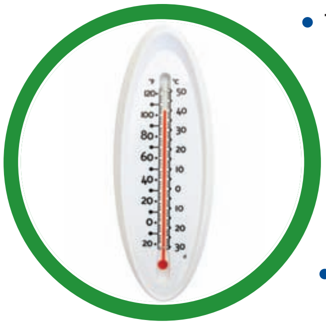
Analytical Balance

When weighing a chemical using a balance, scientists often weigh the empty container first and then weigh the container with its contents.

- The weight of the empty container is called the *tare* weight.
- Some balances provide an option for automatically taring the vessel, so the additional weight of the contents can be read directly on the display.

MEASURING TEMPERATURE

Thermometer: a device used to measure temperature based on the expansion of a fluid.



- Traditional thermometers are thin graduated glass tubes with a bulb at one end to hold the fluid. As the fluid is heated, it expands and is pushed up the tube. Because the tube is so thin, even a small increase in volume causes the fluid to rise noticeably.
- The markings of a thermometer are generally determined based on two fixed reference points, usually the freezing point and boiling point of water.
- The degrees of measurement depend on the temperature measurement scale being used, whether Fahrenheit, Celsius, or Kelvin.

HISTORY: THERMOMETER

In 1593, Galileo Galilei invented a water thermometer. Because water freezes, water thermometers could not measure temperatures less than the freezing point of water.

The first mercury thermometer was invented in 1714 by Daniel Gabriel Fahrenheit. Then in 1724, he introduced the temperature scale that bears his name, the Fahrenheit scale.

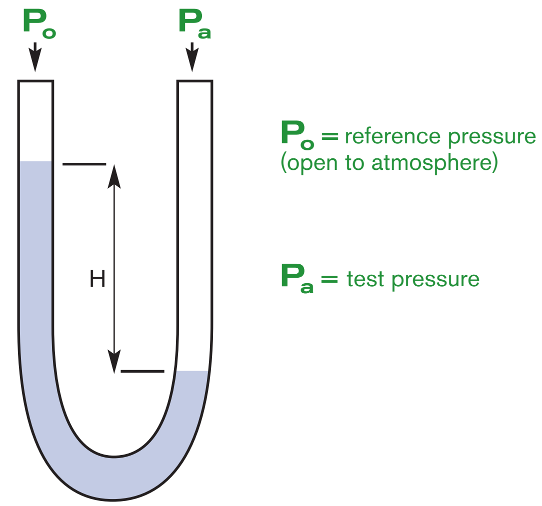
Today, most thermometers use alcohol because, like mercury, alcohol also has a freezing point below the freezing point of water. Since mercury can be dangerous to human health if the thermometer is broken, alcohol is considered to be a safer alternative.

MEASURING PRESSURE

As humans, we often need to measure different forms of pressure, including tire pressure, blood pressure, and atmospheric pressure. To obtain these measurements, scientists use a variety of instruments.

Manometer: a device used to measure the pressure of a fluid.

- Manometers often measure pressure based on differences in the height and position of a column of liquid in a "U" tube.
- The legs of the manometer are connected to separate sources of pressure, causing the liquid to rise in the leg with the lower pressure and drop in the other leg.
- One leg of the tube is a reference leg, often left open to the atmosphere. The other leg of the tube is the measuring leg.
- Manometers are usually used to measure pressures close to atmospheric pressure.



Barometer: a device (a type of manometer) used to measure atmospheric pressure.

- When the water or mercury level rises in a barometer, the air pressure is increasing. When the water or mercury level falls, the air pressure is decreasing.
- Mercury barometers are commonly used in weather reporting.

HISTORY: EVANGELISTA TORRICELLI

Evangelista Torricelli (1608-1647) was an Italian physicist and mathematician. While investigating vacuums, Torricelli developed the first mercury barometer in the early 1640s.

- He determined that the height of mercury was only $^1/_{14}$ the height of water in a water barometer because mercury is fourteen times as dense as water.
- He noticed that the level of mercury varied from day to day and concluded that the difference was caused by changes in atmospheric pressure.
- He also determined that the space above the mercury in the barometer must contain a vacuum.

The "torr," a unit of pressure, is named after him.

Most **pressure gauges** used today are aneroid (meaning "without fluid") and use a coiled tube, called a "Bourdon" tube.

- A Bourdon tube flexes like a spring when pressure is exerted on it.
- As the tube coils or uncoils, it turns a pointer on the face of the gauge to mark the pressure.

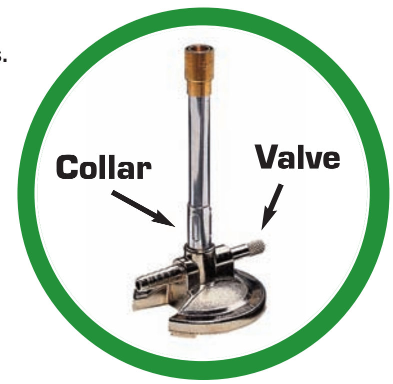
TRANSFERRING LIQUIDS

Pipette: a long, skinny tube with a bulb at one end that is used to extract and transfer liquids from one container to another.

- Liquid is drawn up into the pipette by suction when the bulb, at the opposite end, is squeezed and then released.
- Pipettes are often used to measure (and transfer) small volumes of liquid.

HEATING MATERIALS

In commercial laboratories, the most common ways of heating are with a hot plate, in a temperature controlled oven, or over a Bunsen burner. The first two methods are generally preferred because the heat is provided by an electric source. The lack of a flame reduces the risk of a fire. In classroom laboratories, hot plates and Bunsen burners are commonly used.



Bunsen burner: a gas burner that produces a single, steady flame for laboratory experiments.

- The flame burns at the top of a vertical metal tube connected to a natural gas source.
- A valve on the Bunsen burner controls the amount of gas that flows into it.
- As more gas flows into the burner, the flame increases in size.
- A collar controls the amount of air that mixes with the fuel and can be adjusted to control the quality of the flame.
- The ideal flame is bluish in color, not yellow or smoky.

HISTORY: **ROBERT BUNSEN**

Robert Bunsen (1811-1899) improved the heating burner that bears his name in 1855. Up until his improvements, the flames that were produced had been smoky, flickered excessively, and did not produce much heat.

NOTES

SECTION XV: CHEMICAL SAFETY

OBJECTIVES

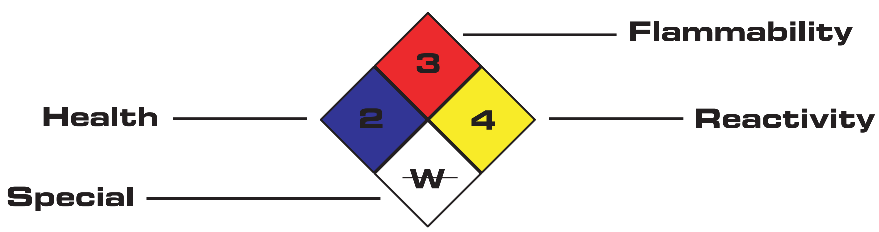
- Explain the importance of chemical safety and where to find chemical safety information.
- Identify common chemical safety and protective equipment symbols.
- List and describe basic laboratory safety guidelines and tips.

Chemicals have many beneficial uses, but some can also be hazardous if mishandled or misused. For this reason, it is very important that anyone using chemicals understand how to safely store, handle, and use them.

WHERE TO FIND CHEMICAL SAFETY INFORMATION

The first place to look for information about chemical safety is on the label of the container. Product container labels include key warnings about storage, handling, and if necessary, first aid and other emergency information.

Most commercial chemical containers will have a label similar to the figure below. The diamond is divided into four sections, each providing a hazard rating from 0 to 4, with 0 indicating no hazard and higher numbers indicating increasing degrees of danger.



- **Health** rating: indicates the degree of harm that exposure to the substance would cause an individual.
- **Flammability** rating: indicates the likelihood of the substance to vaporize, ignite, and burn.
- **Reactivity** rating: indicates the likelihood of the substance to release energy by chemical reaction or explosion.
- **Special Warning** section: provides any additional cautionary information. The symbol shown above (a “W” with a line through it) signifies that the chemical reacts with water.

Another information source is the **material safety data sheet (MSDS)**, available for most chemicals. The MSDS includes:

- Detailed information about the chemical and physical properties of the product.
- Potential hazards associated with the substance (health, storage cautions, flammability, radioactivity, reactivity, etc.).
- Emergency actions that should be taken if an accident occurs.
- Storage and handling precautions.
- Safe disposal information.

An MSDS also often includes manufacturer identification and the manufacturer's address, along with emergency phone numbers.

Although MSDSs are designed for workplaces and emergency personnel, anyone can benefit from having this important product information available. MSDSs are available for most substances, including common household products like cleaners, gasoline, pesticides, certain foods, drugs, and office and school supplies. Familiarity with MSDSs allows us to take precautions when using potentially dangerous products. They also alert you to hazards.

NOTES

WARNING SYMBOLS

Anyone working with chemicals should also become familiar with the common warning symbols found on chemical containers and around laboratories. These symbols may seem frightening, but that is the point! While chemicals have many benefits, warning symbols are used to catch your attention and make you aware of any possible dangers. These symbols help ensure than people use these products responsibly.

The *skull and cross-bones* symbol is used to indicate danger, generally referring to the presence of a poisonous substance. Do not let such a substance come in contact with your skin, do not eat or drink the substance, and do not inhale its vapors.

Quick Fact

Exposure to low temperatures, such as contact with liquid nitrogen or dry ice, may damage the skin just as much as a heat burn. Appropriate hand protection must be worn when working with either hot or cold substances.

The *fire* symbol indicates the presence of an open flame or flammable substance. When working near an open flame, tie back or cover loose hair and avoid loose or bulky clothing. When working with a flammable substance, be sure to avoid anything that could ignite the substance, such as electrical sparks or a hot surface.

The *corrosive* symbol usually indicates the presence of a strong acid or base, which can destroy or cause major damage to other substances. Avoid contact with skin, eyes, and clothing and do not inhale the vapors.

The *explosive* symbol indicates the potential for an explosive situation. Read and follow instructions carefully.

The *biohazard* symbol warns of a biological substance that is dangerous to humans or the environment. Syringes and other medical devices that have come in contact with bodily fluids and could potentially carry harmful diseases are labeled with the biohazard symbol.

Quick Fact

In 2007, a new radiation warning symbol was released by the United Nations. The symbol includes a skull and cross-bones and a person running away from the radiation "trefoil." This symbol was designed to provide a more effective warning of ionizing radiation.

The *radiation* symbol (the "trefoil") indicates a radioactive substance. Instructions should be followed carefully to avoid harmful exposure.

Lasers can damage skin, and even low-powered lasers can cause severe damage to eyesight.

PROTECTIVE EQUIPMENT SYMBOLS

Other symbols are used to indicate the proper protective equipment that you should wear when working with certain chemicals or in certain environments.

The *respiratory protection* symbol indicates that a protective mask or other facepiece is required. The symbol is used in areas where a person may be exposed to contaminated air.

The *hand protection* symbol indicates when gloves must be worn. Be sure to determine what type of gloves are needed. Thermal gloves are used for handling hot materials. Leather gloves are used for handling rough or scratchy materials. Nitrile gloves are generally used to protect against chemical solvents and potentially infectious substances.

The *protective footwear* symbol indicates that protective boots or shoes must be worn. Safety footwear protects feet against injuries, such as impact from falling objects, puncture wounds, or hazardous liquids spilled on the floor.

The *eye protection* symbol indicates that safety goggles or other eye protection is required. Be sure to determine what type of eye protection should be worn before entering an area marked by this symbol. *Safety goggles* are made to fit completely around the eyes and protect people from splashing liquids and some projectiles. *Industrial safety glasses* protect the eyes from sharp, flying projectiles.

The *face protection* symbol indicates that full face protection is required, such as a face shield. Face shields protect your entire face from splashing liquids or other pieces of materials that fly through the air. You should also wear safety glasses or goggles under the face shield for the best protection.

NOTES

GENERAL SAFETY RULES

Always follow directions when performing laboratory experiments. Scientists take special care to choose the safest chemical to do the job, follow safe practices, and understand all hazards.

General safety rules for storing, handling, and using chemicals are guided by three basic principles – keeping people safe around chemicals, keeping chemicals away from each other, and practicing good hygiene. Some general rules to remember are:

Quick Fact

All National Challenge participants receive lab coats!

- **Wear appropriate clothing when working in the laboratory.** Remove jewelry, such as dangling necklaces, chains, and bracelets. Roll up or secure long sleeves, remove loose garments, and never wear open-toe shoes. Wear glasses instead of contact lenses or notify your teacher if you are wearing contact lenses. Tie back or cover long hair, especially near an open flame.
- **Wear a lab coat to keep chemicals off your regular clothes.** Lab coats should be buttoned all the way and sleeves should be rolled down to completely cover your arms. Lab coats should be removed when you leave the lab to avoid contaminating public places or your home.
- **Wear the proper type of gloves for the task at hand.** Different types of gloves, made from different materials, are available for working with certain substances. For example, thermal gloves, rubber gloves, and nitrile gloves are all made for specific tasks. You should also be sure to remove gloves before touching any common surfaces that other people may touch when they are not wearing gloves (such as door handles).
- **Wear safety goggles to protect your eyes.** Safety goggles, unlike regular glasses, cover your eyes completely. Many eye injuries occur to the sides of your eyes, and regular glasses do not provide protection on the sides. Safety goggles and other splash protection are especially important whenever mixing or pouring chemicals
- **Know the location and proper use of safety equipment.** Safety equipment includes the fire extinguisher, safety shower, eye-wash station, fire blanket, first aid kit, and fire alarm. *Notify your teacher immediately of any injury or chemical spill.*
- **Keep your face away from the mouth of a container that holds chemicals and do not smell any chemical directly.** Chemists use a process called "wafting" to smell chemicals. An open chemical container is held at least 18 inches away from your nose. You then use your free hand to "fan" the vapors escaping from the bottle toward your nose. When testing for odors, use a wafting technique to direct the odors to your nose.
- **Do not eat or drink in the laboratory.** Chemicals in the lab, and even particles in the air, may accidentally mix with your food or drink and be ingested. Likewise, food and drinks may contaminate the chemicals in the lab.
- **Do not goof around when conducting experiments.** Chemistry is fun, so be sure to pay attention and focus on the experiment you are performing. Goofing around in the laboratory and not paying attention may have unpleasant consequences.

- **Do not play Mad Scientist!** Never mix chemicals just to "see what happens." Pay attention to the order in which chemicals are to be added to each other and follow such instructions carefully. The chemical reactions and explosions you may see on TV science programs may look cool, but they are created by experts who know how to run the reaction safely and take proper precautions. Do not try to create these reactions or explosions at home or in your class!

WHEN STORING CHEMICAL SUBSTANCES:

- **Label all storage areas and containers clearly.** Be sure that these labels are visible and updated as needed.

- **Avoid storing chemicals in areas that are difficult to see.** Avoid storing chemicals on top of cabinets or on shelves above eye level, so people can easily see chemical labels.

- **Reactive chemicals, such as acids and bases, should never be stored near each other.** Storing acids and bases separately prevents accidental mixing. Therefore, you should not store chemicals alphabetically, as reactive chemicals may end up next to each other in the storage area.

Think About It...

Do you follow the same safe chemical storage practices at home that you follow in the laboratory? What do you have stored under your kitchen sink or in utility room cabinets? Which items are acids? Which are bases?

WHEN POURING LIQUIDS FROM ONE CONTAINER TO ANOTHER:

- **Keep the label of a chemical bottle against the palm of your hand.** By following this procedure, if any of the substance drips from the bottle, it will only touch the opposite side of the bottle. That way, the next person to use the bottle will not risk getting any of the chemical on his/her hands.

- **Use a glass stirring rod to pour a liquid into a beaker.** You will pour the liquid down the glass rod into the beaker to avoid splashing.

- **Pour the liquid into a beaker or graduated cylinder first when transferring a liquid to a small-mouthed vessel, like a test tube.** Beakers and graduated cylinders typically have a pouring spout that will give you better control when transferring the liquid into the smaller vessel.

- **When an acid must be mixed with water, always add the acid slowly to water.** Mixing the other way around, by adding water to an acid, could result in a violent reaction.

WHEN HANDLING GLASSWARE:

- **Do not place hot glassware directly on a table.** Always place some kind of insulating pad between the glass and the table surface.

- **Allow plenty of time for hot glass to cool before touching it.** Remember, hot glass may *look* like it has cooled, but *it can still be very hot!*

WHEN HEATING SUBSTANCES:

- **Never leave a flame or other heating instrument unattended.**
- **Point the mouth of the test tube away from yourself and other people.** When heating a test tube, the substances inside may quickly shoot out of the tube. Be sure the mouth of the tube is pointed away from people and never look into a container that is being heated.
- **Never heat a closed container.** Generally an increase in temperature causes an increase in pressure. Inside a closed container, pressure may increase to the point that the container breaks or bursts apart.
- **Do not bring any substance in contact with a flame.** Unless you are specifically instructed to do so, you should never touch a substance directly to a flame. Heating a substance near a flame and actually touching a substance to a flame can have very different results.
- **Avoid extremes when heating substances.** Do not put cold glassware on hot surfaces.
- **Use tongs when removing glassware from a heat source to avoid burns.**

Quick Fact

Special high-temperature porcelain containers called crucibles are often used to melt metals, dry powders, or accomplish other high-temperature tasks.

- Tongs should be used to hold onto hot containers or crucibles.
- Special “test tube holders” should be used to hold a test tube over the flame of a Bunsen burner.

WHEN CLEANING UP, BE SURE TO:

- **Turn off all heating instruments and disconnect any other electrical equipment.**
- **Return all materials to their proper places.** All equipment should be cleaned and stored in the proper places. Do not return chemicals to the original containers. Your teacher will tell you what to do with the unused chemicals and how to dispose of any other materials.
- **Do not casually dispose of chemicals down the drain!** Some chemicals can be washed down the drain, while others require a different method of disposal. Make sure you ask your teacher what substances can go down the drain. If a chemical can go in the sink, be sure to wash it away completely. Otherwise, if you have not completely washed away one substance, another substance you pour down the drain later may react with the leftovers.
- **Clean and dry your work area.** Do not leave water on the counter or floor.
- **Wash your hands with soap and water after completing an experiment.** Be sure not to touch your eyes or mouth until you have washed your hands.

SECTION XVI: LABORATORY SEPARATIONS

OBJECTIVES

- Define a separation process and provide examples of types of separation processes.
- Identify and distinguish between different types of separation processes.
- Describe specific examples of the uses of different separation processes.

When conducting research and experiments, scientists often need to separate components of a mixture. To do so, they use laboratory separation processes.

- **Separation process:** a process that divides a mixture of substances into two or more different parts by making use of the different properties of those parts.

Separation processes may use physical or chemical means to seprate parts of a mixture (see section on Physical and Chemical Separations), though most common separation processes are physical processes. The following sections describe a variety of physical separation processes commonly used in laboratories.

DISTILLATION

Distillation is a method of separating a liquid mixture based on the differences between the boiling points of the parts of the mixture.

- A liquid mixture is heated to the boiling point of one part of the mixture, causing that part of the mixture to evaporate.
- The vapor (evaporated part of the mixture) is collected in a separate container and cooled, so it condenses into its pure liquid form. The purified liquid is called the **distillate.**

EXAMPLE:

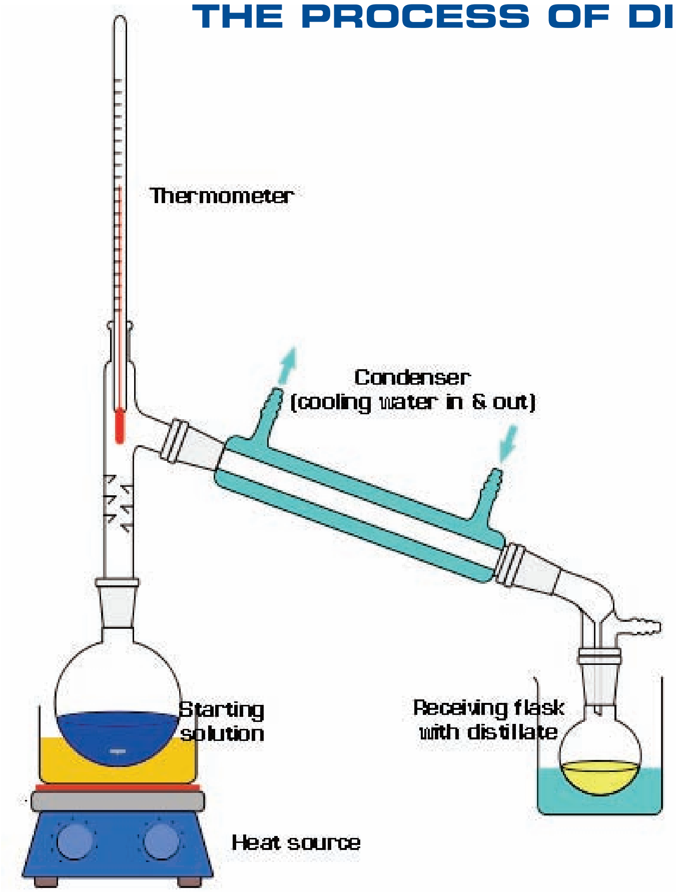
Distillation can be used to purify sea water. Heating sea water to the boiling point of water causes the water to evaporate. When the water evaporates, the salts and other components of the sea water are left behind. The water vapor is then collected in a separate container and cooled. The collected vapor condenses into pure liquid water.

Quick Fact

Fractional distillation is a type of distillation process commonly used in modern refineries.

Fractional distillation is used to separate components of a mixture with similar boiling points. Therefore, to separate those different parts, the temperature of the mixture is increased by small amounts, so each part will evaporate independently.

THE PROCESS OF DISTILLATION



DISTILLATION DEVICE

The image at left shows one type of distillation set up.
(Not all distillation processes are set up in the same way.)

The distillation process begins when the heat source is turned on and begins to heat the starting solution. When the starting solution reaches the boiling point of a liquid in the solution, that liquid will evaporate. The vapor will move up the tube and into the condenser. The condenser will cool the vapor to liquid form. The purified liquid (distillate) will then flow into the receiving flask.

CENTRIFUGATION

Centrifugation is a process for separating a mixture based on the size or density of the particles by spinning the mixture quickly around a center axis.

- A laboratory centrifuge is the device used to spin mixtures at high speeds.
- During centrifugation, a test tube containing a sample of a mixture is placed horizontally inside a centrifuge. The tube is spun around to increase the gravitational force, causing the denser particles to collect at the bottom of the tube. Once the particles collect, the liquid is quickly and carefully poured from the top of the test tube or removed with a pipette.
- Centrifugation is often used to separate denser solids, such as a precipitate, that are suspended in a liquid.

Quick Fact

Centrifugation is used to separate the heavy cream from whole milk.

The process is also used to separate the components of our blood for certain blood tests. Centrifugation causes the dense red blood cells to collect at the bottom of a test tube, with the white blood cells and platelets forming the next layer, and the lighter plasma resting on top.

DECANTATION

Decantation is a process by which the denser particles in a solution are separated from a liquid. Once the denser particles settle to the bottom of the container, the liquid is carefully poured out, leaving the solid particles behind.

- Decantation is often used after centrifugation or a precipitation reaction to remove the separated components of the mixture.
- To properly conduct decantation, a small amount of the solution is often left in the container, so none of the solid particles escape with the decanted liquid.

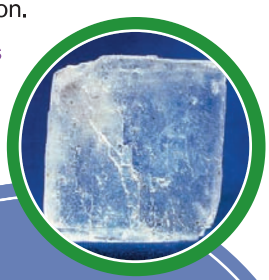
CRYSTALLIZATION

Crystallization is the process by which solid crystals are formed from a homogeneous solution. The process is based on solubility and supersaturation (see Chemicals by Volume - Solutions section).

- As a supersaturated solution cools, the solubility of the solute decreases, causing it to crystallize out of the solution. The crystals will grow in number and size and can eventually be removed by a solid-liquid separation method, such as filtration.

  EXAMPLE:

  If you add sugar to water, and keep adding sugar until no more crystals can be dissolved, the solution is saturated. More crystals can be dissolved into the saturated solution by heating it (since solubility of solutes generally increases with an increase in temperature). The rise in temperature allows more sugar crystals to dissolve in the solution creating a supersaturated solution. As you let the solution cool, the 'excess' sugar begins to crystallize out of the solution.



Quick Fact

Reducing temperature is one way to reduce the solubility of a solute, thus forming crystals from a supersaturated solution. However, the solubility of a solute can also be decreased by initiating a chemical reaction or changing the pH of the solution

- Crystallization is often used to remove salts from solutions. It is also used industrially in the production of pharmaceuticals.

ABSORPTION AND ADSORPTION

Absorption is a process by which matter takes in another substance. The absorbed substance is distributed throughout the absorbing matter and is incorporated into the matter's structure.

- A kitchen sponge soaking up water is an example of absorption.

Adsorption is a process by which a substance clings to the surface of a solid or liquid. The adsorbed substance (called the adsorbate) gathers on the surface but does not enter the solid or liquid.

- During this process, the adsorbate creates a film on the surface of the solid or liquid adsorbent.
- Charcoal is often used in water purification systems because it can adsorb many contaminants. The adsorbed material becomes bound to the surface of the solid.

SOLVENT EXTRACTION

Solvent extraction is a process of removing a component from a mixture by adding a solvent. This process utilizes the differences between the solubility of the components of the mixture. The component to be separated should be able to dissolve easily in the solvent.

EXAMPLE:

Solvent extraction is comonly used for recovering vegetable oil. Vegetable oil is produced by flaking and crushing the oilseed. Some of the oil can be recovered directly from the crushing process. However, much of the oil remains in the seed and is recovered by contact with a solvent (hexane). The oil is highly soluble in hexane and can be separated from the crushed seed for further processing.

Quick Fact

A coffee machine uses water as a solvent to remove the soluble parts of the coffee, leaving the insoluble parts of the coffee grounds behind. This form of extraction is often referred to as solid-liquid extraction.

NOTES

CHROMATOGRAPHY

Chromatography is a group of separation processes used to separate and analyze mixtures based on the phases of the components. During chromatography, a mixture (the "mobile" phase) flows over a material (the "stationary" phase). Components of the mobile phase flow through the stationary phase at different rates, causing the components to separate.

- There are various types of chromatography, including gas chromatography, column chromatography, and thin layer chromatography.
- Paper chromatography is a method that can be easily demonstrated in a classroom.

EXAMPLE:
Black ink is a mixture of many different colors. If a drop of black ink is placed near the bottom of a piece of chromatography paper and the bottom edge of the paper is placed in a solvent, the paper will absorb the solvent. As the solvent moves up the paper, it carries the mixture of black ink up through the paper as well. As the ink moves through the paper, the different color components will move through the paper at different rates, leaving bands of color along the paper based on how far that color component can travel through the paper.

NOTES

SECTION XVII: LABORATORY ANALYSIS

OBJECTIVES
- Distinguish between qualitative and quantitative analysis.
- Identify and describe types of qualitative analysis.
- Identify and describe types of quantitative analysis.

Analytical chemistry is a branch of chemistry focused on understanding the chemical composition (makeup) of matter and creating devices and designing experiments that can take measurements of chemical compositions.

QUALITATIVE VERSUS QUANTITATIVE ANALYSIS

Analytical chemistry measurements may be qualitative or quantitative.

- **Qualitative analysis** is used to determine whether or not a certain substance is present in a sample. Qualitative analysis shows "what" is in a sample.
  – Qualitiative analysis often involves classifying or categorizing data.
- **Quantitative analysis** is used to determine the amount of a certain substances in a sample. Quantitative analysis shows "how much" of something is in a sample.
  – Scientists often conduct quantitative analysis after conducting qualitative analysis.
  – Quantitative data can be counted and should allow for scientists to apply their findings to a larger sample or population.

TYPES OF QUALITATIVE ANALYSIS

Scientists use different qualitative analysis techniques based on what type of substance they are trying to find in a sample.

FLAME TESTS

A **flame test** is used to determine which atoms, specifically which metal atoms, are present in a solution.

- A clean wire is dipped into a solution containing an unknown metal. The wire is then placed into a burner flame. Once in the flame, the metal in the solution will produce a characteristic color that is emitted from the flame. Sodium, for example, creates a bright yellow flame.

Quick Fact

Sodium is present in many compounds. In order to determine what other metal is present in the compound, scientists often view the flame through a special colored glass. Blue cobalt glass is able to filter out the yellow color produced by sodium, showing only the color produced by the other metal.

SPECTROSCOPY

Spectroscopy is used to determine the molecular structure, energy level, and chemical composition of a substance by identifying the way its atoms abosorb and emit energy.

- A device called a spectrometer can be used to record the spectrum of light emitted (or absorbed) by a particular substance. The light can then be used to determine the chemical composition of the substance because many elements emit specific "spectral lines" of light.
- Different types of spectroscopy commonly used in laboratories are: near infra-red spectroscopy, Fourier transform, and X-ray spectroscopy.
- Some types of spectroscopy can be quantitative, such as atomic-absorption spectroscopy.

Quick Fact

Spectroscopy is used by astronomers to determine the chemical makeup of stars.

TYPES OF QUANTITATIVE ANALYSIS

Scientists use different quantitative analysis techniques to measure the amount or the concentration of a certain substance in a sample.

TITRATION

Titration is an analytical method used to determine the concentration of a certain substance in a solution. During a titration, a solution of unknown concentration is added to a solution of known concentration (called the standard solution or titrant) to create a chemical reaction.

- Titration is useful to determine the concentrations of acids and bases in a solution.
- To conduct a titration, a specific volume of the solution of unknown concentration is poured into an Erlenmeyer flask. The titrant, known to react with the solute in the other solution, is added to a buret (see Laboratory Equipment section). Once the buret is positioned over the Erlenmeyer flask, the titrant is added in small amounts to the other solution until a chemical reaction occurs, generally identified through a color change.
- Once the reaction takes place, you can determine the volume of titrant needed to produce the reaction and calculate the concentration of solute in the other solution.

GRAVIMETRIC ANALYSIS

Gravimetric analysis is used to determine the amount of a certain substance in a sample by separating one component of the sample and weighing that component. The weight of the separated component is compared to the original weight of the sample.

- Precipitation is often used to separate a solid sample from the solution to be weighed.

Quick Fact

A type of gravimetric analysis can be used to determine the amount of water in a substance. If you weigh the substance, heat it until all the water evaporates, and then weigh the substance again, you can determine how much water it contained.

SECTION XVIII: APPLICATIONS OF CHEMISTRY IN EVERYDAY LIFE

OBJECTIVES

- Describe chemical processes that occur in your home, such as fermentation.
- Identify common chemicals found in your home.
- Distinguish between different food preservation methods.
- Identify common chemicals found in automobiles.
- Distinguish between gasoline and diesel fuel.

CHEMISTRY IN YOUR KITCHEN

Although you may not think of chemistry when you think of food, chemical reactions play an important role in what you eat. Chemical reactions are involved in making most food products, preparing food to eat, and preserving food so that it does not spoil.

Quick Fact

An entire field of chemistry is devoted to food. Food chemists:

- Develop and improve food and beverage products
- Analyze methods for cooking, storing, and packaging food
- Study the effects of processing on the appearance, taste, aroma, freshness, and vitamin content of food
- Modify foods to control nutrient and calorie content

FERMENTATION

Fermentation is an important food process in which microorganisms are grown on a sugar-based medium. During fermentation, sugar is converted to ethanol and CO_2:

$$C_6H_{12}O_6(aq) \xrightarrow{\text{yeast}} 2CO_2(g) + 2C_2H_5OH(aq)$$

- Fermentation reactions are catalyzed by a group of enzymes that come from yeast, a type of fungus. The yeast written above the arrow in the chemical equation signifies that yeast is the catalyst in the reaction.
- Fermentation processes include:
  - The bacterial fermentation of milk to make cheese and yogurt.
  - The conversion of sugar to alcohol to make beer and wine.
  - The use of yeast to convert carbohydrates to bread. When baking, fermentation produces bubbles of CO_2 that cause the dough to rise.

COMMON CHEMICALS IN YOUR KITCHEN

The table below lists the chemical names and formulas of some common kitchen ingredients:

Common Name	Chemical Name	Chemical Formula	Function/Use
Baking soda	Sodium bicarbonate	$NaHCO_3$	cooking/baking ingredient and odor reducer
Carbonated water (seltzer water)	Carbonic acid	H_2CO_3	soft drink ingredient
Table salt (rock salt)	Sodium chloride	$NaCl$	seasoning
Table sugar (cane sugar)	Sucrose	$C_{12}H_{22}O_{11}$	sweetener

SUGAR

Sugars are water-soluble carbohydrates (see section on Chemistry in the Human Body).

- Most sugars have the general molecular formula $C_n(H_2O)_m$, with n and m representing some number.
- Sugar molecules have many hydroxyl (–OH) groups and can hydrogen-bond with each other and with water.
- They are solids at room temperature and are very soluble in water.
- Sugars such as glucose, fructose, sucrose, and dextrose are commonly used as sweeteners in food and drinks.
- Several sugar substitutes have been developed synthetically, including saccharine, aspartame (NutraSweet®), and sucralose (Splenda®).
  - These artificial sugars have lower calories per gram than regular sugar.
  - Artificial sweeteners also provide more sweetness per gram than regular sugar. When using an artificial sweetener instead of sugar, you will use less sweetener. Therefore, the same volume of an artificially sweetened drink will generally end up being lighter than a sugar-sweetened drink.

Quick Fact

You may have heard people refer to vinegar as acetic acid (CH_3COOH). However, vinegar is not purely acetic acid. It is actually a very diluted solution of acetic acid. Household vinegar, used in cooking, generally only contains between 4% and 8% acetic acid.

Quick Fact

Sodium bicarbonate is a principal component of baking powders. It is used to replace yeast in baking. Baking soda will react with acidic components in the recipe (such as lemon juice, milk, honey, or molasses) to release CO_2. The release of CO_2 causes the cake, muffins, or other baked food to rise:

$$NaHCO_3 + H^+ \rightarrow Na^+ + H_2O + CO_2$$

Fehling's and Benedict's solutions are used to test for sugars.

- **Fehling's solution:** a deep-blue alkaline solution that forms a brick-red precipitate when heated with a substance containing a simple sugar (see section on Physical and Chemical Separations.)
- **Benedict's solution:** a similar blue alkaline solution that also forms a brick-red precipitate in the presence of simple sugars, such as glucose. Benedict's solution is often used in place of Fehling's solution today.

The simple sugar, glucose, is produced by plants through photosynthesis. Both plants and animals need glucose for energy. Glucose that is not needed right away is stored for later use.

Most *plants* store glucose in the form of either cellulose or starch.

Starch: a polymer, consisting of many glucose units (see section on Polymers).

- Starch is formed by plants to store excess glucose.
- Starch occurs in two forms:
  - Alpha-amylose, in which the glucose units are linked together in straight chains
  - Amylopectin, in which the chains are branched
- In the digestive tracts of mammals, starch chains are broken down into individual glucose units by the enzyme, *amylase*.

Quick Fact

Iodine can be used to test for starch. The iodine molecule reacts with the starch to form a compound with a dark blue or black color. If starch is not present, the color of the test solution will remain orange or yellow.

Cellulose and disaccharides, such as sucrose, will not change color when exposed to the test solution.

I

IODINE

Atomic #53

Iodine is a halogen that gets its name from the Greek word "iodes," meaning "violet."

Characteristics:

- Is a bluish-black, shiny solid
- Changes at ordinary room temperatures into a blue-violet gas with a strong odor
- Used as a disinfectant for cuts and wounds
- Is an essential element for human life that is needed for cellular metabolism and normal thyroid function
- Is often added to table salt (iodized salt) to provide the minimum iodine requirements for good health

Cellulose: a polymer, like alpha-amylose starch, that consists of linear chains of glucose molecules. However, the bonds between the glucose units have a different orientation.

- Mammals are unable to digest cellulose without the help of certain enzymes.
- Some mammals, such as cows or deer, are able to survive on cellulose. They cannot digest the cellulose themselves, but their digestive tracts contain microbes that produce the enzyme, *cellulase*. The cellulase works to break down the cellulose.

How is this digestion of cellulose different from the digestion of starch? Cellulose digestion occurs by fermentation. The end product is a fatty acid, not glucose. The process gives off carbon dioxide, not water.

Quick Fact

Just as most sugar names end in "-ose," the names of the enzymes that catalyze reactions of the sugars end in the letters "-ase."

Most *animals* store glucose in the form of glycogen. **Glycogen** is an organic polymer with a branched structure that resembles starch because it consists of chains of glucose units. Glycogen is formed and stored mainly in the liver and muscles.

PEPPER

Pepper is often used as a seasoning. It affects the flavor of food through chemical reactions that occur with the receptors on your tongue. The two kinds of peppers commonly used to season foods are black pepper and chili (or red) pepper.

Black pepper comes from the berry or "peppercorn" of the climbing plant Piper Nigrum.

- When the peppercorn berries first begin to ripen, they are picked and dried to make black pepper.
- Black pepper has a spicy flavor that comes from the chemical compound, piperine ($C_{17}H_{19}NO_3$).

Red pepper comes from the family of plants known as Capsicum.

- Red peppers include bell peppers, paprika, cayenne, tabasco, and habanero peppers.
- Red peppers derive their "heat" from chemicals known as capsaicins.
  - Capsaicins bind with pain sensors in the mouth, releasing pain transmitters that travel to the brain. The brain senses the pain as "heat" and even signals the body to increase perspiration (sweating).
  - Capsaicins are non-polar and are not soluble in water, so drinking water will not help cool your mouth off after hot pepper consumption. Your best bet is to drink tomato juice or lemon juice (something with some acidity to neutralize the alkaline capsaicin). Yogurt or milk, which contain fats that can help dissolve the capsaicin, may also be effective.

Quick Fact

The "heat" of peppers is measured in Scoville Heat Units (SHU).

- Bell peppers measure 0 SHU
- Jalapeños measure about 5,000 SHU
- Habaneros measure about 300,000 SHU.
- Pure capsaicin measures about 16,000,000 SHU!

The pepper spray carried by law enforcement officers is about 5,000,000 SHU.

CAFFEINE

Caffeine is naturally produced in the seeds, leaves, and fruits of many plants. Coffee beans, cacao beans, and tea leaves are well known for their caffeine content. Caffeine acts as a natural pesticide, paralyzing and killing insects that try to feed on the plant.

- When purified, caffeine is an intensely bitter white powder.
- In humans, caffeine stimulates the central nervous system, heart rate, and respiration process. It also acts as a mild diuretic (a substance the increases the rate of urine production).
- Caffeine is what gives coffee and tea their "wake up" boost and slightly bitter flavor. It is also added to colas and other soft drinks as a flavoring ingredient.

A relative of the caffeine molecule, called **theobromine**, is the key ingredient in chocolate.

- Theobromine comes from the beans of the cacao tree (or cocoa tree).
- Theobromine has similar effects on humans as caffeine does, but the effects of theobromine are weaker. Theobromine acts as a mild stimulant but has very little effect on the central nervous system.
- Different types of chocolate contain different amounts of theobromine.
  - "Dark" chocolates usually contain about 10 grams of theobromine in every kilogram of chocolate.
  - "Milk" chocolates contain 1 to 5 grams of theobromine in every kilogram of chocolate.
  - Pure cocoa beans are extremely rich in theobromine. They may contain 25 grams of theobromine in every kilogram of cocoa beans.

Quick Fact

Cocoa and chocolate products may be lethal to dogs and other domestic animals, such as horses, that metabolize theobromine more slowly than humans.

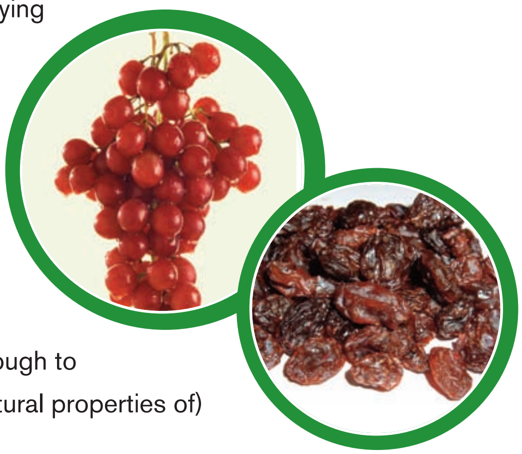
NOTES

PRESERVING FOOD

Most food will spoil if left open to the environment. Preservation methods help to prevent the growth of bacteria, fungi, and other microorganisms that might consume the food before you do. These microorganisms are what cause the bad smell or other unwanted byproducts.

Before refrigeration, ancient cultures found other ways to preserve their food. These methods included drying, salting, smoking, and pickling. More modern preservation methods include canning, freezing, freeze-drying, irradiation, and vacuum packing.

DRYING: a preservation method that removes water from food to prevent bacterial growth.



- Bacterial growth usually requires the presence of water, so drying is one simple way to preserve food.
- Drying has been used since ancient times to preserve:
  - Grains, such as wheat, rice, and oats.
  - Fruits, such as grapes (which makes raisins).
  - Most types of meats.
- Meats may be dried by heating them at a temperature low enough to drive off the water but not too hot to denature (change the natural properties of) the proteins in the meat.
  - Beef jerky is made by drying beef.

SMOKING: a preservation method in which meat or fish is dried in a smoke-house. Smoking dries the meat, while also flavoring it.

SALTING: a preservation method that uses salt to draw water out of the meat through a process called osmosis.

- **Osmosis:** the diffusion of certain molecules across semi-permeable cell membranes (walls).
- When a meat is salted, the concentration of the salt outside the meat is greater than the concentration of salt in the salt-water solution inside the meat. As a result, water moves from the inside of the cells, through the cell membranes. As the water moves through the cell walls to the outside, the meat dries.

FREEZE-DRYING: a special form of drying in which the food is frozen and then placed under a strong vacuum. The water in the food sublimes (going directly from a solid to a vapor) and is immediately removed by the vacuum.

- This method has less of an effect on a food's taste than regular drying does.
- Freeze-drying is most commonly used to make instant coffee.

PICKLING: a method of preserving fruits and vegetables by soaking them in brine (a salt solution) and then storing them in an acid solution, usually vinegar.

- The salt is used to draw out moisture from the vegetables by osmosis. Then, the vegetables are treated with vinegar.

CANNING: a preservation method that involves cooking fruits or vegetables and sealing them in sterile cans or jars. Once the fruits or vegetables are sealed in the cans or jars, the containers are boiled to kill any remaining bacteria.

PASTEURIZATION: the process of heating liquids to kill bacteria, molds, and yeast.

- Milk is pasteurized to kill any microorganisms in it by heating it to about 65°C for thirty minutes.

CARBONATION: a process by which carbon dioxide is dissolved in a beverage under pressure. Carbonation eliminates oxygen and prevents bacterial growth, thus beverages may be preserved by carbonation.

HISTORY: LOUIS PASTEUR

Louis Pasteur (1822-1895) is known for demonstrating that disease is caused by germs and for his development of techniques to establish resistance to disease, most notably the first vaccine against rabies.

Through his studies on fermentation, Pasteur invented a process in which liquids, such as milk, were heated to kill all bacteria and molds already present within them. This process has become known as "pasteurization."

CHEMISTRY IN YOUR AUTOMOBILE

Cars, trucks, buses, and other automobiles help us to get around every day. These automobiles need chemistry to get us where we need to go.

GASOLINE

Lately, you may have heard a lot of talk about gasoline, and you probably know why it is so important. It is what keeps our cars moving on the road!

Gasoline is a mixture of hydrocarbons used mainly as a fuel for automobiles. Gasoline molecules have 5-10 carbons in each hydrocarbon chain, which include heptane, octane, and decane (see section on Naming Organic Compounds).

- Gasoline is made from crude oil, which contains hydrocarbons (see section on Hydrocarbon Resources).
  - The carbon atoms in crude oil link together in chains of different lengths. These different chain lengths can be separated from each other and mixed together to make different fuels.
  - The chains with 5-10 carbons (generally molecule chains from C_7H_{16} through $C_{10}H_{24}$) are blended together to make gasoline.

Quick Fact

The hydrocarbon chain molecules that make up gasoline all evaporate at temperatures below the boiling point of water. If you spill gasoline outside, it will evaporate quickly.

The engine of a car is an internal combustion engine that compresses a cylinder full of air and gasoline. The mixture of compressed air and gasoline is then ignited (set on fire) by a spark from a spark plug.

Octane rating: a rating of how much the gasoline can be compressed before it suddenly ignites on its own.

- Gasoline that ignites from the compression only (instead of from the spark) can cause damage to the car's engine.
- Heptane ignites easily on its own when compressed. Octane can handle much more compression before it ignites.
- Typical "regular" gasoline is 87-octane.
  - This means that the gasoline contains 87-percent octane and 13-percent heptane (or another mixture of fuel that has the same performance of the 87-octane/13-heptane mixture).
  - Lower-octane gasoline, such as "regular" gasoline, ignites on its own with the least amount of compression.
  - "High-performance cars" with "high-performance engines" need higher-octane fuel that can handle a greater amount of compression.

Quick Fact

Before filling stations (gas stations) were built, Americans bought their gasoline in cans to fill their gas tanks.

DIESEL

Have you ever noticed that large tractor-trailers fill up their tanks at a different part of the station than the regular cars? That's because most tractor-trailers and other big trucks use diesel fuel.

Diesel or diesel fuel is also a mixture of hydrocarbons that come from crude oil.

- The hydrocarbon chains that make up diesel fuels contain between 14 and 20 carbons, including cetane ($C_{16}H_{34}$).

The fuel cylinder in a diesel fuel engine is ignited solely by compression instead of a spark. A diesel engine is an internal combustion engine in which the heat from the highly compressed air in the cylinder causes the fuel to ignite.

- Similar to the octane rating of gasoline, diesel fuels are given a cetane rating.
- **Cetane number (cetane rating):** a measure of the combustion (ignition) quality of diesel fuel based on the percentage of cetane in the fuel.
  - The cetane number actually measures the fuel's ignition delay. It measures the time period between when the engine is started and when the fuel is ignited.
    - Higher cetane fuels have a shorter ignition delay than lower cetane fuels.
    - "Regular" diesel fuel generally has a cetane number between 40 and 45. "Premium" diesel fuel generally has a cetane number between 45 and 50.
    - Therefore, good diesel fuels have a higher cetane number, which also means greater fuel efficiency.

Diesel fuel tends to be more fuel-efficient (burning less fuel in the same distance than gasoline). However, burning diesel fuel cleanly is more difficult.

- Diesel fuels typically contain higher quantities of sulfur, which can be harmful to the environment.
- In recent years, diesel vehicles have become much cleaner. A new type of diesel fuel, called ultra-low sulfur diesel (ULSD) is required for new diesel vehicles.

ANTIFREEZE

The engines of cars and trucks produce a lot of heat. Too much heat within the engine of a car can cause damage. Some of the heat must be removed to protect the car and prevent "overheating."

Antifreeze is a chemical that is mixed with water and added to the cooling system of a car.

- The most common chemical used in antifreeze is ethylene glycol ($C_2H_6O_2$).
- The antifreeze-water mixture moves through an automobile engine to remove extra heat.
- Antifreeze gets its name from the fact that it has a lower freezing point than water. The lower freezing point prevents the engine coolant from freezing in the cold temperatures.
- Antifreeze also has a higher boiling point than pure water, which makes it useful in hot temperatures as well.

BATTERY ACID

Not only do cars use gasoline, they also use a battery. A car battery is a type of rechargeable, lead-acid battery that supplies electric energy. A car battery gives power to the lights, radio, and the ignition system of an automobile.

The acid contained in the lead-acid car batteries is commonly called battery acid.

- **Battery acid:** diluted sulfuric acid used in storage batteries.
  - Battery acid is a mixture of about 33% sulfuric acid (H_2SO_4) and water.
  - Although battery acid is diluted sulfuric acid, it is still very corrosive. It can cause severe damage to human skin, so you should wear protective gloves when changing a car battery.

S

SULFUR

Atomic #16

Sulfur is found near hot springs and volcanoes.

Characteristics:

- Is a pale yellow, brittle solid
- Is an essential element for life because it is a component of hair, muscle tissues, and the hormone insulin

Most of the world's sulfur production is used to make sulfuric acid (H_2SO_4), which is used in lead-acid car batteries, fertilizers, the bleaching of paper, and many other industrial processes.

Sulfur dioxide (SO_2) is especially toxic to lower organisms, such as fungi, and for this reason, is used for sterilizing dried fruit and wine barrels.

Most sulfur compounds have foul odors, smelling like rotten eggs or skunk spray.

SECTION XIX:
INDUSTRIAL APPLICATIONS OF CHEMISTRY

OBJECTIVES

- Identify common alloys.
- Define monomers and polymers.
- Describe methods for synthesizing polymers.
- Identify different types of plastics and their uses.
- Describe biopolymers and identify biopolymers in the human body.
- Explain the process of making rubber and the use of sulfur in that process.

ALLOYS

Metals are often mixed together to enhance the properties of the resulting metal.

Alloys: homogeneous mixtures made of two or more metals or of a metal and nonmetal.

- Changing the makeup of an alloy may change properties such as density, hardness, conductivity, melting point, and malleability (see section on Bonding Basics).
- Alloys are also used to keep metals from corroding.
- Some common alloys include:
  - *Brass* (made of copper & zinc) has a gold color and does not corrode. The color of brass depends on the amount of zinc in the alloy – the more zinc, the lighter the color.
  - *Bronze* (made mainly of copper & tin) has good hardness and corrosion resistance. Bronze has been used for over 3,000 years to make statues, weapons, and other household materials. Today, bronze is often used to make machine parts, such as gears, because it creates very little friction.
  - *Steel* (made of iron & carbon) is a strong metal that is often used in construction. *Stainless steels* contain chromium and/or nickel and are resistant to rusting.
  - *White gold* (generally made of gold and nickel or palladium) became popular in the 1920s to make “white” jewelry. It was used as a substitute for platinum, which was much more expensive.

POLYMERS

What do water bottles, tires, contact lenses, and DNA have in common? They are all types of polymers.

Polymers: long, chain-like molecules that are formed by connecting many repeating units (monomer units). Common polymers include:

- Polyester (used for wrinkle-free fabric)
- Polyurethane (used to make flexible foam for bedding and upholstery and rigid foam for wall panels and refrigerators)
- Polymethylmethacrylate (used in contact lenses)

Quick Fact

The word "polymer" comes from the two Latin roots, "poly" which means "many," and "mer" which means "pieces." Literally, polymer translates into "many pieces."

Similarly, "monomer" means "one piece."

Monomer: a single molecule capable of combining with other similar molecules.

- One common type of monomer is an alkene or olefin molecule, which contains a carbon-carbon double bond.

EXAMPLE:

Styrene monomer is a simple liquid. When it polymerizes, it becomes the solid that is used in Styrofoam™ cups or restaurant take-out boxes.

SYNTHESIZING POLYMERS

The two main methods of synthesizing a polymer (combining monomers to create a polymer) are addition polymerization and condensation polymerization.

Addition polymerization (chain polymerization): bonding of monomers to form giant molecules without byproducts.

- Monomers are added one-by-one to the active site on the growing polymer chain.
- Polyethylene and polypropylene are addition polymers used to make various consumer products, including trash bags and plastic containers.
- Polystyrene, as mentioned above, is another classic example of an addition polymer.

Think About It...

What were military helmets made of before the advanced Kevlar design? What design properties does an engineer seek in a helmet, apart from being bullet-resistant?

Condensation polymerization (step polymerization): combining of monomers to produce some additional byproduct, such as water.

- A common condensation polymer is nylon-66, the most widely used of all synthetic polymers.
- Kevlar is an important condensation polymer used for bullet-proof vests and protective helmets.

PLASTICS

Plastics are organic compounds produced by polymerization that can be molded, cast into shapes, or drawn into fibers. The two basic types of plastics are thermosetting plastics (thermosets) and thermoplastics.

Thermosets: plastics that retain their shape after being formed through heat and pressure. Thermosets cannot be remolded.

- Strong cross-linked bonds are formed during the initial molding process, giving the material a stable structure.
- Polyesters are a type of thermosetting plastic that can be combined with fiberglass to produce a glass reinforced plastic (GRP) used in car bodies, sail boats, and furniture.

Quick Fact

One of the most interesting properties of polymers is their toughness. For example, if you've ever dropped your telephone or cell phone, it most likely did not break.

Many adults will cringe when they drop a plastic glass, most likely because when they grew up, glasses were made of glass! If you dropped it, it would break.

Milk jugs used to be made of glass too. Can you imagine dropping your milk jug out of the refrigerator? The old glass containers would shatter, but the new plastic ones will generally survive the fall.

Thermoplastics: plastics that soften when heated and harden again when cooled. Because thermoplastics can be remolded, they are fully recyclable.

- Acrylics are a type of thermoplastic available in a variety of colors and forms. Acrylics withstand weather and are stable in sunlight. Transparent acrylic can be as clear as fine optical glass and is used in some optical equipment, such as cameras.

Some plastics come from materials of biological origin, such as starch and lactic acid.

- **Biodegradable plastics:** plastics that can be broken down by bacterial enzymes.
- Biodegradable plastics have become increasingly popular as people seek to make convenient ways of life fit with "environmentally friendly" lifestyles.

Think About It...

Some would say that plastics are *too* durable. How can you dispose of plastic milk cartons or containers in an environmentally friendly way?

BIOPOLYMERS

Nature relies on polymers too. The larger size of polymeric molecules brings about strength and more advanced functions.

Biopolymer: a group of polymers found in or produced by living organisms.

- Biopolymers include starch, sugar, and cellulose.
- DNA and RNA are examples of biopolymers found inside the body. Other important biopolymers in the human body are proteins.

RUBBER

Natural rubber is a stretchy polymer made from the milky substance (latex) found in certain tropical trees. However, today, many rubber products are made from synthetic rubbers.

In 1839, Charles Goodyear accidentally dropped a mixture of natural rubber and sulfur onto a hot stove and discovered that the rubber became stronger and more elastic. Goodyear called this process *vulcanizatio*n after the Roman god of fire and the volcano.

- This process results from **cross-linking** of the polymer, where the sulfur forms disulfide bridges between strands of the polymer. This prevents strands of the polymer from slipping around too much.
- The properties of rubber can be adjusted by controlling the amount of sulfur used in vulcanization.
  - Rubber made with 1-3% sulfur is soft and stretchy and is used in rubber bands.
  - With 3-10% sulfur, the rubber hardens enough to make good tires.
  - With 20-30% sulfur, the rubber becomes a hard synthetic plastic.
- Goodyear's process made rubber waterproof and winter-proof and opened the door for an enormous rubber goods market.

Quick Fact

The term "elastomer" is often used interchangeably with rubber and means "stretchy pieces."

HISTORY: **SERENDIPITY – LUCK IN SCIENCE**

Just like Charles Goodyear's accidental discovery of the vulcanization process, other scientists have made fortunate accidental discoveries while looking for something else.

- The polymer, Teflon (used to coat non-stick pots and pans), was discovered accidentally by scientists at DuPont who were developing new chlorofluorocarbon (CFC) refrigerants. When they couldn't get one of their reactants to come out of its pressurized tank, they cut open the tank and discovered the gas had polymerized to form a solid that was almost completely unreactive.
- The Dow Chemical Company discovered Saran© wrap by accident. Dow was researching the textile strengths of plastic materials and accidentally synthesized one that tore very easily. Saran© wrap became widely used for keeping food products from spoiling. The easy-to-tear property actually turned out to be a plus.

NOW YOU'RE READY FOR THE

CHALLENGE

GOOD LUCK!

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