

CHM 151: CHEMISTRY AROUND US



Erin Avram
Cleveland State University

Chemistry Around Us

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1: Creating Accounts

Creating Libretext ADAPT Account

You will use the ADAPT system of LibreText to submit your assignments and you will use the Single Sign On option through your school (@vikes.csuohio.edu) email. **Do not create a private account within the ADAPT system.**

1. Go to <https://adapt.libretexts.org>. (Figure 1.3) Click **Log IN** (DO NOT REGISTER!)

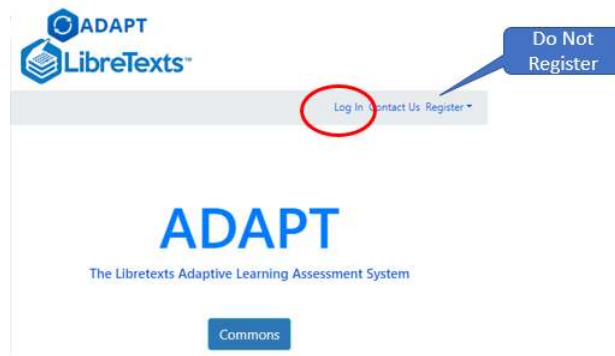


Figure 1.3: Click Log In. Do not Register.

2. Click Campus Login (DO NOT FILL OUT FORM) (figure 1.4).

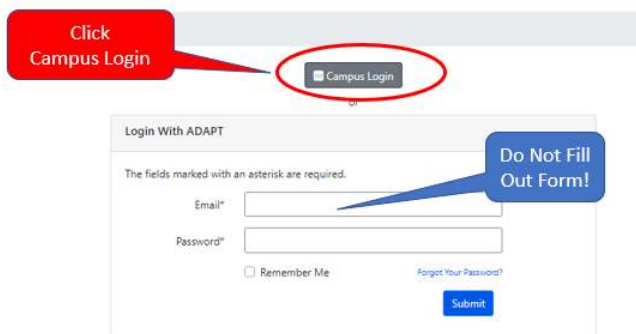


Figure 1.4: Click "Campus Login", this will allow you to set up a Single Sign On (SSO) with your UALR email account.

3. Click "Microsoft Active Directory" (figure 1.5).

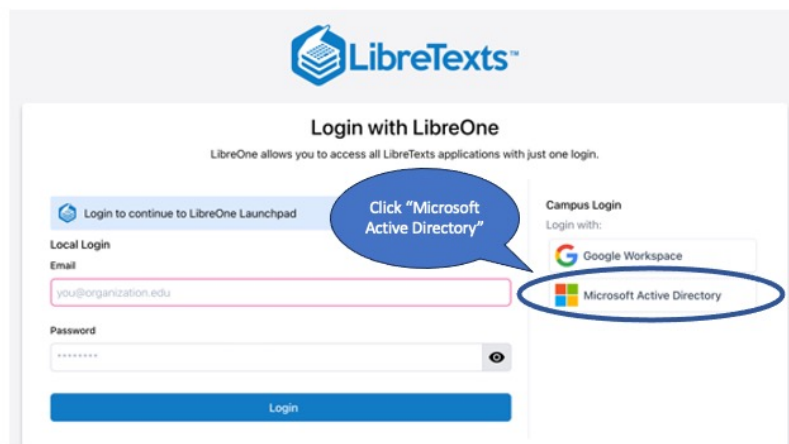


Figure 1.5: You will be able to log in with your CSU email through the Microsoft Active Directory option)

4. Choose your CSU email account and fill out form, choosing the student option (figure 1.6).

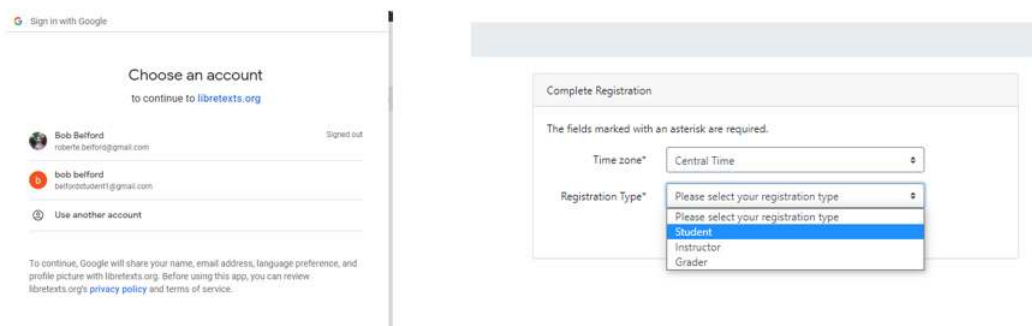


Figure 1.6: Be sure to set up the account with your UALR email address.

5. You now have an account and can enroll in a course.

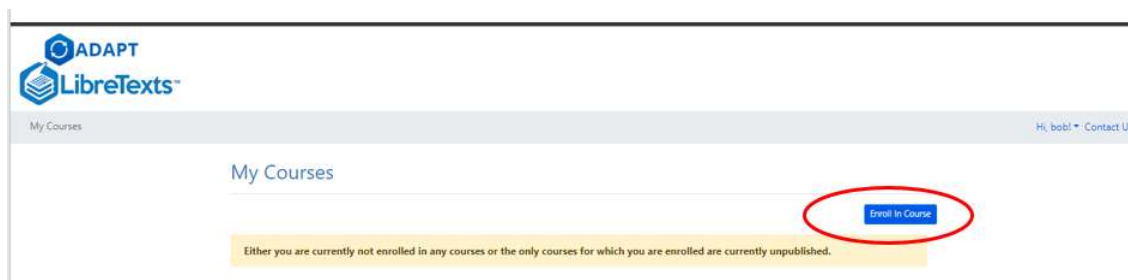


Figure 1.7: You now have an account within the ADAPT system and need to get an access code from your instructor to enroll in a course (Copyright; author via source)

6. You will need an access code for each course you take. This is provided by your instructor and can be found in Blackboard.

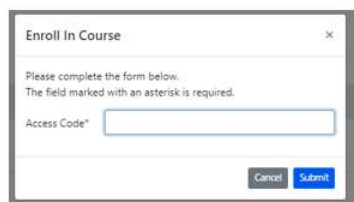


Figure 1.8: Insert your access code and you will become registered for a course.

7. You now can navigate to your course assignments

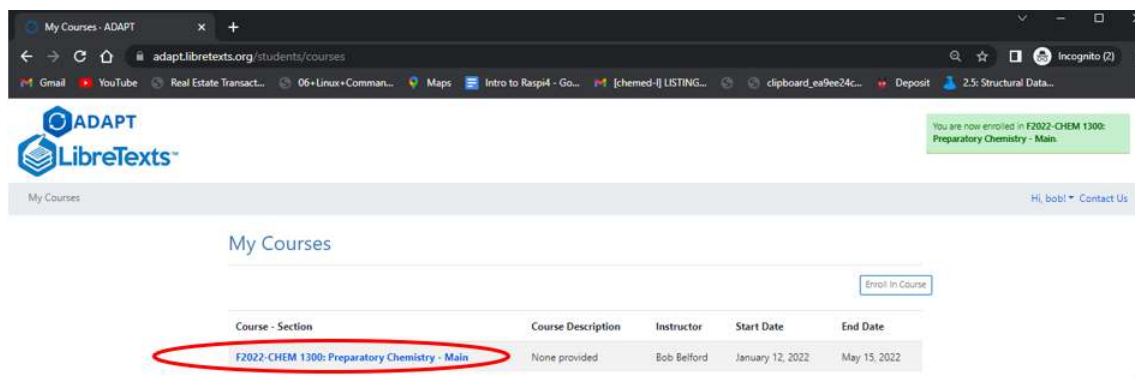


Figure 1.9: Copy and Paste Caption here. (Copyright; author via source)

Setting up Notifications

ADAPT can send an alert to your email when an assignment is due. This is advised as you will lost points for late assignments.

1. Next, Click on your name in the top right corner and go to Settings.

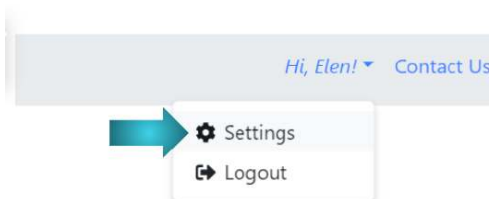


Figure 1.10: Open settings

2. Go to Notifications and choose the time option for your reminders. Click "Update"

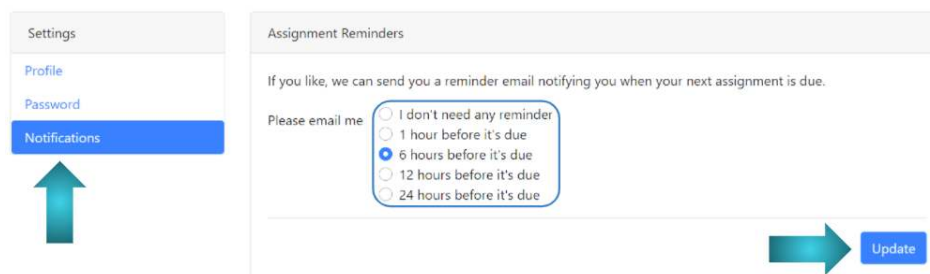


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A detailed breakdown of this resource's licensing can be found in [Back Matter/Detailed Licensing](#).

CHAPTER OVERVIEW

1: Properties and Measurement of Matter

- 1.1: Classification and Properties of Matter
- 1.2: All About the Elements
- 1.3: Measuring Matter
- 1.4: Unit Conversions
- 1.5: Density
- 1.6: Temperature

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1.1: Classification and Properties of Matter

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Learning Objectives

- Use physical and chemical properties to describe matter.
- Distinguish between physical and chemical changes.
- Describe the properties of each state of matter.
- Identify a sample of matter as an element, a compound, or a mixture (homogeneous/heterogeneous).
- Distinguish homogenous and heterogenous mixtures.

1.1.1 Physical and Chemical Properties

The properties that chemists use to describe matter fall into two general categories. Physical properties are characteristics that **describe** matter. They include characteristics such as size, shape, color, and mass. Many of these properties can be quantitative in nature. For example, quantitative physical properties of water would be the boiling point (100 °C / 212 °F) and melting point (0°C / 32 °F).



Figure 1.1.1: Physical Properties. (Copyright;<https://www.slideshare.net/cfoltz/ph...cal-properties>)

Chemical properties are characteristics that describe how matter **changes** its chemical structure or composition. An example of a chemical property is flammability—a material's ability to burn—because burning (also known as combustion) changes the chemical composition of a material. Oxidation, rusting, decomposition, and inertness are chemical properties as well. Click on this [video](#) and record the physical and chemical properties of the element sodium.



Figure 1.1.2: Port Royal, South Carolina oxidation of iron. Image courtesy of Elizabeth R. Gordon.

1.1.2 Elements and Compounds

Any sample of matter that has the same physical and chemical properties throughout the sample is called a substance. There are two types of substances. A substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A substance that consists of two or more atoms is called a **molecule**. Oxygen gas (O₂) is

an example of a molecule because it contains two atoms of oxygen bonded together. A substance that can be broken down into chemically simpler components (because it has more than one element) is a **compound** (Figure 1.1.1). Water (H_2O) is a compound composed of the elements hydrogen and oxygen. Today, there are 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

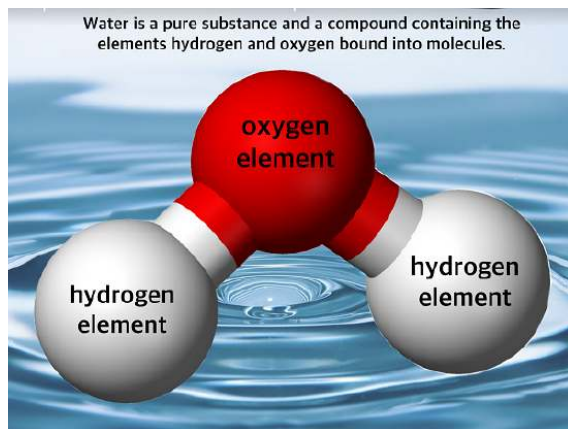


Figure 1.1.3: To a chemist, the pure substance water is a compound (it contains two kinds of atoms bound to one another in just one kind of molecule). An environmentalist might consider water "pure" even if it contains the normal amount of dissolved oxygen and carbon dioxide, but no other "pollutants". To a chemist, water containing oxygen is no longer a pure substance, but a mixture. In pure water, the ratio of hydrogen to oxygen atoms is always 2:1.

Sometimes the word *pure* is added to a substance, but this is not absolutely necessary. By definition, any single substance is pure.

The smallest part of an element that maintains the identity of that element is called an atom. Atoms are extremely tiny; to make a line 1 inch long, you would need 217 million iron atoms. The smallest part of a compound that maintains the identity of that compound is called a molecule. Molecules are composed of atoms that are attached together and behave as a unit. Scientists usually work with millions and millions of atoms and molecules at a time. When a scientist is working with large numbers of atoms or molecules at a time, the scientist is studying the macroscopic view of the universe.

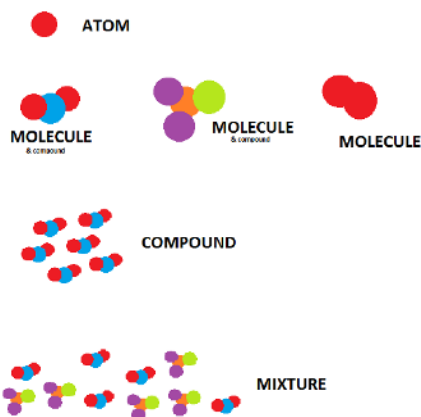


Figure 1.1.4: Appreciating elements, molecules and compounds.

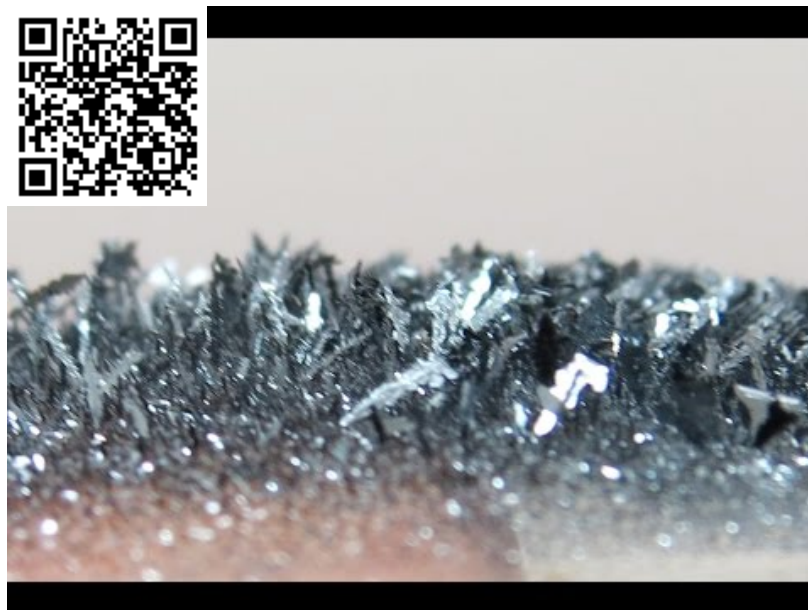
1.1.3 Mixtures

A material composed of two or more substances is a mixture. In a mixture, the individual substances maintain their chemical identities. Many mixtures are obvious combinations of two or more substances, such as a mixture of sand and water. Such mixtures are called **heterogeneous mixtures**. In some mixtures, the components are so intimately combined that they act like a single substance (even though they are not). Mixtures with a consistent composition throughout are called **homogeneous mixtures** (or

solutions). Sugar dissolved in water is an example of a solution. A metal alloy, such as steel, is an example of a solid solution. Air, a mixture of mainly nitrogen and oxygen, is a gaseous solution.

1.1.4 Phases

Another way to classify matter is to describe it as a solid, a liquid, or a gas, which was done in the examples of solutions. These three descriptions, each implying that the matter has certain physical properties, represent the three phases of matter. A solid has a definite shape and a definite volume. Liquids ordinarily have a definite volume but not a definite shape; they take the shape of their containers. Gases have neither a definite shape nor a definite volume, and they expand to fill their containers. We encounter matter in each phase every day; in fact, we regularly encounter water in all three phases: ice (solid), water (liquid), and steam (gas).



Video 1.1.35: Sublimation and deposition of Iodine.

We know from our experience with water that substances can change from one phase to another if the conditions are right. Typically, varying the temperature of a substance (and, less commonly, the pressure exerted on it) can cause a phase change, a physical process in which a substance goes from one phase to another (Figure 1.1.3). Phase changes have particular names depending on what phases are involved, as summarized in Table 1.1.1.

Table 1.1.1: Phase Changes

Change	Name
solid to liquid	melting, fusion
solid to gas	sublimation
liquid to gas	boiling, evaporation
liquid to solid	solidification, freezing
gas to liquid	condensation
gas to solid	deposition

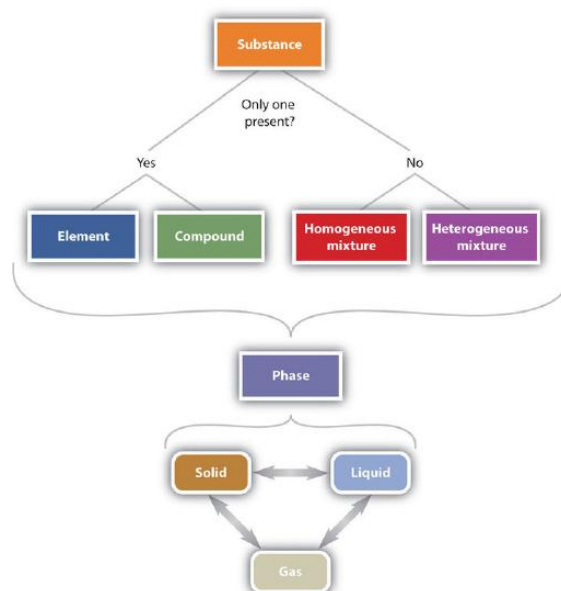


Figure 1.1.6: The Classification of Matter. Matter can be classified in a variety of ways, depending on its properties. (Cc BY-NC-SA; anonymous)

Contributors and Attributions

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- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)
- Erin Avram (Cleveland State University)

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1.2: All About the Elements

Learning Objectives

- Explain how elements are arranged on the periodic table.
- Identify groups and periods on the periodic table.
- Provide the name of an element given its abbreviation.
- Determine the identity and number of atoms in a compound.

Basic Arrangement of Periodic Table

The periodic table has columns (known as **groups**) and horizontal rows are called **periods** and organizes all 118 elements currently identified. Each element is assigned a symbol that is either one or two letters. The first letter of the element symbol will always be capitalized, if there is a second letter, it will always be lowercase. Previously, some of the newly discovered elements (located in period seven on other tables) used three-letter symbols. These three-letter symbols corresponded to a Latin numbering system until the elements were given official names.

Languages of Element Names

Interestingly, different languages are used on the periodic table. Depending on your background, you might be more familiar with some elemental symbols than others. For example, hydrogen is represented by the symbol H. This element's name is derived from the Greek hydro, and *genes* which means to generate water. Other elemental symbols like copper (which has a symbol of Cu) are not so predictable. This particular element's name originated from the Latin *cuprum* (from the island of Cyprus). A variety of languages are displayed on the periodic table. For example, German roots are present in the element tungsten. The symbol for this element is W which corresponds to the German word, wolfram. For this course, you do not need to have these memorized but should be able to use the Periodic Table to convert between the name of the element and the element symbol.

Meanings of Element Names

Many elements are named in honor of people. Some familiar names you might see are curium (Marie and Pierre Curie), einsteinium (Albert Einstein), and mendelevium (Dimitri Mendeleev). IUPAC (International Union of Pure and Applied Chemistry) requires these types of names to be done posthumously. In other words, the scientist who discovered the element could not name it after her/himself or anyone who is still currently living. Over the years, there have been two exceptions. The elements of seaborgium (Sg) and oganesson (Og) were named while Glen Seaborg and Yuri Oganessian were still living. Seaborgium was named in honor of an American nuclear chemist who synthesized the element plutonium. He is also credited with assisting the discovery of other elements like americium (element 95), curium (element 96), and fermium (element 100). His accomplishments encouraged his fellow researcher, Albert Ghiorso, to name Seaborgium in honor of him. As for element 118 (oganessian), a joint research commission of Russian and American scientists named this element for Russian physicist Yuri Oganessian. As of 2017, Oganessian is still living and is responsible for the discovery of superheavy elements.

It is acceptable to name an element in honor of a geographical location. Some notable Earth locations would include Am (americium, American continent), Ge (germanium, Germany), Fr (francium, France), and Eu (europium, European continent). Some scientists have ventured past Earth to name their elements in honor of other places. Examples of these would include Pu (plutonium, Pluto), U (uranium, Uranus), and Np (neptunium, Neptune).

PERIODIC TABLE: ELEMENT NAME ORIGINS

Symbol Key

Table Colour Key

Language of Origin

Name Source



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Interpreting Chemical Formulas

Chemical formulas are used as a shorthand to represent the components of a compound. Each capital letter of a chemical formula indicates a different element. Subscripts and parentheses are used to indicate the number of atoms of each element that is present. If there is no number listed, it is assumed that there is only a single atom of that element. To represent a compound with 1 atom of magnesium and 2 atoms of chlorine the formula would be MgCl_2 ; no subscript is needed after magnesium because only one atom is present, the subscript 2 after the symbol for chlorine indicates that two atoms of chlorine are present. For further information, see the example below.

✓ Example 1.2.1

Using the chemical formula below, identify the number of atoms for each element present.



Solution

The chemical formula shown contains three capital letters, one for each element it contains: sodium (Na), phosphorous (P), and oxygen (O). The subscripts after an element symbol indicate the number of atoms for that element. In this example, it means there are 3 atoms of sodium (Na). Parentheses are used to indicate that a portion of the formula is repeated. The number of

repeats is indicated by the subscript after the parentheses. In this example, the subscript 2 indicates there are two sets of PO_4 , and therefore 2 atoms of phosphorous (1×2) and 8 atoms of oxygen (4×2).

- 3 atoms of sodium
- 2 atoms of phosphorous
- 8 atoms of oxygen

Contributors and Attributions

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- Erin Avram (Cleveland State University)

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1.3: Measuring Matter

Learning Objectives

- Connect basic glassware and equipment with metric units.
- Provide SI units for mass, temperature, and length.
- Convert between standard and scientific notation.

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A variety of instruments are available for making direct measurements of the macroscopic properties of a chemical substance. For example, we usually measure the **volume** of a liquid sample with pipettes, burets, graduated cylinders, and volumetric flasks (Figure 1.3.1), whereas we usually measure the mass of a solid or liquid substance with a balance (Figure 1.3.2).

1.3.1 Instruments that Measure Volume

Many academic scientists report their volume measurements with the milliliter unit. On the other hand, analytical chemists tend to work with much smaller volumes, and they would use nanoliter and microliter samples in their laboratory. When considering the industrial realm, scientists work with larger volumes and would thus employ the liter unit of measurement.

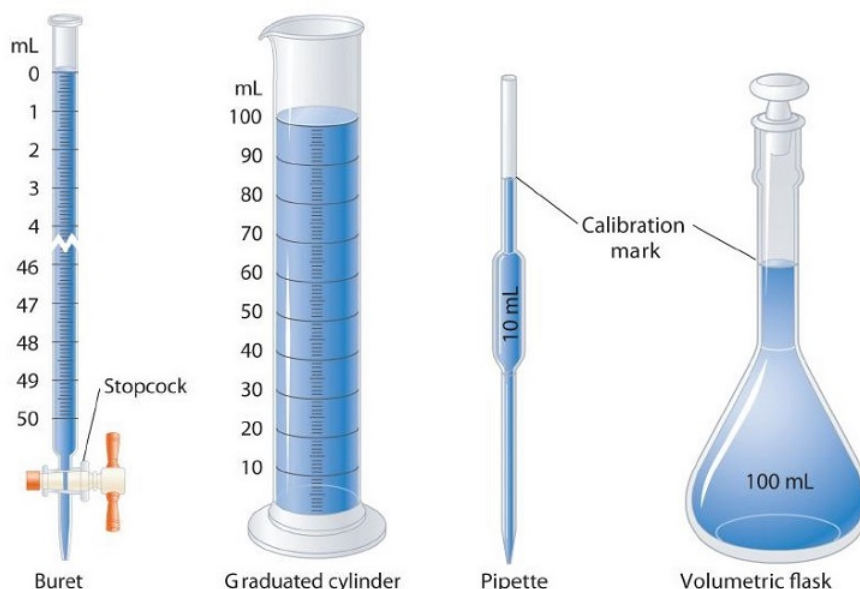


Figure 1.3.1: (left two) Graduated glassware is used to deliver variable volumes of liquid. (right two) Volumetric glassware is used to deliver (pipette) or contain (volumetric flask) a single volume accurately when filled to the mark. (CC BY-SA-NC; anonymous).

1.3.2 Instruments that Measure Mass

Balances are used to measure the **mass** of matter. Typically, high school and middle schools will use triple beam balances for laboratory activities. Colleges will rely on digital balances. For larger quantities of materials, top-loading digital balances are employed. For higher sensitivity and accuracy, some chemists use analytical balances.



Figure 1.3.2: Triple beam balance (image from: upload.wikimedia.org/wikiped...m_balance2.JPG. Digital top-loading balance (upload.wikimedia.org/wikiped...00-2M_-_01.jpg) and an analytical balance (upload.wikimedia.org/wikiped...-NaCl-1mol.jpg)

In the academic laboratory (like Cleveland State University's), a chemist will commonly use the gram, or the even smaller unit, the milligram. Industrial chemists making larger quantities of material would work with kilograms of materials. Analytical chemists (industrial or academic), environmental scientists, and toxicologists are typically more concerned with the smaller units of the metric system. For example, many chemicals found in drinking water and air are regulated in concentrations like nanograms or micrograms.

1.3.3 Tools that Measure Length

Length is measured using a metric stick or a metric ruler marked in millimeters and centimeters. Most devices used to measure length (in the United States) contain both English and metric markings.

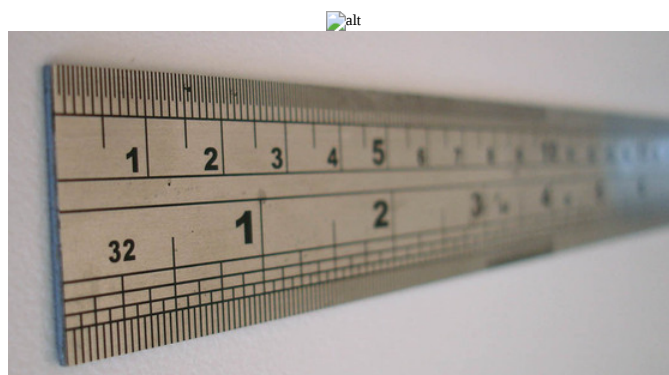


Figure 1.3.3: Metric and English measuring devices. (CC BY-SA 4.0; Ejay).

1.3.3.1 SI Units

All reported measurements must include an appropriate unit of measurement. To say that a substance has “a mass of 10” does not tell whether the mass was measured in grams, pounds, tons, or some other unit. For example, a nurse giving morphine to a patient needs to be specific in the unit dosages. A volume of liters would kill a patient quickly while a milliliter amount would provide sufficient pain relief. To establish worldwide standards for the consistent measurement of significant physical and chemical properties, an international body called the General Conference on Weights and Measures devised the *Système Internationale d'unités* (or SI). The International System of Units is based on metric units and requires that measurements be expressed in decimal form. Table 1.3.1 lists the seven base units of the SI system; all other SI units of measurement are derived from them.

Table 1.3.1: *SI Units*

Base Quantity	Unit Name	Abbreviation
mass	kilogram	kg
length	meter	m
time	second	s
temperature	Kelvin	K
amount of substance	mole	mol
intensity of light	candela	cd

1.3.3.2 The Metric System

The metric system consists of base units such as meter, liter, and gram. Sometimes, measurements can be very large or very small. To keep numbers manageable and easily readable, metric prefixes can be used. These prefixes are used to express values based on the power of 10 and serve as a way to limit the number of zeros needed to report a number. Rather than reporting 0.000000056 g, 56 ng can be used instead. Table 1.3.2 lists the common metric prefixes.

Table 1.3.2: Prefixes Used with metric bases of meter, liter, and gram

Prefix	Symbol	Value	Power of 10	Meaning
tera	T	1,000,000,000,000	10^{12}	trillion
giga	G	1,000,000,000	10^9	billion
mega	M	1,000,000	10^6	million
kilo	k	1000	10^3	thousand
hecto	h	100	10^2	hundred
deca	da	10	10^1	ten
—	—	1 (this is base unit which could be gram, liter, or meter)	10^0	one
deci	d	0.1	10^{-1}	tenth
centi	c	0.01	10^{-2}	hundredth
milli	m	0.001	10^{-3}	thousandth
micro	μ (medical community uses mcg)	0.000001	10^{-6}	millionth
nano	n	0.000000001	10^{-9}	billionth
pico	p	0.000000000001	10^{-12}	trillionth
femto	f	0.000000000000001	10^{-15}	quadrillionth



Figure 1.3.4: Pharmaceuticals use a different notation for the microgram unit than traditional chemists. (CC BY-SA 3.0; Cronnie).

Two methods of Memorization for the Metric System

Using negative exponents (see above) is the traditional way to memorize the metric system. Many students who have been taught outside of the United States approach the metric system this way. For their conversion factors, students who have been taught the metric system in the United States tend to use positive exponents. Either method is fine, pick one method and use it consistently.

Two different ways to memorize the metric systems (be sure to know all bolded conversion factors listed in section 2.1 of this OER)

- 1 gigabase = 1×10^9 base
- 1 megabase = 1×10^6 base
- 1 kilobase = 1×10^3 base
- 1 decibase = 1×10^{-1} base or 1×10^1 decibase = 1 base
- 1 centibase = 1×10^{-2} base or 1×10^2 centibase = 1 base
- 1 millibase = 1×10^{-3} base or 1×10^3 millibase = 1 base
- 1 microbase = 1×10^{-6} base or 1×10^6 microbase = 1 base
- 1 nanobase = 1×10^{-9} base or 1×10^9 nanobase = 1 base
- 1 picobase = 1×10^{-12} base or 1×10^{12} picobase = 1 base

Note: A base unit can be gram, liter, or meter

1.3.4 Additional Units

When traveling to different regions of the globe, it is helpful to be able to convert from English to Metric units.

Unit Conversion
2.21 pounds = 1.00 kilograms
1 mile = 1.61 kilometers
1.00 inch = 2.54 centimeters
1.00 mL = 1.00 cc (cubic centimeters) = 1.00 cm³
1 teaspoon (tsp) = 5.00 mL
12 in = 1.00 ft.
16 oz. = 1.00 pounds

1.3.5 Scientific Notation

Chemists often work with numbers that are exceedingly large or small. For example, entering the mass in grams of a hydrogen atom into a calculator requires a display with at least 24 decimal places. The tedium and awkwardness of manipulating numbers with large or small magnitudes can be avoided by the use of metric prefixes, as previously discussed, or with scientific notation. In scientific notation, these numbers are expressed in the form

$$N \times 10^n$$

where N is greater than or equal to 1 and less than 10 ($1 \leq N < 10$), and n is a positive or negative integer ($10^0 = 1$). The number 10 is called the base because it is this number that is raised to the power n . Although a base number may have values other than 10, the base number in scientific notation is always 10.

A simple way to convert numbers to scientific notation is to move the decimal point as many places to the left or right as needed to give a number between 1 to 10 (N). The magnitude of n is then determined as follows:

- If the decimal point is moved to the left n places, n is positive.
- If the decimal point is moved to the right n places, n is negative.

Another way to remember this is to recognize that as the number N decreases in magnitude, the exponent increases, and vice versa. The application of this rule is illustrated in Example 1.3.1.

✓ Example 1.3.1: scientific notation

Convert each number to scientific notation.

- 637.8
- 0.0479
- 7.86
- 12,378
- 0.00032
- 61.06700
- 2002.080
- 0.01020

Solution

Explanation	Solution
To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 6.378 Because the decimal point was moved two places to the left, $n = 2$.	$637.8 = 6.378 \times 10^2$
To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 4.79 Because the decimal point was moved two places to the right, $n = -2$.	$0.0479 = 4.79 \times 10^{-2}$
This is usually expressed simply as 7.86. (Recall that $10^0 = 1$).	7.86×10^0
Move the decimal point four places to the left and note a positive exponent.	1.2378×10^4
Move the decimal place four places to the right a note a negative exponent.	3.2×10^{-4}
Move the decimal place one space to the left and note a positive exponent.	6.106700×10^1
Move the decimal place three places to the left and note a positive exponent.	2.002080×10^3
Move the decimal place two places to the right and note a negative exponent.	1.020×10^{-2}

1.3.6 Using the Scientific Calculator

Please be sure to obtain a calculator that has an exponent button on it. It will either say "exp" or have an "ee" key indicating this function. This key represents the function of 10^x . You do not need to enter a 10 for an exponent when using this particular function. Read the steps below and practice on your calculator.

Use of Calculator

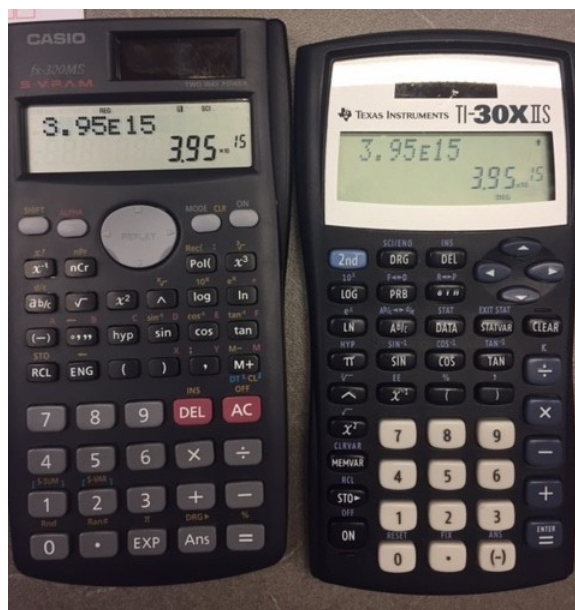
To input powers of ten:

1. Locate the ee or exp button on your calculator. Do not use the carrot (looks like the $\wedge$ button).
2. Put the coefficient number in and then hit "ee" or "exp."
3. If you need to have a negative "exp," then locate the +/- button.
4. Lastly, input your power of ten. You do not need to put the actual number ten in for the exponent.

Table 1.3.3: Examples

Number	How to Input in a calculator
3.95×10^{15}	Put 3.95 in as the coefficient number. Hit the "ee" or "exp" button. Lastly, enter the 15 for the exponent. Your calculator should say $3.95e^{15}$.
9.26×10^{-4}	Put 9.26 in as the coefficient number. Hit the "ee" or "exp" button. Locate the small +/- button for negative (not minus function). Next, enter the number four. Your calculator should say $9.26e^{-4}$.

This is an image of suggested calculators for CHM 151. You will need a non-programmable, scientific calculator that does not have metric conversion functions. In this picture, you will see the correct way to input one of the numbers shown in the example above. Please take time to practice with your calculator before tests and quizzes.



1.3.7 Contributors and Attributions

- [Elizabeth R. Gordon \(Furman University\)](#)
- [Hayden Cox \(Furman University\)](#)
- [Erin Avram \(Cleveland State University\)](#)

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1.4: Unit Conversions

Learning Objective

- Convert metric and English-based units.
- Use dimensional analysis (factor-label method) to solve quantitative problems
- Master using your calculator.

Conversion Factors

Recall your metric table (provided on the exam cover sheet) while working on these types of problems.

Two methods of Using Metric Prefixes

Using negative exponents is one way to use the metric prefixes. Many students who have been taught outside of the United States approach the metric system this way. For their conversion factors, students who have been taught the metric system in the United States tend to use positive exponents. Either method is fine, just pick one way and use it consistently.

Two different ways to approach metric prefix conversions

- **1 gigabase = 1×10^9 base**
- **1 megabase = 1×10^6 base**
- **1 kilobase = 1×10^3 base**
- **1 decibase = 1×10^{-1} base or 1×10^1 decibase = 1 base**
- **1 centibase = 1×10^{-2} base or 1×10^2 centibase = 1 base**
- **1 millibase = 1×10^{-3} base or 1×10^3 millibase = 1 base**
- **1 microbase = 1×10^{-6} base or 1×10^6 microbase = 1 base**
- **1 nanobase = 1×10^{-9} base or 1×10^9 nanobase = 1 base**
- **1 picobase = 1×10^{-12} base or 1×10^{12} picobase = 1 base**

Note: A base unit can be gram, liter, or meter

We have an equivalence, 1 yard = 3 feet. This can also be written:

$$\frac{3 \text{ ft}}{1 \text{ yd}} = 1 \quad (1.4.1)$$

Equation 1.4.1 is a strange way to write 1, but it makes sense: 3 ft equal 1 yd, so the quantities in the numerator and denominator are the same quantity, just expressed with different units. The expression in Equation 1.4.1 is called a conversion factor and it is used to formally change the unit of a quantity into another unit. (The process of converting units in such a formal fashion is sometimes called *dimensional analysis* or the *factor label method*.)

To see how this happens, let us start with the original quantity: **4 yd** and let us multiply this quantity by 1. Rather than multiplying by just 1, let us write 1 as

$$1 = \frac{3 \text{ ft}}{1 \text{ yd}}$$

When you multiply anything by 1, you don't change the value of the quantity.

$$\begin{aligned} 4 \text{ yd} &= 4 \text{ yd} \times 1 \\ &= 4 \text{ yd} \times \frac{3 \text{ ft}}{1 \text{ yd}} \end{aligned}$$

The 4 yd term can be thought of as yd/1; that is, it can be thought of as a fraction with 1 in the denominator. We are essentially multiplying fractions. If the same thing appears in the numerator and denominator of a fraction, they cancel. In this case, what cancels is the unit yard:

$$4 \cancel{yd} \times \frac{3 ft}{1 \cancel{yd}}$$

That is all that we can cancel. Now, multiply and divide all the numbers to get the final answer:

$$\frac{4 \times 3 ft}{1} = \frac{12 ft}{1} = 12 ft$$

Again, we get an answer of 12 ft, just as we did originally. But in this case, we used a more formal procedure that is applicable to a variety of problems.

How many millimeters are in 14.66 m? To answer this, we need to construct a conversion factor between millimeters and meters and apply it correctly to the original quantity. One method would entail using the conversion factor of $1 mm = 1 \times 10^{-3} m$

$$14.66 \cancel{m} \times \frac{1 mm}{10^{-3} \cancel{m}} = 14660 mm \quad (1.4.2)$$

$$= 1.46 \times 10^4 mm \quad (1.4.3)$$

Which conversion factor do we use? The answer is based on *what unit you want to get rid of in your initial quantity*. The original unit of our quantity is meters, which we want to convert to millimeters. Because the original unit is assumed to be in the numerator, to get rid of it, we want the meter unit in the *denominator*; then they will cancel. Using the positive metric exponents, we have an alternative method to calculate the same answer as shown previously. Canceling units and performing the mathematics, we get:

$$14.66 \cancel{m} \times \frac{1000 mm}{1 \cancel{m}} = 14660 mm \quad (1.4.4)$$

$$= 1.46 \times 10^4 mm \quad (1.4.5)$$

Note how m cancels, leaving mm , which is the unit of interest.

The ability to construct and apply proper conversion factors is a very powerful mathematical technique in chemistry.

Video: Unit Conversion and Dimensional Analysis

Check out [this video by Melissa Maribel](#) and answer the questions below.

<https://www.youtube.com/watch?v=0W1e-dAnRrE>



The following two questions come directly from the video.

Below are some additional examples on converting units for you to practice.

✓ Example 1.4.1

- Convert 35.9 kL to liters.
- Convert 555 nm to meters.

Solution

- We will use the fact that 1 kL = 1,000 L. Of the two conversion factors that can be defined, the one that will work is 1000L/1kL. Applying this conversion factor, we get:

$$35.9 \cancel{\text{kL}} \times \frac{1000 \text{ L}}{1 \cancel{\text{kL}}} = 35,900 \text{ L}$$

- We will use the fact that 1 nm = 1/1,000,000,000 m, which we will rewrite as 1,000,000,000 nm = 1 m, or $10^9 \text{ nm} = 1 \text{ m}$. Of the two possible conversion factors, the appropriate one has the nm unit in the denominator:

$$\frac{1 \text{ m}}{10^9 \text{ nm}}$$

Applying this conversion factor, we get

$$555 \cancel{\mu m} \times \frac{1 m}{10^9 \cancel{\mu m}} = 0.000000555 m$$

$$= 5.55 \times 10^{-7} m$$

In the final step, we expressed the answer in scientific notation.

? Exercise 1.4.1

- Convert 67.08 μL to liters.
- Convert 56.8 m to kilometers.

Answer a

$$6.708 \times 10^{-5} \text{ L}$$

Answer b

$$5.68 \times 10^{-2} \text{ km}$$

? Exercise 1.4.2

How many milliliters are in 607.8 kL?

Answer

This problem involves multiple steps that require going through the metric base of liters. Think of a pathway of a conversion before setting up these types of problems. The suggested pathway would be $\text{mL} \Rightarrow \text{L} \Rightarrow \text{kL}$.

$$607.8 \text{ kL} \times \frac{1000 \text{ L}}{1 \text{ kL}} = 6.08 \times 10^5 \text{ L}$$

$$= 6.08 \times 10^5 \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}}$$

$$= 608000000 \text{ mL}$$

$$= 6.08 \times 10^8 \text{ mL}$$

? Exercise 1.4.3

Kathleen Hsu Barron (Furman University, Class of 2011) is exploring Switzerland with her husband Patrick Barron (Furman University, Class of 2011). The sign she is pointing to is a typical speed in km/hr. Convert this distance to miles/hr and feet/hr.



Answer

Again, this problem is a multi-step problem. There are multiple ways that one could proceed to solve this problem. With the conversion factors in this course, the suggested round would be miles $\Rightarrow$ kilometers $\Rightarrow$ meters $\Rightarrow$ centimeters $\Rightarrow$ inches $\Rightarrow$ feet. Another method could utilize miles $\Rightarrow$ feet.

18.6 miles/hr, 9.8×10^4 ft/hr

In this course, we will not use significant figures. You will need to read the directions on your assignment, quiz, or test and round accordingly.

Applications of Conversions and the Metric System

The skill of converting units can be helpful in many aspects of your life. Measuring medications can be done precisely by using milliliters instead of teaspoons (recall 1 tsp = 5.0 mL). Of course, a pharmacist would provide you with the appropriate syringe if this was needed. Sizes of syringes can vary from 1.0 to 10.0 mL for most at-home applications.

Marathons are measured using the kilometer unit. If you were raised outside the United States, you would be quite familiar with how long a kilometer truly is. To appreciate this distance, most Americans need to convert this to mileage by dividing the distance by 1.61 km.

When traveling abroad, keep in mind the majority of countries (except Myanmar and Liberia) commonly use the metric system in day to day living. Understanding how to convert Metric to English units will help you to appreciate the measurements used by a different society.

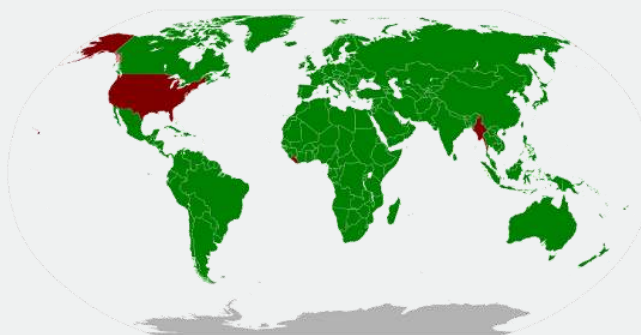


Figure 1.4.1: Adoption of the metric system of measurement. Countries that have officially adopted the metric system (green) and countries which have not officially adopted the metric system (US, Myanmar, Liberia)

Why has the United States not adopted the metric system as a country? Watch this [video](#) and see if you can derive some appropriate hypotheses.

Contributors and Attributions

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)
- Makena Song (Furman University)
- Erin Avram (Cleveland State University)

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1.5: Density

Learning Objectives

- Calculate density, mass, or volume when given 2 of these three variables.
- Identify what units are required for the density equation.
- Compare densities of different chemical substances.
- Compare any chemical substance's density to the density of water.

[Template:HideTOC](#)

Density is a physical property that is defined as a substance's mass divided by its volume:

$$\text{density} = \frac{\text{mass}}{\text{volume}} \quad (1.5.1)$$

$$d = \frac{m}{V} \quad (1.5.2)$$

Density is usually a measured property of a substance. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has *derived units*, just like velocity. Common units for density include g/mL, g/cm³, g/L, kg/L, and even kg/m³. Densities for some common substances are listed in Table 1.5.1.

Table 1.5.1: Densities of Some Common Substances

Substance	Density (g/mL or g/cm ³)
water	1.0
gold	19.3
mercury	13.6
air	0.0012
cork	0.22–0.26
aluminum	2.7
iron	7.87



Figure 1.5.1: Separatory Funnel containing oil and colored water to display density differences. (CC BY-SA 3.0; PRHaney via [Wikipedia](#)).

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3 . How can you determine what mass of aluminum you have without measuring it? You can use the volume and density to calculate it. If you multiply the given volume by the known density (Table 1.5.1), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

Start with Equation 1.5.1

$$\text{density} = \frac{m}{V}$$

and insert the relevant numbers

$$\frac{2.7g}{\text{cm}^3} = \frac{\text{mass}}{7.88 \text{ cm}^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

$$7.88 \cancel{\text{cm}^3} \times \frac{2.7g}{\cancel{\text{cm}^3}} = 21g \text{ of aluminum}$$

✓ Example 1.5.1: Mercury

What is the mass of 44.6 mL of mercury?

Solution

Use the density value for mercury from Table 1.5.1 and the definition of density (Equation 1.5.1)

$$\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow d = \frac{m}{V}$$

$$\frac{13.6g}{\text{mL}} = \frac{m}{44.6 \text{ mL}}$$

Remember to cross multiply here in order to isolate variable. Then, report answer with correct units.

$$44.6 \cancel{\text{mL}} \times \frac{13.6g}{\cancel{\text{mL}}} = 607g$$

The mass of the mercury is 607 g.

? Exercise 1.5.1

What is the mass of 25.0 cm^3 of iron?

Answer

Use the density value for iron from Table 1.5.1

$$\text{density} = \frac{\text{mass}}{\text{volume}} \Rightarrow d = \frac{m}{V}$$

$$\frac{7.87g}{\text{cm}^3} = \frac{m}{25.0 \text{ cm}^3}$$

Cross multiplying both sides (right numerator x left denominator = left numerator x right denominator), we get the following expression with answer and appropriate unit.

$$25.0 \cancel{\text{cm}^3} \times \frac{7.87g}{\cancel{\text{cm}^3}} = 197g \text{ of iron}$$

Another way of looking at density (some students choose to perform calculations using this method)

Density can also be used as a conversion factor to convert mass to volume—but care must be taken. We have already demonstrated that the number that goes with density normally goes in the numerator when density is written as a fraction. Take the density of gold, for example:

$$d = 19.3 \text{ g/mL} = \frac{19.3 \text{ g}}{\text{mL}}$$

Although this was not previously pointed out, it can be assumed that there is a 1 in the denominator:

$$d = 19.3 \text{ g/mL} = \frac{19.3 \text{ g}}{\text{mL}}$$

That is, the density value tells us that we have 19.3 grams for every 1 milliliter of volume, and the 1 is an exact number. When we want to use density to convert from mass to volume, the numerator and denominator of density need to be switched—that is, we must take the *reciprocal* of the density. In so doing, we move not only the units but also the numbers:

$$\frac{19.3 \cancel{\text{g}}}{\text{mL}} = \frac{45 \cancel{\text{g}}}{V}$$

Cross-multiplying denominators with numerators, we obtain the following algebraic equation.

$$19.3V = 45.9 \text{ mL}$$

then you will need to isolate the variable (volume)

$$V = \frac{45.9 \text{ mL}}{19.3}$$

After multiplication, the answer would be

$$V = 2.38 \text{ mL}$$

✓ Example 1.5.2: Wine Cork

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/mL, what is the volume of the cork? *Regardless of the method that is used, you should still be able to obtain the same (and correct) answer.*

Solution

To use density as a conversion factor, we need to take the reciprocal so that the mass unit of density is in the denominator. Taking the reciprocal, we find

$$\frac{0.22 \cancel{\text{g}}}{\text{mL}} = \frac{3.78 \cancel{\text{g}}}{V}$$

Cross multiplying denominators with numerators, we obtain the following algebraic equation.

$$0.22V = 3.78 \text{ mL}$$

then you will need to isolate the variable (volume)

$$V = \frac{3.78 \text{ mL}}{0.22}$$

so, the volume of the cork is 17.2 mL.

? Exercise 1.5.2

What is the volume of 3.78 g of gold?

Answer

Before attempting this question, be sure to obtain the density of gold in the table above. If you were to need this value on a quiz or a test, then it would be provided. Once you have this value, plug it into the density equation. Next, you will need to isolate the volume variable (basic algebra). The final answer should be 0.196 cm^3 .

Care must be used with density as a conversion factor. Make sure the mass units are the same or the volume units are the same, before using density to convert to a different unit. Often, the unit of the given quantity must be first converted to the appropriate unit before applying density as a conversion factor.

Using Density in Environmental Applications

Along with solubility, density can help determine how a contaminant could affect an aquatic system. For example, imagine mercury has been spilled in Lake Erie. Looking at this element's density value and comparing it to liquid water, one could determine the location of the insoluble mercury layer (you would also need solubility information). Since mercury is more dense than water, a layer of mercury would settle on the bottom of the lake. If one were to take a cross-section of the lake, you could see that a heterogeneous mixture would result.

In contrast, spilling ethanol (density = 0.789 g/mL) would result in the formation of a homogeneous mixture. Ethanol (grain alcohol) is soluble in water. This would make it miscible (mixable to form a solution) in water and one would not be able to denote separate layers. According to the density, an alcohol layer would remain on top, but would ultimately dissolve.

Applications

What difficulties would arise from the separation and removal of contaminants?

1. Hg in Lake Erie
2. Ethanol in Lake Erie
3. oil (less dense, insoluble) in Lake Erie

Watch this [video](#) and record your observations.

1. What component was different in the two types of beverages (mass or volume)?
2. How does the above-mentioned difference affect the density equation?
3. Which beverage is denser than water?

Contributors and Attributions

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- Hayden Cox (Furman University)
- Erin Avram (Cleveland State University)

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1.6: Temperature

Learning Objectives

- Learn about the various temperature scales that are commonly used in chemistry and other cultures.
- Use temperature equations for **Celsius and Fahrenheit conversions**.
- Understand how body temperature can vary.

Temperature Scales

One of the fundamental quantities in science is temperature, which is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called degrees. There are three scales used for reporting temperatures. Figure 1.6.1 compares the three temperature scales.

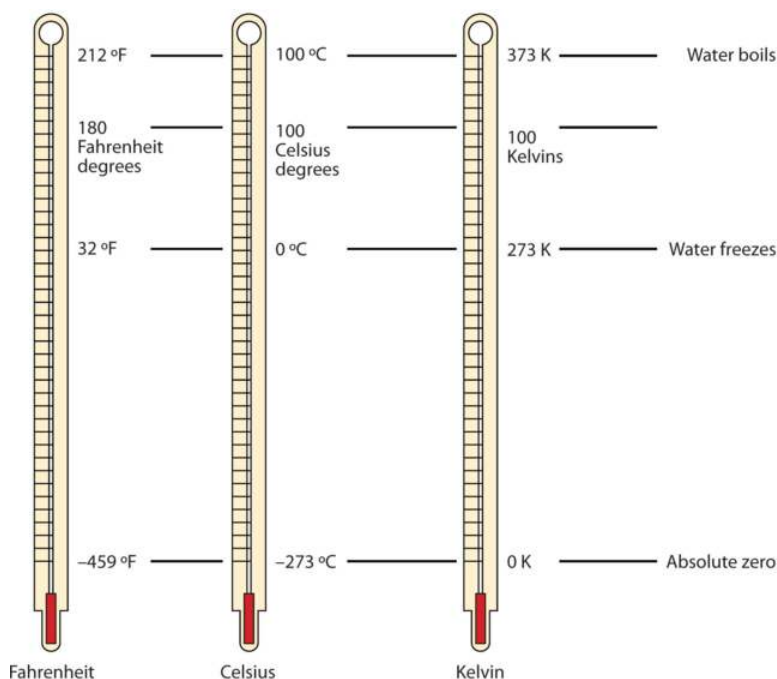


Figure 1.6.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales. These thermometers have a red-colored liquid that contains alcohol and food coloring. Silver-looking thermometers contain mercury, which is a neurotoxin.

In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as “degrees Fahrenheit”). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32 °F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212 °F.

Scientists also use other scales to express temperature. For example, the Celsius scale (symbolized by °C and spoken as “degrees Celsius”) defines 0°C as the freezing point of water and 100 °C as the boiling point of water. This scale is divided into 100 divisions between these two landmarks and extended higher and lower as well. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

$$^{\circ}\text{C} = (^{\circ}\text{F} - 32) \times \frac{5}{9} \quad (1.6.1)$$

$$^{\circ}\text{F} = \left(^{\circ}\text{C} \times \frac{9}{5}\right) + 32 \quad (1.6.2)$$

Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. (In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. People driving near the borders of Canada or Mexico may pick up local radio

stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.)

✓ Example 1.6.1: Conversions

- What is 98.6°F in degrees Celsius?
- What is 25.0°C in degrees Fahrenheit?

Solution

- Using Equation 1.6.1, we have

$$\begin{aligned}\text{°C} &= (98.6 - 32) \times \frac{5}{9} \\ &= 66.6 \times \frac{5}{9} \\ &= 37.0\text{°C}\end{aligned}$$

- Using Equation 1.6.2, we have

$$\begin{aligned}\text{°F} &= (25.0 \times \frac{9}{5}) + 32 \\ &= 45.0 + 32 \\ &= 77.0\text{°F}\end{aligned}$$

? Exercise 1.6.1

- Convert 0°F to degrees Celsius.
- Convert 212°C to degrees Fahrenheit.

Answer a

−17.8°C

Answer b

414°F

The fundamental unit of temperature in SI is the Kelvin (K). The Kelvin temperature scale (note that the name of the scale capitalizes the word *Kelvin*, and the unit itself is also uppercase) uses degrees that are the same size as the Celsius degree, but the numerical scale is shifted up by 273.15 units. That is, the conversion between the Kelvin and Celsius scales is as follows:

$$K = \text{°C} + 273.15 \quad (1.6.3)$$

For most purposes, it is acceptable to use 273 instead of 273.15 in Equation 1.6.3.

Note that the Kelvin scale does not use the word *degrees* or the degree symbol °; a temperature of 295 K is spoken of as “two hundred ninety-five kelvin” and not “two hundred ninety-five degrees Kelvin.”

The reason that the Kelvin scale is defined this way is that there exists a minimum possible temperature called absolute zero (zero Kelvin). The Kelvin temperature scale is set so that 0 K is absolute zero, and the temperature is counted upward from there. Normal room temperature is about 295 K, as seen in the following example.

✓ Example 1.6.2: Room Temperature

If the normal room temperature is 72.0°F, what is room temperature in degrees Celsius and kelvin?

Solution

First, we use Equation 1.6.1 to determine the temperature in degrees Celsius:

$$\begin{aligned} ^\circ C &= (72.0 - 32) \times \frac{5}{9} \\ &= 40.0 \times \frac{5}{9} \\ &= 22.2^\circ C \end{aligned}$$

Then we use Equation 1.6.3 to determine the temperature in the Kelvin scale:

$$\begin{aligned} K &= 22.2^\circ C + 273.15 \\ &= 295.4K \end{aligned}$$

So, the room temperature is about 295 K.

📌 Health Application: Body temperature

Normal body temperature is defined as being 98.6°F (+/- 1.0°F). To determine body temperature, thermometers can be placed inside or on the surface of the body. The two best methods of obtaining body temperature are placing the thermometer either under the tongue or inside the rectum. Typically, children are capable of holding a thermometer in their mouths around the age of four (have fun before that age).

Fever is defined as body temperature being above 100°F (adults). High fever status occurs at and above 104°F. For adults, these adults should seek medical attention immediately if the fever exceeds 104°F. For children, these values are much lower.

Hyperthermia (inability of the body to regulate heat) occurs when normal body temperature is exceeded. Conditions that can cause hyperthermia are fever (infection), heat stroke, thyroid disorders, heart attack, or traumatic injury. Medications for cancer, arthritis, and thyroid patients can cause the body temperature to rise. Symptoms of hyperthermia include sweating, confusion, nausea, and dizziness.

Hypothermia (exposure to cold environments) occurs when normal body temperature dips below 95.0°F. When this occurs, the affected person(s) should seek immediate medical attention. During hypothermia, the body has problems producing heat. Medical conditions like diabetes, infection, and thyroid dysfunction can cause hypothermia. Watch this [video](#) of American Marines attempting to survive extreme environments. Symptoms of this condition involve shivering, confusion, and sluggish behavior.

Treating hyperthermia could involve hydrating a patient. If a fever is due to infection, analgesics (like Tylenol, Advil, aspirin, or Aleve) can help alleviate fever. Placing someone in a cool bath can also relieve symptoms.

To combat hypothermia, one must remove wet clothing, redress in warm materials, and participate in physical movement.

The Science of Cryogenics

Cryogenics is a branch of physics that incorporates the production and effects of substances at temperatures ranging between -150°C to -273°C. In the late 1870s, cryogenic science began with the cooling of gaseous oxygen. At -183°C, O₂ condensed from a clear gas into a blue liquid. Changing the state of oxygen enables it to be transported more easily. Both the aircraft and medical industries rely on the liquid form of this molecule to provide breathable oxygen for pilots, travelers, and patients.



Figure 1.6.2: Liquid oxygen (O_2) exhibits a bluish hue.

Besides oxygen, other gases have been condensed as well. These permanent gases, meaning these substances normally exist in the gaseous state, include nitrogen (N_2), hydrogen (H_2), chlorine (Cl_2), and helium (He). By manipulating pressure and/or temperature, many chemists and physicists have attempted to produce the lowest, theoretical temperature possible. At absolute zero ($-273^\circ C$ or $0 K$), the particles of a material have a minimum kinetic energy. Click on this [video](#) to determine if scientists can ever reach absolute zero in their laboratories.

📌 Investigating a Cryogenic material: $N_2(l)$

With a boiling point of $-196^\circ C$ ($-321^\circ F$), liquid nitrogen has a variety of cryogenic uses. This clear, odorless liquid is used widely in the field of dermatology. For patients that have warts, skin tags (see image below), or precancerous moles, physicians can spray a selected area of skin. The unwanted skin lesion will typically fall off the body after exposure to this cryogenic material. If this does not occur, then the dermatologist will use a scalpel to remove the frozen area of the skin.



Figure 1.6.1: Liquid nitrogen is stored in insulated flasks called Dewars. Image courtesy of Elizabeth R. Gordon

Liquid N_2 is also utilized to freeze biological samples. Fertility clinics freeze semen, eggs, and embryos that can be used at a later date for couples who choose to undergo In Vitro Fertilization (IVF). Umbilical cord blood and stem cells can be preserved in liquid nitrogen for future applications.

Culinary Arts employ liquid nitrogen for preparing various dishes and interesting cocktails. In the following video, chefs are freezing components of salads and desserts for their customers. When using liquid nitrogen for any application, chemists, chefs, and other individuals should protect their skin and eyes from the extremely cold ($-196^\circ C$) temperature. Wearing thick rubber gloves and goggles can prevent frostbite. When ingesting materials that have been frozen with liquid nitrogen, there is

concern that internal burns can occur. Lastly, indoor use of liquid nitrogen could induce asphyxiation. Kitchens and labs should have adequate ventilation to ensure sufficient oxygen remains in the atmosphere.



Video 1.6.2: At the Bazaar restaurant: Cooking with liquid nitrogen, more

References

- <https://www.bbc.com/news/magazine-19870668>
- <https://ehs.research.uiowa.edu/liqui...rogen-handling>

Contributors and Attributions

- Elizabeth R. Gordon (Furman University)
- Hayden Cox (Furman University)
- Erin Avram (Cleveland State University)

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CHAPTER OVERVIEW

2: Naming and Structure of Compounds

2.1: Subatomic Particles - Electrons, Protons, and Neutrons

2.2: Atomic Number and Mass Number

2.3: Isotopes

2.4: The Importance of Ions to a Chemist

2.5: Lewis Dot and Bonding

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2.8: Molar Mass to Mole Conversions

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2.1: Subatomic Particles - Electrons, Protons, and Neutrons

Learning Objectives

- Know the basics of the experiments involving the discoveries of the three subatomic particles
- Memorize relative charge values and amu masses of the three subatomic particles.
- Describe the location of the proton, electron, and neutron inside the atomic model.

Identifying and Locating the Subatomic Particles

In 1897, the British physicist J. J. Thomson (1856–1940) proved that atoms were not the most basic form of matter. His experiments involved the use of cathode ray tubes (CRT). Under a vacuumed environment, these devices produce invisible rays that originate from a negative electrode (cathode) and continue to a positive electrode (anode). In these experiments, Thomson demonstrated that cathode rays could be deflected, or bent, by magnetic or electric fields, which indicated that cathode rays consist of charged particles. More importantly, by measuring the extent of the deflection of the cathode rays in magnetic or electric fields of various strengths, Thomson was able to calculate the mass-to-charge ratio of the particles. Because similar charges repel each other and opposite charges attract, Thomson concluded that the particles had a net **negative charge**; these particles are now called **electrons**. Most relevant to the field of chemistry, Thomson found that the mass-to-charge ratio of cathode rays is independent of the nature of the metal electrodes or the gas, which suggested that electrons were fundamental components of all atoms.



Video 2.1.1: An instructor explains nature of a cathode ray tube and the discovery of the electron.

During the 1910's, experiments with x-rays led to this useful conclusion: the magnitude of the **positive charge** in the nucleus of every atom of a particular element is the same. In other words, all atoms of the same element have the same number of **protons**. Furthermore, different elements have a different number of protons in their nuclei, so the number of protons in the nucleus of an atom is characteristic of a particular element. This discovery was so important to our understanding of atoms that the number of protons in the nucleus of an atom is called the atomic number. These experiments were performed by Ernest Rutherford. He is credited with defining a central location of the atom which has now been named the nucleus.

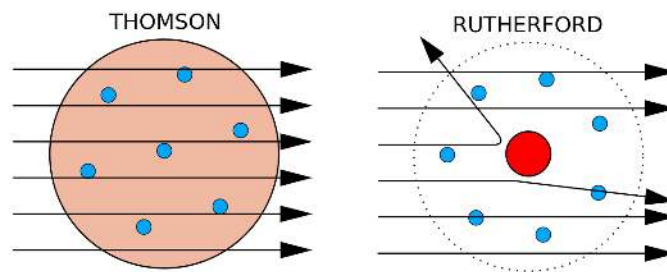


Figure 2.1.1: Expected results: alpha particles passing through the plum pudding model of the atom with negligible deflection. Bottom: Observed results: a small portion of the particles were deflected by the concentrated positive charge of the nucleus. (Public Domain).

Rutherford was aware that the protons were inside this region. One of Ernest Rutherford's colleagues, James Chadwick, performed and studied radioactive reactions of beryllium. Upon hitting a beryllium nucleus with alpha particles, a **neutron** particle was emitted. In 1932, James Chadwick announced the existence of a third subatomic particle, the neutron. This particle has a mass of 1 atomic mass unit, but **does not have a charge**. After this discovery, nuclear science and technology started immediately.

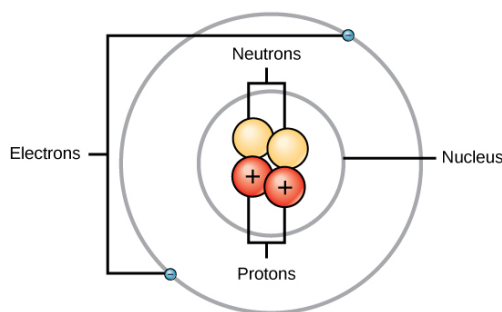


Figure 2.1.2: Elements, such as helium, depicted here, are made up of atoms. Atoms are made up of protons and neutrons located within the nucleus, with electrons in orbitals surrounding the nucleus.

Masses for the three subatomic particles can be expressed in amu (**atomic mass units**) or grams. For simplicity, we will use the amu unit for the three subatomic particles. Both neutrons and protons are assigned as having masses of 1 amu each. In contrast, the electron has a negligible mass of 0.0005 amu. This subatomic particle's mass is not represented in the decimal mass that is displayed on the periodic table. Only protons and neutrons contribute to an atom's mass value.

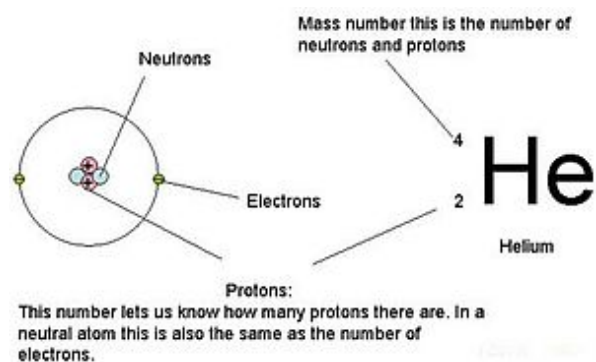


Figure 2.1.3: An explanation of the superscripts and subscripts seen in atomic number notation. (Copyright;

en>User:MaterialsScientist)

For atoms (neutral species), the number of electrons is equivalent to the number of protons. As a result, most atoms have an overall charge of zero. If an atom gains or loses an electron, it forms an ion. Ions will be discussed later in a later section of the text. Atoms of an element can also differ in the number neutrons and mass values. Two atoms of the same element (same number of protons) but different mass numbers (different number of neutrons) are called isotopes. These particular species are discussed in the following section of the text.

Table 2.1.1: Properties of Subatomic Particles

Subatomic particle	Symbol	Relative charge	Location	Amu mass
Proton	p	+1	inside the nucleus	1
Neutron	n	0	inside the nucleus	1
Electron	e ⁻	-1	outside the nucleus	0.0005 (~0)

✓ Example 2.1.1

Which is true?

- a. The nucleus contains protons and electrons.
- b. The mass of an electron is included in the mass of the atom
- c. For a neutral atom, the number of protons = number of electrons.

Solution

- a. False, the nucleus contains the protons and the neutrons. Electrons are outside the core.
- b. False, electrons do not contribute to the mass number of an atom.
- c. True, all neutral atoms must have the same number of protons as electrons.

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2.2: Atomic Number and Mass Number

Learning Outcomes

- Define atomic number and mass number.
- Determine the number of protons, neutrons, and electrons in an atom.
- Identify the charge and relative mass of subatomic particles.
- Label the location of subatomic particles in the atom.
- Determine the mass of an atom based on its subatomic particles.
- Write A/Z and symbol-mass format for an atom.

Atoms are the fundamental building blocks of all matter and are composed of protons, neutrons, and electrons. Because atoms are typically electrically neutral, the number of positively charged protons must be equal to the number of negatively charged electrons. Since neutrons do not affect the charge, the number of neutrons is not dependent on the number of protons and will vary even among atoms of the same element.

Atomic Number

The **atomic number (represented by the letter Z)** of an element is the number of protons in the nucleus of each atom of that element. An atom can be classified as a particular element based solely on its atomic number. For example, any atom with an atomic number of 8 (its nucleus contains 8 protons) is an oxygen atom, and any atom with a different number of protons would be a different element. The periodic table (see figure below) displays all of the known elements and is arranged in order of increasing atomic number. In this periodic table, an element's atomic number is indicated above the elemental symbol. Hydrogen, at the upper left of the table, has an atomic number of 1. Every hydrogen atom has one proton in its nucleus. Next on the table is helium, whose atoms have two protons in the nucleus. Lithium atoms have three protons, beryllium atoms have four, and so on.

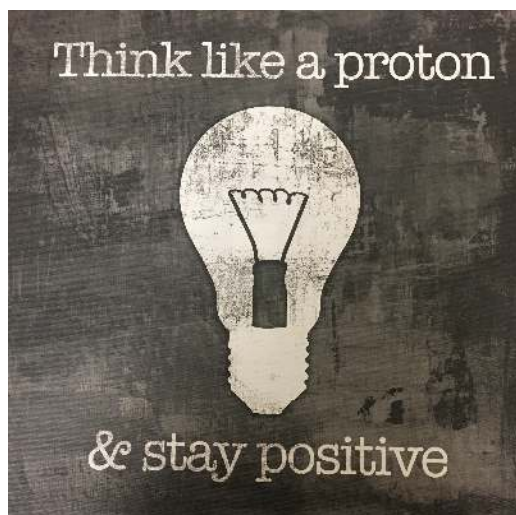


Figure 2.2.1: The subatomic social security number: the proton.

Since atoms are typically neutral (overall charge = 0), the number of electrons (each with a charge of -1) in an atom is equal to the number of protons (each with a charge of +1). Hydrogen atoms all have one electron occupying the space outside of the nucleus. Helium, with two protons, will have two electrons. In the chemical classroom, the number of protons will always be equivalent to an atom's atomic number. This value will not change unless the nucleus decays or is bombarded (nuclear chemistry).

$$\text{Charge} = \text{number of protons} - \text{number of electrons}$$

	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
①	Hydrogen 1 H 1.008																	Helium 2 He 4.0026
②	Lithium 3 Li 6.94	Beryllium 4 Be 9.0122											Boron 5 B 10.81	Carbon 6 C 12.011	Nitrogen 7 N 14.007	Oxygen 8 O 15.999	Fluorine 9 F 18.998	Neon 10 Ne 20.180
③	Sodium 11 Na 22.990	Magnesium 12 Mg 24.305											Aluminum 13 Al 26.982	Silicon 14 Si 28.085	Phosphorus 15 P 30.974	Sulfur 16 S 32.06	Chlorine 17 Cl 35.45	Argon 18 Ar 39.948
④	Potassium 19 K 39.098	Calcium 20 Ca 40.078	Scandium 21 Sc 44.956	Titanium 22 Ti 47.867	Vanadium 23 V 50.942	Chromium 24 Cr 51.996	Manganese 25 Mn 54.938	Iron 26 Fe 55.845	Cobalt 27 Co 58.933	Nickel 28 Ni 58.693	Copper 29 Cu 63.546	Zinc 30 Zn 65.38	Gallium 31 Ga 69.723	Germanium 32 Ge 72.630	Arsenic 33 As 74.922	Selenium 34 Se 78.971	Bromine 35 Br 79.904	Krypton 36 Kr 83.798
⑤	Rubidium 37 Rb 85.468	Strontium 38 Sr 87.62	Yttrium 39 Y 88.906	Zirconium 40 Zr 91.224	Niobium 41 Nb 92.906	Molybdenum 42 Mo 95.95	Technetium 43 Tc [98]	Ruthenium 44 Ru 101.07	Rhodium 45 Rh 102.91	Palladium 46 Pd 106.42	Silver 47 Ag 107.87	Cadmium 48 Cd 112.41	Indium 49 In 114.82	Tin 50 Sn 118.71	Antimony 51 Sb 121.76	Tellurium 52 Te 127.60	Iodine 53 I 126.90	Xenon 54 Xe 131.29
⑥	Cesium 55 Cs 132.91	Barium 56 Ba 137.33	Lanthanum 57 La 138.91	* 72 Hf 178.49	Tantalum 73 Ta 180.95	Tungsten 74 W 183.84	Rhenium 75 Re 186.21	Osmium 76 Os 190.23	Iridium 77 Ir 192.22	Platinum 78 Pt 195.08	Gold 79 Au 196.97	Mercury 80 Hg 200.59	Thallium 81 Tl 204.38	Lead 82 Pb 207.2	Bismuth 83 Bi 208.98	Polonium 84 Po [209]	Astatine 85 At [210]	Radon 86 Rn [222]
⑦	Francium 87 Fr [223]	Radium 88 Ra [226]	Actinium 89 Ac [227]	* 104 Rf [267]	Dubnium 105 Db [268]	Seaborgium 106 Sg [269]	Bohrium 107 Bh [270]	Hassium 108 Hs [270]	Meitnerium 109 Mt [278]	Darmstadtium 110 Ds [281]	Roentgenium 111 Rg [282]	Copernicium 112 Cn [285]	Nihonium 113 Nh [286]	Flerovium 114 Fl [289]	Moscovium 115 Mc [290]	Livermorium 116 Lv [293]	Tennessine 117 Ts [294]	Oganesson 118 Og [294]
				* 58 Ce 140.12	Praseodymium 59 Pr 140.91	Neodymium 60 Nd 144.24	Promethium 61 Pm [145]	Samarium 62 Sm 150.36	Europium 63 Eu 151.96	Gadolinium 64 Gd 157.25	Terbium 65 Tb 158.93	Dysprosium 66 Dy 162.50	Holmium 67 Ho 164.93	Erbium 68 Er 167.26	Thulium 69 Tm 168.93	Ytterbium 70 Yb 173.05	Lutetium 71 Lu 174.97	
				* 90 Th 232.04	Protactinium 91 Pa 231.04	Uranium 92 U 238.03	Neptunium 93 Np [237]	Plutonium 94 Pu [244]	Americium 95 Am [243]	Curium 96 Cm [247]	Berkelium 97 Bk [247]	Californium 98 Cf [251]	Einsteinium 99 Es [252]	Fermium 100 Fm [257]	Mendelevium 101 Md [258]	Nobelium 102 No [259]	Lawrencium 103 Lr [260]	

Figure 2.2.3: The periodic table of the elements. (CC BY-SA 4.0 International; DePiep via Wikipedia).

Mass Number

Experimental data showed that the vast majority of the mass of an atom is concentrated in its nucleus, which is composed of protons and neutrons. The **mass number (represented by the letter A)** is defined as the total number of protons and neutrons in an atom. Consider the table below, which shows data from the first six elements of the periodic table.

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

Table 2.2.1: Atoms of the First Six Elements

Name	Symbol	Atomic Number (Z)	Protons	Neutrons	Electrons	Mass Number (A)
hydrogen	H	1	1	0	1	1
helium	He	2	2	2	2	4
lithium	Li	3	3	4	3	7
beryllium	Be	4	4	5	4	9
boron	B	5	5	6	5	11
carbon	C	6	6	6	6	12

Consider the element helium. Its atomic number is 2, so it has two protons in its nucleus. Its nucleus also contains two neutrons. Since $2 + 2 = 4$, we know that the mass number of the helium atom is 4. Finally, the helium atom also contains two electrons, since the number of electrons must equal the number of protons. This example may lead you to believe that atoms have the same number of protons and neutrons, but a further examination of the table above will show that this is not the case. Lithium, for example, has three protons and four neutrons, giving it a mass number of 7.

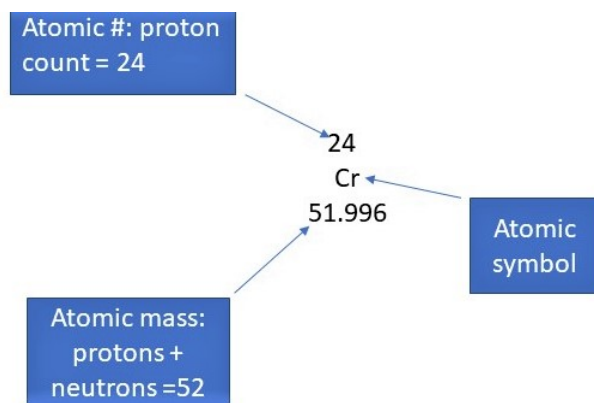


Figure 2.2.4: Determining the subatomic particles for the element Chromium from the periodic table (Copyright; Elizabeth R. Gordon)

Knowing the mass number and the atomic number of an atom allows you to determine the number of neutrons present in that atom by subtraction.

$$\text{Number of neutrons} = \text{mass number} - \text{atomic number}$$

Atoms of the element chromium (**Cr**) have an atomic number of 24 and a mass number of 52. How many neutrons are in the nucleus of a chromium atom? To determine this, you would subtract as shown:

$$52 - 24 = 28 \text{ neutrons in a chromium atom}$$

The composition of any atom can be illustrated with a shorthand notation called A/Z format. Both the atomic number and mass are written to the **left** of the chemical symbol. The "A" value (mass number) is written as a superscript while the "Z" value (atomic number) is written as a subscript. For an example of this notation, look to the chromium atom shown below:



Another way to refer to a specific atom is to write the mass number of the atom after the name, separated by a hyphen. Symbol-mass format for the above atom would be written as chromium-52 or Cr-52. In this notation, the atomic number is not included. You will need to refer to a periodic table for the number of protons.

✓ Example 2.2.1

For each atom, determine the number of protons, neutrons, and electrons.

- mercury-200
- platinum-195
- bromine-80

Solutions

- Hg (transition metal)- has 80 electrons, 80 protons, and 120 neutrons
- Pt (transition metal)- has 78 electrons, 78 protons, and 117 neutrons
- Br (halogen)- has 35 electrons, 35 protons, and 45 neutrons

✓ Example 2.2.2

Write both A/Z and symbol-mass formats for the atoms in Example 2.2.1.

Solutions

- ${}^{200}_{80}\text{Hg}$ and Hg-200
- ${}^{195}_{78}\text{Pt}$ and Pt-195

c. $^{80}_{35}\text{Br}$ and Br-80

✓ Example 2.2.3

Identify the elements based on the statements below.

- Which element has 25 protons?
- Which element has 0 neutrons?
- Which element has 83 electrons?

Solutions

- manganese
- hydrogen
- bismuth

Contributors and Attributions

- Elizabeth Gordon (Furman University)
- Allison Soult, Ph.D. (Department of Chemistry, University of Kentucky)
- Erin Avram (Cleveland State University)

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2.3: Isotopes

Learning Objectives

- Calculate subatomic particles (protons, neutrons, and electrons) for any element by looking at the periodic table.
- Calculate subatomic particles (protons, neutrons, and electrons) for any element by looking at symbol-mass and A/Z format.
- Understand how isotopes differ in particles and mass.
- Identify the most abundant isotope when given specific values.
- Calculate the atomic mass of an element from the masses and relative percentages of the isotopes of the element.

Isotopes are atoms of the same element that contain different numbers of neutrons. For these species, the number of electrons and protons remains constant. This difference in the number of neutrons affects the atomic mass (**A**) but not the atomic number (**Z**). In a chemical laboratory, isotopes of an element appear and react the same. For this reason, it is difficult to distinguish between an atom's isotopes. In contrast, nuclear scientists can identify and separate different types of atomic nuclei (Figure 2.3.1). The technology required for this process is more sophisticated than what could be found in a typical chemical laboratory.

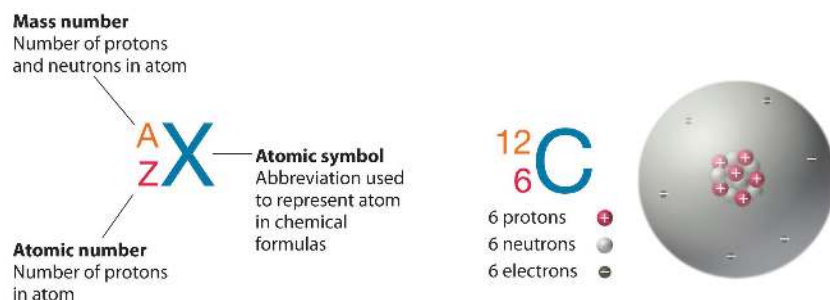


Figure 2.3.1: Unlike protons, the number of neutrons is not absolutely fixed for most elements. Atoms that have the same number of protons, and hence the same atomic number, but different numbers of neutrons are called isotopes. All isotopes of an element have the same number of protons and electrons, which means they have the same chemical properties. The isotopes of an element differ only in their atomic mass, which is given by the mass number (**A**), the sum of the numbers of protons and neutrons. (CC BY-NS-SA; anonymous by request)

The element carbon (**C**) has an atomic number of 6, which means that all neutral carbon atoms contain 6 protons and 6 electrons. In a typical sample of carbon-containing material, 98.89% of the carbon atoms also contain 6 neutrons, so each has a mass number of 12. An isotope of any element can be uniquely represented as ${}_Z^AX$ where **X** is the atomic symbol of the element. The isotope of carbon that has 6 neutrons is therefore ${}^{12}_6\text{C}$. The subscript indicating the atomic number is actually redundant because the atomic symbol already uniquely specifies **Z**. Consequently, it is more often written as ${}^{12}\text{C}$, which is read as “carbon-12.” Nevertheless, the value of **Z** is commonly included in the notation for *nuclear* reactions because these reactions involve changes in **Z**.

Most elements on the periodic table have at least two stable isotopes. For example, in addition to ${}^{12}\text{C}$, a typical sample of carbon contains 1.11% ${}^{13}\text{C}$, with 7 neutrons and 6 protons, and a trace of ${}^{14}\text{C}$, with 8 neutrons and 6 protons. The nucleus of ${}^{14}\text{C}$ is not stable, however, but undergoes a slow radioactive decay that is the basis of the carbon-14 dating technique used in archaeology. Many elements other than carbon have more than one stable isotope; tin, for example, has 10 isotopes. There are about twenty elements that exist in only one isotopic form (sodium and fluorine are examples of these). Most scientists cannot tell you how many isotopic forms exist unless they consult an isotope table.

An important series of isotopes is found with hydrogen atoms. Most hydrogen atoms have a nucleus with only a single proton. About 1 in 10,000 hydrogen nuclei, however, also has a neutron; this particular isotope is called *deuterium* (a hydrogen atom with 1 proton, 1 electron, and 1 neutron). An extremely rare hydrogen isotope, *tritium*, has 1 proton and 2 neutrons in its nucleus. Figure 2.3.2 compares the three isotopes of hydrogen.

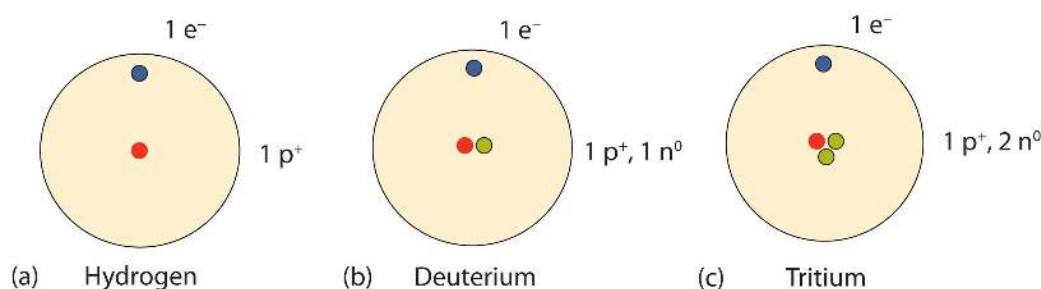


Figure 2.3.2: a). A small amount of hydrogen exists as the isotope deuterium, which has one proton and one neutron in its nucleus (b). A tiny amount of the hydrogen isotope tritium, with one proton and two neutrons in its nucleus, also exists on Earth (c). The nuclei and electrons are proportionately much smaller than depicted here.

There are currently over 3,500 isotopes known for all the elements. When scientists discuss individual isotopes, they need an efficient way to specify the number of neutrons in any particular nucleus. *A/Z* and symbol-mass formats (refer to Section 3.4) can be used to display periodic table information. When viewing either of these two notations, isotopic differences can be obtained.

Look at the **A/Z formats** for the three isotopes of hydrogen in Table 2.3.1. Note how the atomic number (bottom value) remains the same while the atomic masses (top number) are varied. All isotopes of a particular element will vary in neutrons and mass. This variance in mass will be visible in the **symbol-mass format** of isotopes of the same element as well.

Table 2.3.1

Common Name	A/Z formats	symbol-mass format	Expanded Name
Hydrogen	${}^1_1\text{H}$	H-1	hydrogen-1
Deuterium	${}^2_1\text{H}$	H-2	hydrogen-2
Tritium	${}^3_1\text{H}$	H-3	hydrogen 3

Using a periodic table, A/Z or symbol-mass formats can be utilized to determine the number of subatomic particles (protons, neutrons, and electrons) contained inside an isotope. When given either format, these mass values should be used to calculate the number of neutrons in the nucleus.

✓ Exercise 2.3.1

How many neutrons are in each atom?

- ${}^{36}_{17}\text{Cl}$
- ${}^{58}_{26}\text{Fe}$
- C-14

Solution

- In ${}^{36}_{17}\text{Cl}$ there are 19 neutrons in this nucleus. To find this value, subtract A-Z or $36 - 17$.
- In ${}^{58}_{26}\text{Fe}$ there are 32 neutrons in this nucleus. Again, subtract A-Z or $58 - 26$.
- In this example, the C-14 represents symbol-mass format. Once the atomic number is located (look at periodic table), subtract $14 - 6$. The final answer will be 8 neutrons.

? Exercise 2.3.2

For the species below, translate the A/Z format to symbol-mass. Then, calculate the number of subatomics for each atom.

- ${}^{197}_{79}\text{Au}$
- ${}^{23}_{11}\text{Na}$
- ${}^{239}_{94}\text{Pu}$

Answer a

Au-197 contains 79 protons, 79 electrons, and 118 neutrons.

Answer b

Na-23 contains 11 protons, 11 electrons, and 12 neutrons.

Answer c

Pu-239 contains 94 protons, 94 electrons, and 145 neutrons.

Calculating Atomic Mass

The **atomic mass** of an element is the weighted mass of all the naturally presented isotopes. On the periodic table, it is assumed that this mass has units of amu (atomic mass unit) which can be abbreviated by using the letter u. To determine the most abundant isotopic form of an element, compare given isotopes to the weighted average on the periodic table. For example, the three hydrogen isotopes (shown above) are H-1, H-2, and H-3. The atomic mass or weighted average of hydrogen is around 1.008 amu (look again at the periodic table). Of the three hydrogen isotopes, H-1 is closest in mass to the weighted average; therefore, it is the most abundant. The other two isotopes of hydrogen are rare but are very exciting in the world of nuclear science.

✓ Example 2.3.2

Identify the true statements:

- a. Al-27 is more abundant than Al-25.
- b. An appropriate isotope of bromine could be Br-35.
- c. A chemist could easily distinguish between Cs-132 and Cs-133 by noting chemical and physical properties.
- d. Most scientists know that calcium has 24 isotopes.

Answer a

This statement is true. Aluminum with a mass of 27 is closest to the mass on the periodic table. It would be more abundant than Al-25.

Answer b

This statement is false. An appropriate mass number of an isotope of bromine would be in around 80 amu (atomic mass units), not 35 (which is the atomic number).

Answer c

Chemists cannot distinguish between isotopes by looking at various properties. Most isotopes have similar solubilities, densities, and colors.

Answer d

Unless a scientist works heavily with a particular atom, he or she is not aware of the many forms that cannot exist. They would need a reference guide to know how many natural and artificial isotopes exist for a particular element.

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

$$\text{Atomic mass} = (\%_1) (\text{mass}_1) + (\%_2) (\text{mass}_2) + \dots \quad (2.3.1)$$

To confirm your answer, compare the calculated value to the weighted mass displayed on the periodic table.

An element's atomic mass can be calculated provided the relative abundances of the element's naturally occurring isotopes and the masses of those isotopes are known. If all

the abundances are not provided, it is safe to assume that all numbers should add up to 100%.

✓ Example 2.3.3: Atomic mass of Boron

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

Solution

Boron has two isotopes. We will use the Equation [2.3.1](#):

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

Substitute these into the equation, and we get:

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

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

✓ Example 2.3.4: Atomic mass of Neon

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

Solution

Neon has three isotopes. We will use the equation:

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

Substitute these into the equation, and we get:

$$\begin{aligned}\text{Atomic mass} &= (0.9092)(19.99) + (0.003)(20.99) + (0.0885)(21.99) \\ &= 20.18 \text{ amu}\end{aligned}$$

The mass of an average neon atom is **20.18 amu**

📌 Applications of Isotopes

During the Manhattan Project, the majority of federal funding was dedicated to the separation of uranium isotopes. The two most common isotopes of uranium are U-238 and U-235. About 99.3% of uranium is of the U-238 variety, this form is not fissionable and will not work in a nuclear weapon or reaction. The remaining 0.7% is U-235 which is fissionable, but first had to be separated from U-238. This separation process is called *enrichment*. During World War II, a nuclear facility was built in Oak Ridge, Tennessee to accomplish this project. At the time, the enrichment process only produced enough U-235 for one nuclear weapon. This fuel was placed inside the smaller of the two atomic bombs (Little Boy) and dropped over Hiroshima, Japan on August 6, 1945, and is estimated to have caused between 70,000 and 140,000 deaths.



Figure 2.3.3: A billet of highly enriched uranium that was recovered from scrap processed at the Y-12 National Security Complex Plant. Original and unrotated.

Uranium is a natural element that can be found in several different countries. Countries that do not have natural uranium supplies would need to obtain it from one of the countries below. Most nuclear reactors that provide energy rely on U-235 as a source of fuel. Fortunately, reactors only need 2-5% U-235 for the production of megawatts or even gigawatts of power. If the purification process exceeds this level, then it is likely a country that is focusing on making nuclear weapons. For example, the Manhattan Project scientists enriched U-235 up to 90% to produce the Little Boy weapon.

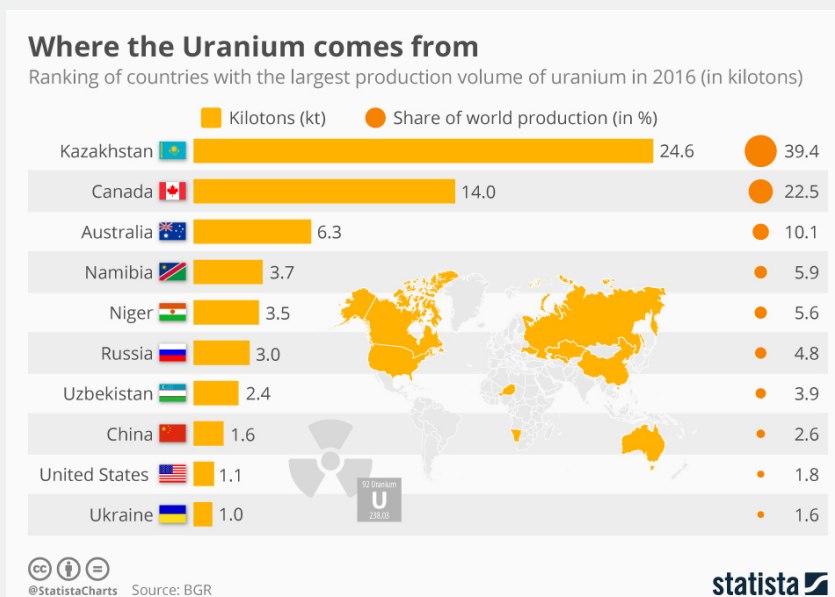


Figure 2.3.4: This infographic shows a ranking of countries with the biggest production volume of uranium in 2016 (in kilotons). (CC BY-ND; statista).

Abbreviations like HEU (highly enriched uranium) and LEU (low-enriched uranium) are used frequently by nuclear scientists and groups. HEU is defined as being over 20% pure U-235 and would not be used in most commercial nuclear reactors. This type of material is used to fuel larger submarines and aircraft carriers. If the purification of U-235 reaches 90%, then the HEU is further classified as being weapons-grade material. This type of U-235 could be used to make a nuclear weapon (fission or even fusion-based). As for LEU, its U-235 level would be below this 20% mark. LEU would be used for commercial nuclear reactors and smaller, nuclear-powered submarines. LEU is not pure enough to be used in a conventional nuclear weapon but could be used in a dirty bomb. This type of weapon uses conventional explosives like dynamite to spread nuclear material. Unlike a nuclear weapon, dirty bombs are not powerful enough to affect large groups of buildings or people. Unfortunately, the spread of nuclear material would cause massive chaos for a community and would result in casualties.

Contributors and Attributions

- cK-12.org
- [Elizabeth R. Gordon \(Furman University\)](#)
- Erin Avram (Cleveland State University)

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2.4: The Importance of Ions to a Chemist

Learning Objectives

- Identify how a charge affects the number of electrons of an atom.
- Apply a provided charge when calculating the number of subatomic particles.
- Define and identify cations and anions
- Locate regions of the periodic table where cations and anions tend to form.
- Compare and contrast chemical and nuclear energy.

[Template:HideTOC](#)

Chemists appreciate isotopes and use them in basic and applied research. However, they are more concerned with the movement of electrons. When an atom gains or loses electrons, it becomes a charged species or an ion. When this occurs, the nucleus is not changed. For atoms that lose electrons, an overall positive charge will result (**#protons > #electrons**). Atoms that form these types of ions are called **cations**. Metal atoms (located on the left side of the periodic table) always lose electrons to become cations.

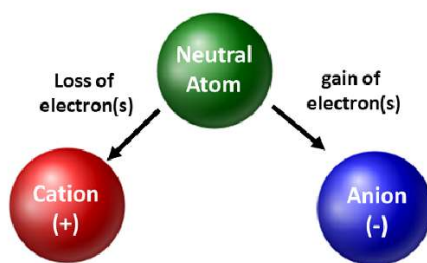


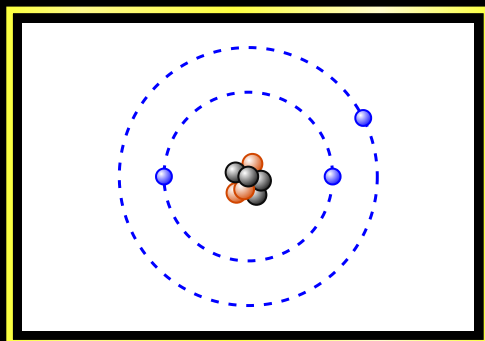
Figure 2.4.1: Neutral atoms become ions by either losing an electron (to form positively charged cations) or gaining electrons (to form negatively charged anions). (Public Domain; Delmar Larsen)

Unlike metal atoms, nonmetals will gain electrons to become **anions**. These types of ions have an overall negative charge (**#electrons > #protons**). With the exception of the noble gases, all atoms on the periodic table will lose or gain electrons to achieve electronic stability. Different types of bonding occur when atoms lose, gain, or share electrons. These types of atomic connections will be further discussed later in the textbook.

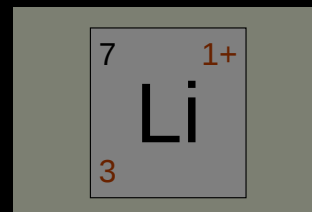
Interactive: Building an Atom

Build an atom out of protons, neutrons, and electrons, and see how the element, charge, and mass change. Then play a game to test your ideas!

Build an Atom



Atom



Symbol

At this point, you should be able to calculate all subatomic particles when given a specific ion charge. In addition, you should master classifications of ions (recognizing cations are positive and anions are negative). Note and appreciate the ion charges for main group elements in the periodic table below. Lastly, chemists do not write a number for charges involving the number one (1). For these species, you will only see a + or - sign. Charges of higher magnitude will always have a number associated with the elemental symbol.

1A						8A
H ⁺	2A					
Li ⁺						
Na ⁺	Mg ²⁺					
K ⁺	Ca ²⁺					
Rb ⁺	Sr ²⁺					

	3A	4A	5A	6A	7A	
			N ³⁻	O ²⁻	F ⁻	
	Al ³⁺		P ³⁻	S ²⁻	Cl ⁻	
				Se ²⁻	Br ⁻	
					I ⁻	

Figure 2.4.2: Periodic Table with ion charges of common main group elements. The charge that an atom acquires when it becomes an ion is related to the structure of the periodic table. Within a group (family) of elements, atoms form ions of a certain charge. Please note that hydrogen can adopt a + or - charge. When the word hydride is used, this means that hydrogen has a - charge.

✓ Example 2.4.1: Ionic Classification

Calculate the subatomic particles for the species below using the information in Figure 2.4.2. Label each as being an atom, cation, or anion. Refer to the periodic table for masses, atomic numbers, and specific ion charges.

- Aluminum ion
- Zirconium atom
- Sulfur ion

Solutions

- There are 13 protons, 14 neutrons, and 10 electrons in the aluminum ion. This ion is positively charged which means it has lost electrons and forms a cation.
- There are 40 protons, 51 neutrons, and 40 electrons in the zirconium atom. This is an atom that has no overall charge.
- There are 16 protons, 16 neutrons, and 18 electrons in the sulfur ion. This ion is negatively charged which means it has gained electrons and forms an anion.

✓ Example 2.4.2: A/Z and symbol mass notations

Write A/Z and symbol mass notations for the ions/atom in the previous question.

Solutions

- $^{27}_{13}\text{Al}^{3+}$ and Al-27
- $^{91}_{40}\text{Zr}$ and Zr-91
- $^{32}_{16}\text{S}^{2-}$ and S-32

If an atom gains or loses electrons, this should be represented on the A/Z format. The symbol-mass format does not display ion charge. Remember, neutrons and protons are not affected by changes of electrons.

Contributors

- Elizabeth R. Gordon (Furman University)
- Erin Avram (Cleveland State University)

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2.5: Lewis Dot and Bonding

Learning Objectives

- Identify and describe covalent bonds and ionic bonds.
- Draw the Lewis dot structure for an atom.

Bonding (Ionic and Covalent Basics)

When two atoms approach each other, they have the potential to bond (or connect). If a **metal and a nonmetal interact**, then an **ionic bond** will result. These types of bonds involve the metal donating its valence electron(s) to a nonmetal, forming an ionic compound. As the transfer of electrons occurs, both atoms will achieve a more stable confirmation. The end result will be a less reactive compound. These ionic species are composed of both cations (positively charged ions) and anions (negatively charged ions). In addition, ionic compounds are crystalline and solid in nature. A few examples of real-world ionic compounds would be NaCl (sodium chloride, table salt) and NaF (sodium fluoride, the active ingredient in toothpaste).



Figure 2.5.1: NaCl crystals. (Public Domain; NASA).

If **two nonmetals** interact, then a **covalent bond** will result. The connection that forms is due to each atom *sharing* its valence electron(s). In this type of bonding, each atom (except for H, B, and Be) obtains an octet (eight valence electrons) and becomes stable. Metalloids and nonmetals can bond to form covalent compounds as well. Solid covalent compounds are not crystalline in texture, unlike ionic compounds. Covalent compounds can be liquids or even gases.



Figure 2.5.2: Covalent compounds NO_2 and N_2O_4 contained in glass bottles. Image taken from: commons.wikimedia.org/wiki/File:NO2-N2O4.jpg

Table 2.5.1: Comparing ionic and covalent compounds

Ionic Bonding	Covalent Bonding
Metal and nonmetal combination	nonmetal and nonmetal or metalloid and nonmetal combination
Metal loses valence electrons and nonmetal gains up to eight.	Both elements share valence electrons to achieve stability
Individual atoms have charges (anions or cations)	No charges are present in these compounds
All ionic compounds are solids	Covalent compounds can be solid, liquid, or gas.
Examples could include: Li_2CO_3 , Fe_2O_3 , and MgSO_4	Examples include: CO_2 , H_2O , and $\text{CH}_3\text{CH}_2\text{OH}$

Lewis Symbols

At the beginning of the 20th century, an American physical chemist G. N. Lewis (1875–1946) devised a system of symbols—now called *Lewis electron dot symbols* (often shortened to *Lewis dot symbols*) that can be used for predicting the number of bonds formed by most elements in their compounds. Each Lewis dot symbol consists of the chemical symbol for an element surrounded by dots that represent its valence electrons. Lewis knew that incorporating models into his teaching would enable students to visualize chemical bonding more easily. Gilbert Lewis was known for his interactive teaching methods and his interest in student success. Unfortunately, his nervous personality limited his ability to lecture in front of large groups of people.

Regarding his research, Lewis constructed his own theory to explain the nature of acids and bases. He was the first to synthesize D_2O (heavy water) and wrote numerous textbooks for his courses. In World War I, he served as a Major in the Gas Service unit of the American Army. He trained over 200 soldiers a week on how to adequately protect themselves from gas warfare. From these efforts, he helped reduce fatalities due to chemical gas exposures and was awarded the Distinguish Service Medal in 1922.

Lewis was nominated over thirty times for the Nobel Prize. He never received this award and does not have an element named in honor of him. For more on Gilbert Lewis, click on this [link](#) to read a bibliographic memoir published by the National Academy of Sciences.

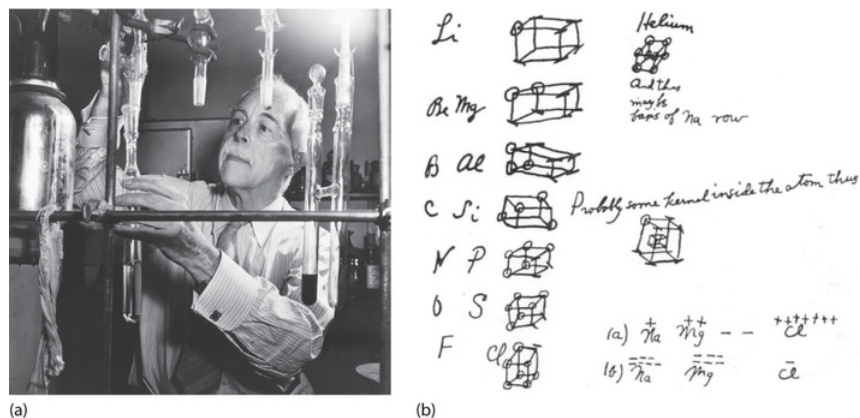


Figure 2.5.3: G. N. Lewis and the Octet Rule. (a) Lewis is working in the laboratory. (b) In Lewis's original sketch for the octet rule, he initially placed the electrons at the corners of a cube rather than placing them as we do now.

📌 Lewis Dot symbols

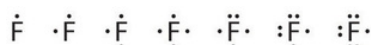
- convenient representation of valence electrons
- allows scientists to keep track of valence electrons during bond formation
- consists of the chemical symbol for the element plus a dot for each **valence** electron

To write an element's Lewis dot symbol, we place dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more

than four valence electrons, are again distributed one at a time, each paired with one of the first four. For example, the element sulfur has six valence electrons (this can be determined by looking at the Roman numeral listed above the element's group on the periodic table) and its Lewis symbol would be:



Fluorine, for example, has seven valence electrons, so its Lewis dot symbol is constructed as follows:



The number of dots in the Lewis dot symbol is the same as the number of valence electrons, which is the same as the last digit of the element's group number in the periodic table. Lewis dot symbols and electron configurations (not a topic for this course) for the elements in period 2 are given in Table 2.5.2.

Table 2.5.2: Lewis Dot Symbols for the Elements in Period 2

Group	1	2	13	14	15	16	17	18
Electron Configuration	[He]2s ¹	[He]2s ²	[He]2s ² 2p ¹	[He]2s ² 2p ²	[He]2s ² 2p ³	[He]2s ² 2p ⁴	[He]2s ² 2p ⁵	[He]2s ² 2p ⁶
Lewis Dot Symbol	Li ·	·Be·	·B·	·C·	·N·	:O·	:F·	:Ne:

Ionic compounds are produced when a metal bonds with a nonmetal. Stability is achieved for both atoms once the transfer of electrons has occurred. The image below shows how sodium and chlorine bond to form the compound sodium chloride. Unlike sodium metal, the resulting compound is not explosive and is less corrosive than chlorine gas. Ionic bonding can be viewed by noting the donation of valence electrons from a metal atom to a nonmetal atom by using the Bohr model. From this theory, we will move on to the Lewis structure with an understanding that metals will always lose valence electrons, and nonmetals will gain up to eight to form stable compounds.

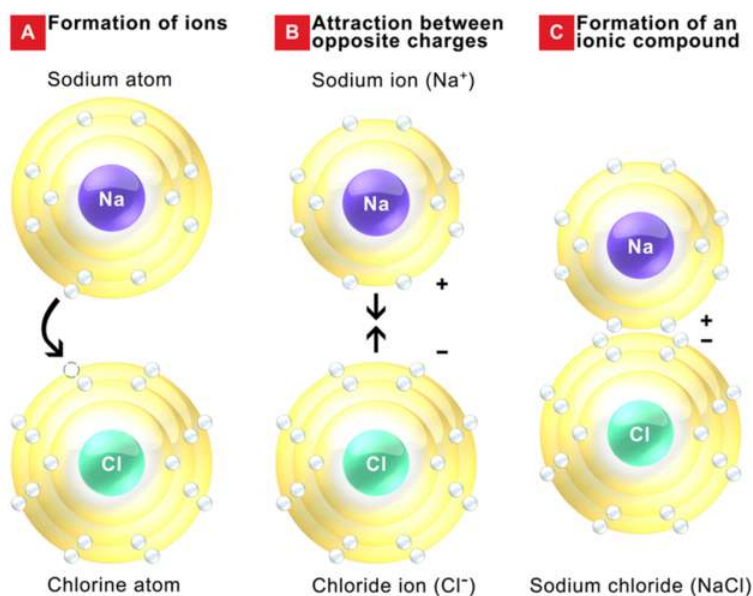


Figure 2.5.4: Formation of an ionic bond between sodium and chlorine. Image courtesy of By BruceBlais - Own work, CC BY-SA 4.0, commons.wikimedia.org/w/index.php?curid=44968540

Rules for Drawing Ionic Lewis structures

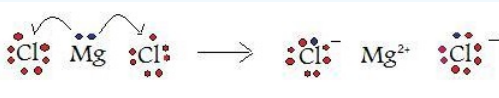
When drawing Lewis structures for compounds formed from the combination of a metal and a nonmetal, use the list of rules shown below:

- Locate the elements of interest and recall how many valence electrons each species has (look at the Roman numeral listed above the element's group).
- Draw dots around each element individually, one dot for each valence electron.
- To obtain an octet, the metal atom must transfer all of its valence electrons to the nonmetal.
- At this point, the metal should have no valence electrons around it (and have a positive charge). With the newly acquired electrons from the metal, the nonmetal atom will have an octet (and a negative charge). If this does not occur, then refer to step e.
- Additional atoms of the same two elements may be brought in to accomplish rules a - c.
- Cation and anion charges must be on Lewis structure to indicate the number of electrons gained or lost (this is an ionic compound).**
- To check the structure, write a formula for the ionic compound, the charges of all atoms in the compound must add up to zero.

✓ Example 2.5.1: Chloride Salts



In this example, the sodium atom donates its 1 valence electron to the chlorine atom. This creates a sodium cation (+ charge) and a chloride anion (- charge). Notice that the net charge of the resulting compound is 0 (+1 - 1 = 0). In this example, only one atom of each element is needed to obtain a stable compound.



In this next example, the magnesium atom is donating both of its valence electrons, one electron to each chlorine atom. Each chlorine atom can only accept 1 electron before it achieves an octet; therefore, 2 atoms of chlorine are required to accept the 2 electrons donated by the magnesium. Each chloride anion has a charge of -1, the magnesium ion has a charge of 2+. Notice that the net charge of the compound is 0 (+2 - 1 - 1 = 0). This Lewis structure requires two chloride anions per one magnesium cation to produce a stable compound.

Contributors and Attributions

- Mike Blaber (Florida State University)
- Wikipedia
- National Programme on Technology Enhanced Learning (India)
- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)
- Erin Avram (Cleveland State University)

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2.6: Ionic Formula Writing

Learning Objectives

- Write the names and formulas for binary ionic compounds
- Relate the Lewis structure to the chemical formula for binary ionic compounds.

Ionic Formulas

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula contains a cation and an anion. An ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is the total positive charge must balance the total negative charge. Sometimes this requires more than one of a cation or an anion to balance the overall positive and negative charges. It is conventional to use the lowest ratio of ions that are needed to balance the charges.

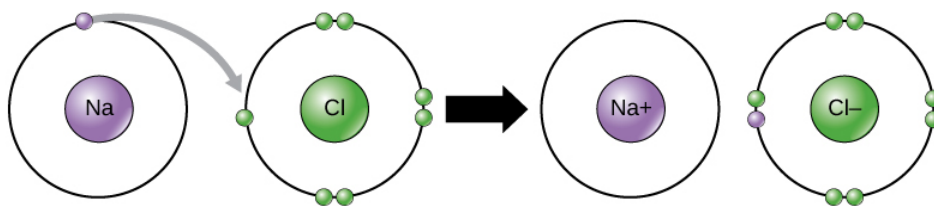


Figure 2.6.1: Formation of an ionic bond between sodium and chlorine to form sodium chloride. https://commons.wikimedia.org/wiki/File:NaCl_formation.jpg (CC BY-SA 4.0)

For example, consider the ionic compound between Na^+ and Cl^- . Each ion has a single charge, one positive and one negative, so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions:

1. write the formula for the cation first and the formula for the anion next
2. cross the charges down diagonally so the integer of the charge becomes the subscript for the other element

Thus, for the compound between Na^+ and Cl^- , we have the ionic formula NaCl (Figure 2.6.1). The formula Na_2Cl_2 also has balanced charges, but the convention is to use the lowest ratio of ions, which would be one of each. (Remember from our conventions for writing formulas that we do not write a 1 subscript if there is only one atom of a particular element present.) For the ionic compound between magnesium cations (Mg^{2+}) and oxide anions (O^{2-}), again we need only one of each ion to balance the charges. By convention, the formula is MgO .



Figure 2.6.2: NaCl is the formula for Table Salt. The ionic compound NaCl is very common. Red rock salt from the Khewra Salt Mine in Pakistan. (CC BY-SA 4.0; Hubertl).

Rules for Writing Ionic Formulas

Recall ionic compound can be of the following combination: metal/nonmetal or ammonium/nonmetal

1. Determine the charge on the cation. Then write the symbol of the cation with the charge written as a superscript to the right of the symbol of the cation (Ca^{2+}). If a metal can form different charges (most transition metals and metals below the metalloids), then the charge is indicated in the name of the compound by a Roman numeral (lead IV would translate to Pb^{4+}).
2. Determine the charge on the anion. Then write the symbol of the anion with the charge written to the right of the anion (S^{2-}). If a polyatomic ion is present, then write the formula for the polyatomic ion in parentheses with the charge written as a superscript to the right of the formula (SO_4^{2-}).
3. Cross the charges down diagonally so the integer of the charge becomes the subscript for the other element. Subscripts that are divisible by the same number must be reduced. Do not change the formula for the polyatomic ion. The parentheses around a polyatomic ion only stay if more than 1 ion is required to balance the charges between the anions and cations.

* Binary compound names will end with ide (except cyanide and hydroxide).

** Compounds that contain polyatomic ions will end in -ate, -ite, or -ide. A reference table of polyatomic ions is provided on your exam cover sheet**

✓ Example 2.6.1:

Write the formulas for the combinations of ions below (make sure the charge of all atoms together adds up to zero):

- a. Ca^{2+} and Cl^-
- b. Al^{3+} and F^-
- c. Li^+ and O^{2-}

Solution

- a. The proper ionic formula is CaCl_2 .
- b. The formula for this compound is AlF_3 .
- c. After crossing charges diagonally down, the formula will be Li_2O .

? Exercise 2.6.2

Write the proper ionic formulas for each pair of ions.

- a. Fe^{2+} and S^{2-}
- b. Fe^{3+} and S^{2-}

Answer a

FeS , Subscripts must be reduced for ionic compounds. If this is not done, then the answer will be incorrect.

Answer b

Fe_2S_3 , These subscripts cannot be reduced because they are not divisible by the same number.

There also exists a group of ions that contain more than one atom. These are called **polyatomic ions**. Table 2.6.1 lists the formulas, charges, and names of some common polyatomic ions. Only one of them, the ammonium ion, is a cation; the rest are anions. Most of them also contain oxygen atoms, so sometimes they are referred to as **oxyanions**. Some of them, such as nitrate and nitrite, and sulfate and sulfite, have very similar formulas and names, so care must be taken to get the formulas and names correct. Note that the -ite polyatomic ion has one less oxygen atom in its formula than the -ate ion but with the same ionic charge.

Table 2.6.1: Common Polyatomic Ions

Name	Formula and Charge	Name	Formula and Charge
ammonium	$(\text{NH}_4)^+$	nitrate	$(\text{NO}_3)^-$
acetate	$(\text{C}_2\text{H}_3\text{O}_2)^-$	nitrite	$(\text{NO}_2)^-$
bicarbonate (hydrogen carbonate)	$(\text{HCO}_3)^-$	carbonate	$(\text{CO}_3)^{2-}$
hydroxide	$(\text{OH})^-$	sulfate	$(\text{SO}_4)^{2-}$
chlorate	$(\text{ClO}_3)^-$	sulfite	$(\text{SO}_3)^{2-}$
hypochlorite	$(\text{ClO})^-$	phosphate	$(\text{PO}_4)^{3-}$
cyanide	$(\text{CN})^-$	phosphite	$(\text{PO}_3)^{3-}$

There are more polyatomic ions than the ones listed in the table above. You will be provided with a list of relevant polyatomic ions for use on exams.

Writing the formulas of ionic compounds has one important difference. If more than one polyatomic ion is needed to balance the overall charge in the formula, enclose the formula of the polyatomic ion in parentheses and write the proper numerical subscript to the right and *outside* the parentheses. Thus, the formula between calcium ions, Ca^{2+} , and nitrate ions, NO_3^- , is properly written $\text{Ca}(\text{NO}_3)_2$, not CaNO_3_2 or CaN_2O_6 . Refer back to your rules for writing ionic formulas.

✓ Example 2.6.4:

Write the formulas for each of the following ionic compounds:

- ammonium sulfide
- aluminum phosphate
- iron II phosphite

Solution

- Ammonium is a polyatomic ion (cation). These species will go first in the formula. Take ammonium $(\text{NH}_4)^+$ and combine it with sulfide S^{2-} (ide = periodic table, except for hydroxide/cyanide). While keeping the polyatomic formula within parentheses, cross the ion charges down and lose signs. The resulting formula will be **$(\text{NH}_4)_2\text{S}$** .
- Aluminum is Al^{3+} and phosphate is $(\text{PO}_4)^{3-}$. While keeping the polyatomic ion inside parentheses, cross charges down and reduce subscripts outside. The final formula for the compound will be **$\text{Al}(\text{PO}_4)$** . For more advanced chemistry courses, students are encouraged to drop the parentheses if a subscript is not present once charges have been crossed down. I am fine with you always keeping parentheses.
- Transition metal charges (except for silver and zinc) will always be noted by a Roman numeral. This charge is always positive for the metal ion. Place this charge to the top right corner of the metal atom, Fe^{2+} . Next, recall that phosphite is the polyatomic, $(\text{PO}_3)^{3-}$. Cross the charges down diagonally and keep the polyatomic inside the parentheses. The final result will be **$\text{Fe}_3(\text{PO}_3)_2$** .

? Exercise 2.6.4

Write the formulas for the following ionic compounds:

- ammonium bicarbonate
- cobalt III nitrite
- lead IV oxide
- potassium hydroxide
- silver phosphide

f. mercury I sulfite

Answer a

(NH_4)(HCO_3), there are no numbers outside the parentheses so you could write NH_4HCO_3 for the answer. Do not combine the hydrogen atoms in the formula.

Answer b

$\text{Co}(\text{NO}_2)_3$, the parentheses must be kept for this answer.

Answer c

PbO_2 , after the charges have been crossed, subscripts must be reduced by the division of common factor

Answer d

$\text{K}(\text{OH})$, you could drop ()

Answer e

Ag_3P , phosphide is from the periodic table, not the polyatomic table. Know your endings!!

Answer f

$\text{Hg}_2(\text{SO}_3)$, you could drop ()

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions (cations) and negative ions (anions).
- Ionic formulas balance the total positive and negative charges.

Contributors and Attributions

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)
- Erin Avram (Cleveland State University)

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2.7: Covalent Bonding and Formula Writing

Learning Objectives

- Describe the properties of a covalent bond
- Use Lewis structures to represent covalent bonding in molecular substances
- Use a Lewis structure to determine how many single, double, and triple bonds are present in a molecule
- Write the formulas and names for binary covalent compounds

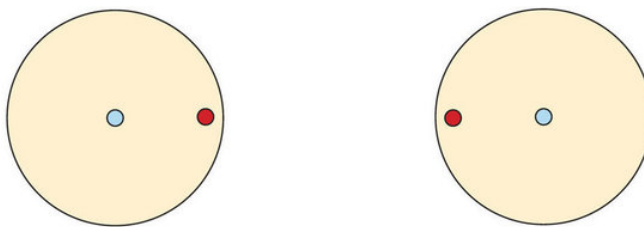
Electron Sharing

Previously, we discussed ionic bonding where electrons can be transferred from one atom to another so that both atoms have an energy-stable outer electron shell. Because most filled electron shells have eight electrons in them, chemists called this tendency the octet rule. However, there is another way an atom can achieve a full valence shell: atoms can *share* electrons. This type of bonding is called a covalent bond. Two combinations of atoms can produce this type of bonding: nonmetal with nonmetal or metalloid with nonmetal. **In this class, we will not discuss the option of metallic bonding which is a form of covalent bonding.**

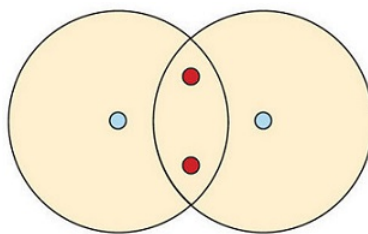


Figure 2.7.1: Sharing is caring, especially for atoms that participate in covalent bonding. (CC BY-2.0; [Carlos Maya](#)).

This concept can be illustrated by using two hydrogen atoms, each of which has a single electron in its valence shell. (For small atoms such as hydrogen atoms, the valence shell will be the first shell, which holds a maximum of only two electrons.) We can represent the two individual hydrogen atoms as follows (the blue circle represents the single proton found in each atomic nucleus, the red circle represents the single electron found in its valence shell):



In contrast, when two hydrogen atoms get close enough together to share their electrons, they can be represented as follows:



By sharing their valence electrons, both hydrogen atoms now have two electrons in their respective valence shells. Because each valence shell is now filled, this arrangement is more stable than when the two atoms are separate. The *sharing of electrons* between atoms is called a **covalent bond**, and the two electrons that join atoms in a covalent bond are called a bonding pair of electrons. A discrete group of atoms connected by covalent bonds is called a **molecule** — the smallest part of a compound that retains the chemical identity of that compound.

Chemists frequently use Lewis structures to represent covalent bonding in molecular substances. For example, the Lewis dot structures of two separate hydrogen atoms are as follows:



The Lewis diagram of two hydrogen atoms sharing electrons looks like this:



This depiction of molecules is simplified further by using a dash to represent a covalent bond. The Lewis structure of the hydrogen molecule is then represented as follows:



Remember that the dash, also referred to as a single bond, represents a *pair* of electrons.

The bond in a hydrogen molecule, measured as the distance between the two nuclei, is about 7.4×10^{-11} m, or 74 picometers (pm; $1 \text{ pm} = 1 \times 10^{-12} \text{ m}$). This particular bond length represents a balance between several forces: the attraction between oppositely charged electrons and nuclei, the repulsion between two negatively charged electrons, and the repulsion between two positively charged nuclei. If the nuclei were closer together, they would repel each other more strongly; if the nuclei were farther apart, there would be less attraction between the positive and negative particles.

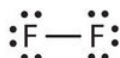
Fluorine is another element whose atoms bond together in pairs to form *diatomic* (two-atom) molecules. Two separate fluorine atoms have the following electron dot diagrams:



Each fluorine atom contributes one valence electron, making a single bond and giving each atom a complete valence shell, which fulfills the octet rule:



The circles show that each fluorine atom has eight electrons around it (a full octet). As with hydrogen, we can represent the Lewis structure of the fluorine molecule with a dash in place of the bonding electrons:



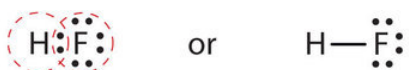
Each fluorine atom has six electrons, or three pairs of electrons, that are not participating in the covalent bond. Rather than being shared, they are considered to belong to a single atom. These are called nonbonding pairs (or **lone pairs**) of electrons.

Covalent Bonds between Different Atoms

Now that we have looked at electron sharing between atoms of the same element, let us look at covalent bond formation between atoms of different elements. Consider a molecule composed of one hydrogen atom and one fluorine atom:



Each atom needs one additional electron to complete its valence shell. By each contributing one electron, they make the following molecule:



In this molecule, the hydrogen atom does not have nonbonding electrons, while the fluorine atom has six nonbonding electrons (three lone electron pairs). The circles show how the valence electron shells are filled for both atoms.

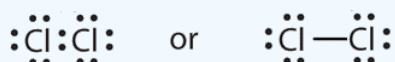
✓ Example 2.7.1

Draw the Lewis structure for each compound.

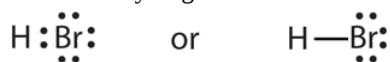
- a molecule composed of two chlorine atoms
- a molecule composed of a hydrogen atom and a bromine atom

Solution

- Chlorine has the same valence shell electron configuration as fluorine, so the Lewis structure for a molecule composed of two chlorine atoms is similar to the one for fluorine:



- Bromine has the same valence shell electron configuration as fluorine, so the Lewis structure for a molecule composed of a hydrogen atom and a bromine atom is similar to that for hydrogen and fluorine:

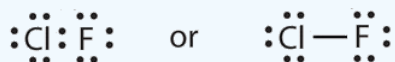


? Exercise 2.7.1

Draw the Lewis structure for each compound.

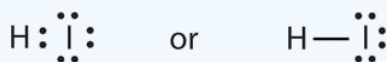
- a molecule composed of one chlorine atom and one fluorine atom
- a molecule composed of one hydrogen atom and one iodine atom

Answer a



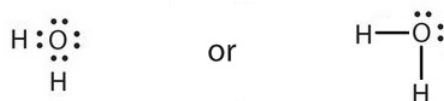
Chlorine monofluoride is a volatile interhalogen compound.

Answer b



HI is used in organic and inorganic synthesis as one of the primary sources of iodine.

Larger molecules are constructed in a similar fashion, with some atoms participating in more than one covalent bond. For example, water, with two hydrogen atoms and one oxygen atom can be represented as follows:



and methane (CH_4), with one carbon atom and four hydrogen atoms, can also be represented as follows:

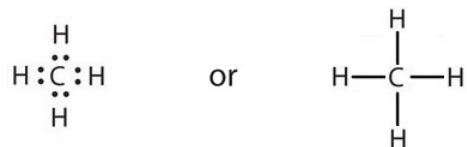


Figure 2.7.1 shows the number of covalent bonds various atoms typically form.

Rules for drawing Covalent Lewis Structures

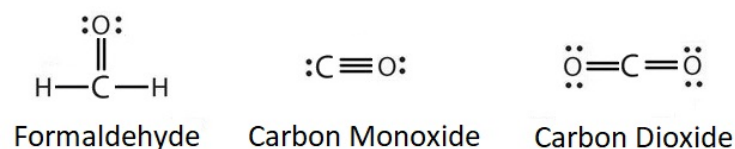
These apply to either nonmetal/nonmetal or metalloid/nonmetal combinations.

1. Tally up all valence electrons that are to be used. Remember valence electrons are the Roman numerals above the column the element is located in.
2. Determine which atom is to be central. Carbon is always central and hydrogen is a terminal atom. If these atoms are not present, then go with the atom you have the least of in the formula (typically this atom is listed first in the formula).
3. Connect each atom to the central atom using single bonds first.
4. For simple molecules (molecules with only two atoms), try single bonds, then double, and finally triple to achieve octets around each atom (except for B, Be, and H).
5. For more complex molecules (3 or more atoms), try one multiple bond at a time if single bonds don't achieve octets on every atom.
6. My series for molecules/ions that need more than single bonds can be seen below (keeping in mind H, B, and Be only have single bonds)
 - Single/single (—, —),
 - then single/double (—, =),
 - then double/double (=, =),
 - and if that does not work try triple/single ($\equiv$, —)
7. In our class, we do not use multiple bonds for halogens.

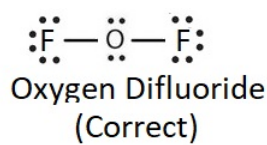
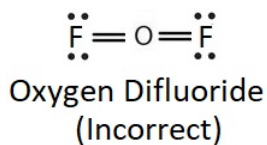
Double and Triple Bonds

In covalent bonding, using double (four total electrons) or triple (six total electrons) is an option to achieve a stable octet. If a single bond does not work, then you may incorporate a multiple bond. Please note the rules above on which atoms will never achieve octets (hydrogen, beryllium and boron). Although it is mathematically feasible, halogens will not use multiple bonds to become stable either.

Let's look at a few molecules that utilize multiple bonds:



Atoms capable of employing multiple bonds would be oxygen, nitrogen, sulfur, phosphorus, and carbon. If an octet can be achieved with a single bond, do not incorporate multiple bonds into the structure.



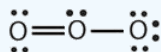
Multiple bonds involve more than one type of connection. For this reason, it is more difficult to break or cleave a multiple bond than it would be a single bond or a bond of a lower order (when comparing the same connected atoms). For example, the triple bond found in carbon monoxide is harder to break than one of the double bonds found in carbon dioxide. When reading about smog production or ozone depletion, consider the bond strengths of the chemicals involved in these environmental problems.

✓ Example 2.7.1

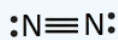
Draw the Lewis structure for each molecule:

- O₃ (ozone)
- N₂ (nitrogen)
- SO₃ (sulfur trioxide)
- CCl₄ (carbon tetrachloride)

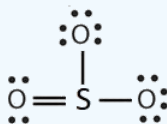
Solutions



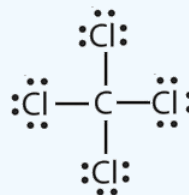
Ozone



Nitrogen



Sulfur Trioxide



Carbon Tetrachloride

Formula Writing for Covalent Compounds

Covalent compounds use prefixes to indicate the number of atoms in their formulas. In order to write correct formulas, you will need to **memorize** Table 2.7.1 below.

Table 2.7.1: Prefix to occurrence correlations.

Prefix name	Number of Atoms
Mono	1
Di (not bi)	2
Tri	3
Tetra (not quad)	4
Penta	5
Hexa	6
Hepta (not septa)	7
Octa	8
Nona	9
Deca	10

📌 Rules for Writing Covalent formulas

1. There are no charges involved with covalent compounds.
2. Interpret the prefixes to represent the number of atoms or the subscript.
3. Do not cross or reduce subscripts (unlike ionic compounds).
4. The formula should reflect the name of the compound.

✓ Example 2.7.1

Write formulas for the following covalent compounds:

- a. dinitrogen pentoxide
- b. boron trihydride
- c. sulfur hexafluoride
- d. carbon tetrachloride
- e. silicon dioxide
- f. dihydrogen dioxide

Answer a



Answer b



Answer c



Answer d



Answer e



Answer f

H_2O_2 , don't reduce subscripts in covalent compounds

Contributors and Attributions

- Elizabeth R. Gordon (Furman University)
- Isabella Quiros (Furman University)
- Erin Avram (Cleveland State University)

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2.8: Molar Mass to Mole Conversions

Learning Objectives

- Perform conversions between mass and moles of a substance.
- Convert from mass or moles of one substance to mass or moles of another substance in a chemical reaction.
- Use a balanced chemical equation to determine molar relationships between substances.

Molar Mass

The **molar mass** of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. To calculate the molar mass of a substance, add the atomic weight of every atom found in that substance. For example, water (or H_2O) consists of two atoms of hydrogen and 1 atom of oxygen. The molar mass equals the sum of the atomic weight of each atom in the molecule ($1.008 \text{ amu} + 1.008 \text{ amu} + 15.999 \text{ amu}$). For water, the molar mass is equal to 18.015 amu (or g/mol).

The simplest type of manipulation using molar mass as a conversion factor is a mole-mass conversion (or its reverse, a mass-mole conversion). In such a conversion, we use the molar mass of a substance as a conversion factor to convert mole units into mass units (or, conversely, mass units into mole units).

When considering the element, aluminum, we can use the periodic table of elements to look up the atomic weight of this element (26.98 amu). This value also means that 1 mol of Al has a mass of 26.98 g . Stated mathematically,

$$1 \text{ mol Al} = 26.98 \text{ g Al}$$

We can divide both sides of this expression by either side to get one of two possible conversion factors:

$$\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \text{ and } \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \quad (2.8.1)$$

The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do “by eye.”

✓ Example 2.8.1

What is the mass of 3.987 mol of Al ?

Solution

Example 2.8.1: Steps for Problem Solving the mass of 3.987 moles of aluminum (Al).

Steps for Problem Solving	Example
Identify the "given" information and what the problem is asking you to "find."	Given: 3.987 mol of Al Find: g Al
List other known quantities	1 mol Al = 26.98 g Al
Prepare a concept map and use the proper conversion factor.	 $\frac{26.98 \text{ g Al}}{1 \text{ mol Al}}$
Cancel units and calculate.	$3.987 \cancel{\text{ mol Al}} \times \frac{26.98 \text{ g Al}}{1 \cancel{\text{ mol Al}}} = 107.6 \text{ g Al}$
Think about your result.	The calculated value makes sense because it is almost four times times the mass for 1 mole of aluminum. Our final answer is expressed to four significant figures.

? Exercise 2.8.1

How many moles are present in 100. g of Al? (Hint: you will have to use the other conversion factor we obtained for aluminum.)

Answer

3.71 g Al

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 2.8.1 is a chart for determining what conversion factor is needed, and Figure 2.8.2 is a flow diagram for the steps needed to perform a conversion.

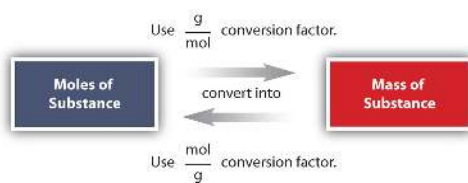


Figure 2.8.1 A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles.

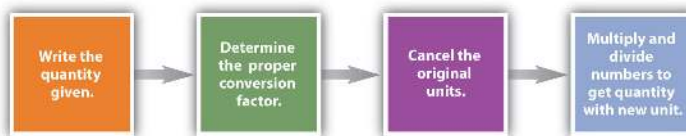


Figure 2.8.2 A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

Suppose that for a certain experiment you need 3.00 moles of calcium chloride (CaCl_2). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. Dimensional analysis will allow you to calculate the mass of CaCl_2 that you should measure as show in Example 2.8.3.

✓ Example 2.8.2: Calcium Chloride

Calculate the mass of 3.00 moles of calcium chloride (CaCl_2).



Figure 2.8.1: Calcium chloride is used as a drying agent and as a road deicer.

Solution

Example 2.8.2: Steps for Problem Solving the mass of 3.00 moles of calcium chloride (CaCl_2).

Steps for Problem Solving	Example
Identify the "given" information and what the problem is asking you to "find."	Given: 3.00 moles of CaCl_2 Find: g CaCl_2
List other known quantities (molar mass is calculated by adding the atomic weight of each atom in the molecule)	1 mol CaCl_2 = 110.98 g CaCl_2
Prepare a concept map and use the proper conversion factor.	 $\frac{110.98 \text{ g CaCl}_2}{1 \text{ mol CaCl}_2}$
Cancel units and calculate.	$3.00 \text{ mol CaCl}_2 \times \frac{110.98 \text{ g CaCl}_2}{1 \text{ mol CaCl}_2} = 333 \text{ g CaCl}_2$
Think about your result.	

? Exercise 2.8.2: Calcium Oxide

What is the mass of 7.50 mol of (calcium oxide) CaO ?

Answer:

420.60 g

✓ Example 2.8.3: Water

How many moles are present in 108 grams of water?

Solution

Example 2.8.3: Steps for Problem Solving moles in 108 grams of water?

Steps for Problem Solving	Example
Identify the "given" information and what the problem is asking you to "find."	Given: 108 g H ₂ O Find: mol H ₂ O
List other known quantities (molar mass is calculated by adding the atomic weight of each atom of the molecule or compound)	1 mol H ₂ O = 18.02 g H ₂ O
Prepare a concept map and use the proper conversion factor.	 $\frac{1 \text{ mol } H_2O}{18.02 \text{ g } H_2O}$
Cancel units and calculate.	$108 \cancel{\text{ g } H_2O} \times \frac{1 \text{ mol } H_2O}{18.02 \cancel{\text{ g } H_2O}} = 5.99 \text{ mol } H_2O$
Think about your result.	

? Exercise 2.8.3: Nitrogen Gas

What is the mass of 7.50 mol of Nitrogen gas N₂?

Answer:

210 g

Summary

- The molar mass of an atom, molecule, or compound can be used as a conversion factor to convert from grams to moles or from moles to grams.

Contributors and Attributions

- TextMap: The Basics of GOB Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Erin Avram (Cleveland State University)

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2.9: Chemical Equations

Learning Objectives

- To identify the parts of a chemical reaction (reactants, products, reaction arrow, subscripts and superscripts)
- To apply the Law of Conservation of Matter to balance chemical equations

What happens to matter when it undergoes chemical changes? The [Law of conservation of matter](#) says that "**Atoms are neither created, nor destroyed, during any chemical reaction.**" Thus, the same collection of atoms present before a chemical reaction must still be present afterward. Therefore, the changes that occur during a reaction will involve the *rearrangement* of atoms. In this section, we will discuss *stoichiometry* (the "measurement of elements").

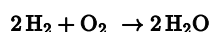
Chemical Equations

As shown in Figure 2.9.1, applying a small amount of heat to a pile of ammonium dichromate powder results in a vigorous reaction known as the ammonium dichromate volcano. Heat, light, and gas are produced as a large pile of fluffy green chromium (III) oxide forms. This reaction is described with a chemical equation, an expression that gives the identities and quantities of the substances in a chemical reaction.



Figure 2.9.1: An Ammonium Dichromate Volcano: Change during a Chemical Reaction. The starting material is solid ammonium dichromate. A chemical reaction transforms it to solid chromium(III) oxide, depicted showing a portion of its chained structure, nitrogen gas, and water vapor (in addition, energy in the form of heat and light is released). During the reaction, the distribution of atoms changes, but the number of atoms of each element does not change. Because the numbers of each type of atom are the same in the reactants and the products, the chemical equation is balanced. (CC BY-SA 3.0; Mikk Mihkel Vaabel via Wikipedia). See video here: www.youtube.com/watch?v=CW4hN0dYnkM

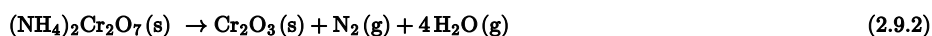
Chemical reactions are represented on paper by **chemical equations**. For example, hydrogen gas (H_2) can react (burn) with oxygen gas (O_2) to form water (H_2O). The *chemical equation* for this reaction is written as:



Chemical formulas and other symbols are used to indicate the starting materials, or **reactants**, which by convention are written on the left side of the equation, and the final compounds, or **products**, which are written on the right. A **reaction arrow** points from the reactant to the products. The chemical reaction for the ammonium dichromate volcano in Figure 2.9.1 is



The reaction arrow is read as "yields" or "reacts to form." Equation 2.9.1 indicates that ammonium dichromate (the reactant) *yields* chromium(III) oxide, nitrogen, and water (the products). The equation for this reaction is even more informative when written as follows:



Equation 2.9.2 is identical to Equation 2.9.1 except for the addition of abbreviations in parentheses to indicate the physical state of each species. The abbreviations are (s) for solid, (l) for liquid, (g) for gas, and (aq) for an aqueous solution, a solution of the substance in water.

Consistent with the law of conservation of matter, the numbers of each type of atom are the same on both sides of Equations 2.9.1 and 2.9.2. Each side of the reaction has two chromium atoms, seven oxygen atoms, two nitrogen atoms, and eight hydrogen atoms.

In a balanced chemical equation, both the numbers of each type of atom and the total charge are the same on both sides. Equations 2.9.1 and 2.9.2 are balanced chemical equations. What is different on each side of the equation is how the atoms are arranged to make molecules or ions. A chemical reaction represents a change in the distribution of atoms, but not in the number of atoms. In this reaction, and in most chemical reactions, bonds are broken in the reactants (here, Cr–O and N–H bonds), and new bonds are formed to create the products (here, O–H and N≡N bonds). If the numbers of each type of atom are different on the two sides of a chemical equation, then the equation is unbalanced, and it cannot correctly describe what happens during the reaction. To proceed, the equation must first be balanced. A chemical reaction changes only the distribution of atoms, not the number of atoms.



Introduction to Chemical Reaction Equations: [Introduction to Chemical Reaction Equations, YouTube](#)(opens in new window) [youtu.be]

Balancing Simple Chemical Equations

When a chemist encounters a new reaction, it does not usually come with a label that shows the balanced chemical equation. Instead, the chemist must identify the reactants and products and then write them in the form of a chemical equation that may or may not be balanced as first written. Consider, for example, the combustion of n-heptane (C_7H_{16}), an important component of gasoline:



The complete combustion of any hydrocarbon in the presence of sufficient oxygen will always yield carbon dioxide and water.



Figure 2.9.2: An Example of a Combustion Reaction. The wax in a candle is a high-molecular-mass [hydrocarbon](#), which produces gaseous carbon dioxide and water vapor in a combustion reaction (Equation 2.9.3).

Equation 2.9.3 is not balanced: the numbers of each type of atom on the reactant side of the equation (7 carbon atoms, 16 hydrogen atoms, and 2 oxygen atoms) is not the same as the numbers of each type of atom on the product side (1 carbon atom, 2 hydrogen atoms, and 3 oxygen atoms). Consequently, the coefficients of the reactants and products must be adjusted to give the same numbers of atoms of each type on both sides of the equation. Because the identities of the reactants and products are fixed, the equation cannot be balanced by changing the subscripts of the reactants or the products. To do so would change the chemical identity of the species being described, as illustrated in Figure 2.9.3.

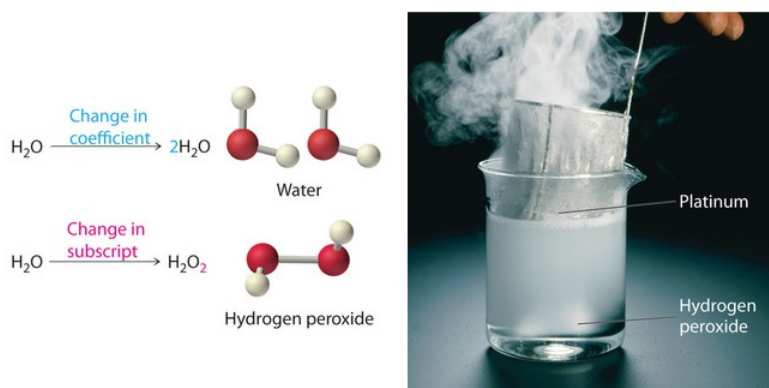
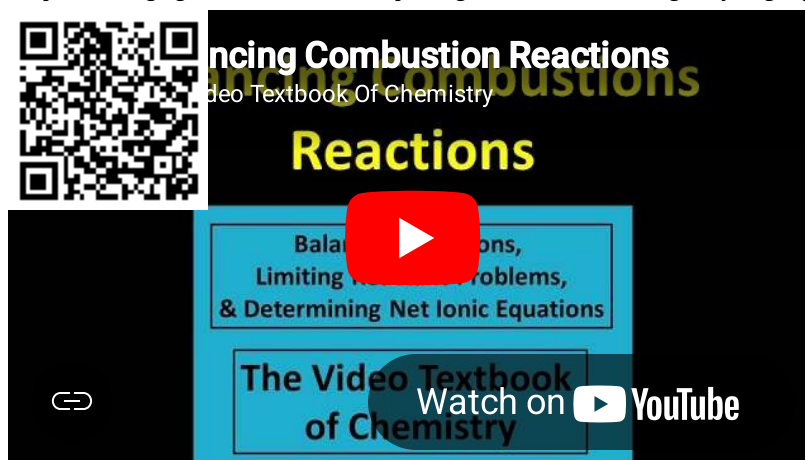


Figure 2.9.3: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances. H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum, whereas no such reaction occurs between water and platinum.

Left: Example of changing coefficients or subscripts. Right: Platinum dissolving in hydrogen peroxide.



Balancing Combustion Reactions: [Balancing Combustions Reactions, YouTube\(opens in new window\)](https://www.youtube.com/watch?v=466746) [youtu.be]

The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Balancing Chemical Equations

1. Identify the most complex substance.

2. Beginning with that substance, choose an element that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element on both sides.
3. Balance polyatomic ions (if present) as a unit.
4. Balance the remaining atoms, usually ending with the least complex substance and using fractional coefficients if necessary. If a fractional coefficient has been used, multiply both sides of the equation by the denominator to obtain whole numbers for the coefficients.
5. Check your work by counting the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

✓ Example 2.9.1A: Combustion of Heptane

To demonstrate this approach, let's use the combustion of n-heptane (Equation 2.9.3) as an example.

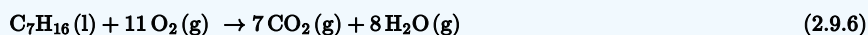
1. **Identify the most complex substance.** The most complex substance is the one with the largest number of different atoms, which is C_7H_{16} . We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. **Adjust the coefficients.** Try to adjust the coefficients of the molecules on the other side of the equation to obtain the same numbers of atoms on both sides. Because one molecule of n-heptane contains 7 carbon atoms, we need 7 CO_2 molecules, each of which contains 1 carbon atom, on the right side:



3. **Balance polyatomic ions as a unit.** There are no polyatomic ions to be considered in this reaction.
4. **Balance the remaining atoms.** Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which contains 2 hydrogen atoms, on the right side:



The carbon and hydrogen atoms are now balanced, but we have 22 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex substance, O_2 , on the reactant side:



5. **Check your work.** The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen atoms, and 22 oxygen atoms on each side. Always check to be sure that a chemical equation is balanced. The assumption that the final balanced chemical equation contains only one molecule or formula unit of the most complex substance is not always valid, but it is a good place to start.

✓ Example 2.9.1B: Combustion of Isooctane

Consider, for example, a similar reaction, the combustion of isooctane (C_8H_{18}). Because the combustion of any hydrocarbon with oxygen produces carbon dioxide and water, the unbalanced chemical equation is as follows:



1. **Identify the most complex substance.** Begin the balancing process by assuming that the final balanced chemical equation contains a single molecule of isooctane.
2. **Adjust the coefficients.** The first element that appears only once in the reactants is carbon: 8 carbon atoms in isooctane means that there must be 8 CO_2 molecules in the products:



3. **Balance polyatomic ions as a unit.** This step does not apply to this equation.

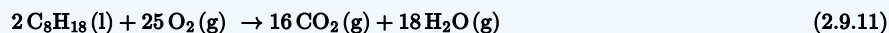
4. **Balance the remaining atoms.** Eighteen hydrogen atoms in isooctane means that there must be 9 H₂O molecules in the products:



The carbon and hydrogen atoms are now balanced, but we have 25 oxygen atoms on the right side and only 2 oxygen atoms on the left. We can balance the least complex substance, O₂, but because there are 2 oxygen atoms per O₂ molecule, we must use a fractional coefficient (25/2) to balance the oxygen atoms:



Equation 2.9.10 is now balanced, but we usually write equations with whole-number coefficients. We can eliminate the fractional coefficient by multiplying all coefficients on both sides of the chemical equation by 2:



5. **Check your work.** The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side.



Balancing Complex Chemical Equations: [Balancing Complex Chemical Equations, YouTube\(opens in new window\)](#) [youtu.be]

Balancing equations requires some practice on your part as well as some common sense. If you find yourself using very large coefficients or if you have spent several minutes without success, go back and make sure that you have written the formulas of the reactants and products correctly.

✓ Example 2.9.1C: Hydroxyapatite

The reaction of the mineral hydroxyapatite ($\text{Ca}_5(\text{PO}_4)_3(\text{OH})$) with phosphoric acid and water gives $\text{Ca}(\text{H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}$ (calcium dihydrogen phosphate monohydrate). Write and balance the equation for this reaction.



Hydroxyapatite (Extra close brace or missing open brace) crystal

Given: reactants and product

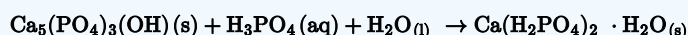
Asked for: balanced chemical equation

Strategy:

- A. Identify the product and the reactants and then write the unbalanced chemical equation.
- B. Follow the steps for balancing a chemical equation.

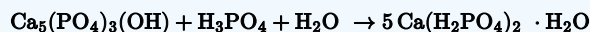
Solution:

A We must first identify the product and reactants and write an equation for the reaction. The formulas for hydroxyapatite and calcium dihydrogen phosphate monohydrate are given in the problem (recall that phosphoric acid is H_3PO_4). The initial (unbalanced) equation is as follows:

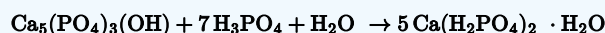


1. **B Identify the most complex substance.** We start by assuming that only one molecule or formula unit of the most complex substance, $\text{Ca}_5(\text{PO}_4)_3(\text{OH})$, appears in the balanced chemical equation.

2. **Adjust the coefficients.** Because calcium is present in only one reactant and one product, we begin with it. One formula unit of $\text{Ca}_5(\text{PO}_4)_3(\text{OH})$ contains 5 calcium atoms, so we need 5 $\text{Ca}(\text{H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}$ on the right side:

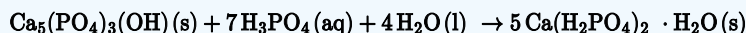


3. **Balance polyatomic ions as a unit.** It is usually easier to balance an equation if we recognize that certain combinations of atoms occur on both sides. In this equation, the polyatomic phosphate ion (PO_4^{3-}), shows up in three places. In H_3PO_4 , the phosphate ion is combined with three H^+ ions to make phosphoric acid (H_3PO_4), whereas in $\text{Ca}(\text{H}_2\text{PO}_4)_2 \cdot \text{H}_2\text{O}$ it is combined with two H^+ ions to give the dihydrogen phosphate ion. Thus it is easier to balance PO_4 as a unit rather than counting individual phosphorus and oxygen atoms. There are 10 PO_4 units on the right side but only 4 on the left. The simplest way to balance the PO_4 units is to place a coefficient of 7 in front of H_3PO_4 :



Although OH^- is also a polyatomic ion, it does not appear on both sides of the equation. So oxygen and hydrogen must be balanced separately.

4. **Balance the remaining atoms.** We now have 30 hydrogen atoms on the right side but only 24 on the left. We can balance the hydrogen atoms using the least complex substance, H_2O , by placing a coefficient of 4 in front of H_2O on the left side, giving a total of 4 H_2O molecules:



The equation is now balanced. Even though we have not explicitly balanced the oxygen atoms, there are 41 oxygen atoms on each side.

5. **Check your work.** Both sides of the equation contain 5 calcium atoms, 10 phosphorus atoms, 30 hydrogen atoms, and 41 oxygen atoms.

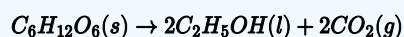
? Exercise 2.9.1: Fermentation

Fermentation is a biochemical process that enables yeast cells to live in the absence of oxygen. Humans have exploited it for centuries to produce wine and beer and make bread rise. In fermentation, sugars such as glucose are converted to ethanol ($\text{CH}_3\text{CH}_2\text{OH}$) and carbon dioxide CO_2 . Write a balanced chemical reaction for the fermentation of glucose.



Commercial use of fermentation. (a) Microbrewery vats are used to prepare beer. (b) The fermentation of glucose by yeast cells is the reaction that makes beer production possible.

Answer



Balancing Reactions Which Contain Polyatomics: [Balancing Reactions Which Contain Polyatomics, YouTube](https://www.youtube.com/watch?v=466746)(opens in new window) [youtu.be]

Interpreting Chemical Equations

In addition to providing qualitative information about the identities and physical states of the reactants and products, a balanced chemical equation provides quantitative information. Specifically, it gives the relative amounts of reactants and products consumed or produced in a reaction. The number of atoms, molecules, or formula units of a reactant or a product in a balanced chemical

equation is the coefficient of that species (e.g., the 4 preceding H_2O in Equation 2.9.1). When no coefficient is written in front of a species, the coefficient is assumed to be 1. As illustrated in Figure 2.9.4, the coefficients allow Equation 2.9.1 to be interpreted in any of the following ways:

- Two NH_4^+ ions and one $\text{Cr}_2\text{O}_7^{2-}$ ion yield 1 formula unit of Cr_2O_3 , 1 N_2 molecule, and 4 H_2O molecules.
- One mole of $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ yields 1 mol of Cr_2O_3 , 1 mol of N_2 , and 4 mol of H_2O .
- A mass of 252 g of $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ yields 152 g of Cr_2O_3 , 28 g of N_2 , and 72 g of H_2O .
- A total of 6.022×10^{23} formula units of $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ yields 6.022×10^{23} formula units of Cr_2O_3 , 6.022×10^{23} molecules of N_2 , and 24.09×10^{23} molecules of H_2O .

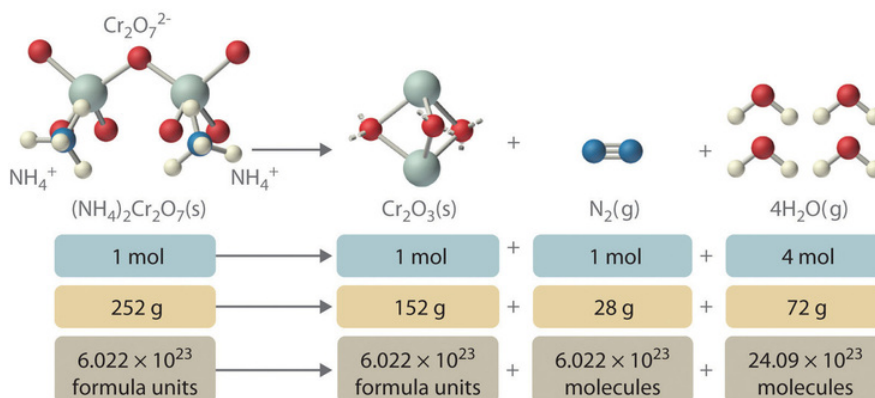


Figure 2.9.4: The Relationships among Moles, Masses, and Formula Units of Compounds in the Balanced Chemical Reaction for the Ammonium Dichromate Volcano

Chemical equation: $(\text{NH}_4)_2\text{Cr}_2\text{O}_7$ dissociates into Cr_2O_3 , N_2 , and H_2O . Conversions are given between moles, mass, and molecules.

These are all chemically equivalent ways of stating the information given in the balanced chemical equation, using the concepts of the mole, molar or formula mass, and Avogadro's number. The ratio of the number of moles of one substance to the number of moles of another is called the mole ratio. For example, the mole ratio of H_2O to N_2 in Equation 2.9.1 is 4:1. The total mass of reactants equals the total mass of products, as predicted by Dalton's law of conservation of mass:

$$\begin{aligned}
 &252 \text{ g of } (\text{NH}_4)_2\text{Cr}_2\text{O}_7 \\
 &\quad \text{yield} \\
 &152 + 28 + 72 = 252 \text{ g of products.}
 \end{aligned}$$

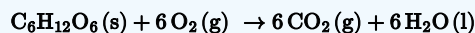
The chemical equation does not, however, show the rate of the reaction (rapidly, slowly, or not at all) or whether energy in the form of heat or light is given off.

An important chemical reaction was analyzed by Antoine Lavoisier, an 18th-century French chemist, who was interested in the chemistry of living organisms as well as simple chemical systems. In a classic series of experiments, he measured the carbon dioxide and heat produced by a guinea pig during respiration, in which organic compounds are used as fuel to produce energy, carbon dioxide, and water. Lavoisier found that the ratio of heat produced to carbon dioxide exhaled was similar to the ratio observed for the reaction of charcoal with oxygen in the air to produce carbon dioxide—a process chemists call combustion. Based on these experiments, he proposed that "Respiration is a combustion, slow it is true, but otherwise perfectly similar to that of charcoal." Lavoisier was correct, although the organic compounds consumed in respiration are substantially different from those

found in charcoal. One of the most important fuels in the human body is glucose ($C_6H_{12}O_6$), which is virtually the only fuel used in the brain. Thus combustion and respiration are examples of chemical reactions.

✓ Example 2.9.2: Combustion of Glucose

The balanced chemical equation for the combustion of [glucose](#) in the laboratory (or in the brain) is as follows:



Construct a table showing how to interpret the information in this equation in terms of

- a single molecule of glucose.
- moles of reactants and products.
- grams of reactants and products represented by 1 mol of glucose.
- numbers of molecules of reactants and products represented by 1 mol of glucose.



The combustion of a sugar cube consisting of [sucrose](#) with a similar reaction to the combustion of glucose. from Wikipedia.

Given: balanced chemical equation

Asked for: molecule, mole, and mass relationships

Strategy:

- Use the coefficients from the balanced chemical equation to determine both the molecular and mole ratios.
- Use the molar masses of the reactants and products to convert from moles to grams.
- Use Avogadro's number to convert from moles to the number of molecules.

Solution:

This equation is balanced as written: each side has 6 carbon atoms, 18 oxygen atoms, and 12 hydrogen atoms. We can therefore use the coefficients directly to obtain the desired information.

- One molecule of glucose reacts with 6 molecules of O_2 to yield 6 molecules of CO_2 and 6 molecules of H_2O .
- One mole of glucose reacts with 6 mol of O_2 to yield 6 mol of CO_2 and 6 mol of H_2O .
- To interpret the equation in terms of masses of reactants and products, we need their molar masses and the mole ratios from part b. The molar masses in grams per mole are as follows: glucose, 180.16; O_2 , 31.9988; CO_2 , 44.010; and H_2O , 18.015.

mass of reactants = mass of products

$$g \text{ glucose} + g O_2 = g CO_2 + g H_2O$$

$$1 \text{ mol glucose} \left(\frac{180.16 \text{ g}}{1 \text{ mol glucose}} \right) + 6 \text{ mol } O_2 \left(\frac{31.9988 \text{ g}}{1 \text{ mol } O_2} \right)$$

$$= 6 \text{ mol } CO_2 \left(\frac{44.010 \text{ g}}{1 \text{ mol } CO_2} \right) + 6 \text{ mol } H_2O \left(\frac{18.015 \text{ g}}{1 \text{ mol } H_2O} \right)$$

$$372.15 \text{ g} = 372.15 \text{ g}$$

C One mole of glucose contains Avogadro's number (6.022×10^{23}) of glucose molecules. Thus 6.022×10^{23} glucose molecules react with $(6 \times 6.022 \times 10^{23}) = 3.613 \times 10^{24}$ oxygen molecules to yield $(6 \times 6.022 \times 10^{23}) = 3.613 \times 10^{24}$ molecules each of CO_2 and H_2O .

In tabular form:

Solution to Example 3.1.2

	$\text{C}_6\text{H}_{12}\text{O}_6 (s)$	+	$6\text{O}_2 (g)$	→	$6\text{CO}_2 (g)$		$6\text{H}_2\text{O} (l)$
a.	1 molecule		6 molecules		6 molecules		6 molecules
b.	1 mol		6 mol		6 mol		6 mol
c.	180.16 g		191.9928 g		264.06 g		108.09 g
d.	6.022×10^{23} molecules		3.613×10^{24} molecules		3.613×10^{24} molecules		3.613×10^{24} molecule

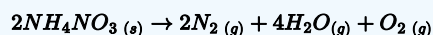
? Exercise 2.9.2: Ammonium Nitrate Explosion

Ammonium nitrate is a common fertilizer, but under the wrong conditions, it can be hazardous. In 1947, a ship loaded with ammonium nitrate caught fire during unloading and exploded, destroying the town of Texas City, Texas.



Ammonium nitrate can be hazardous. This aerial photograph of Texas City, Texas, shows the devastation caused by the explosion of a shipload of ammonium nitrate on April 16, 1947. For a video click [here](#).

The explosion resulted from the following reaction:



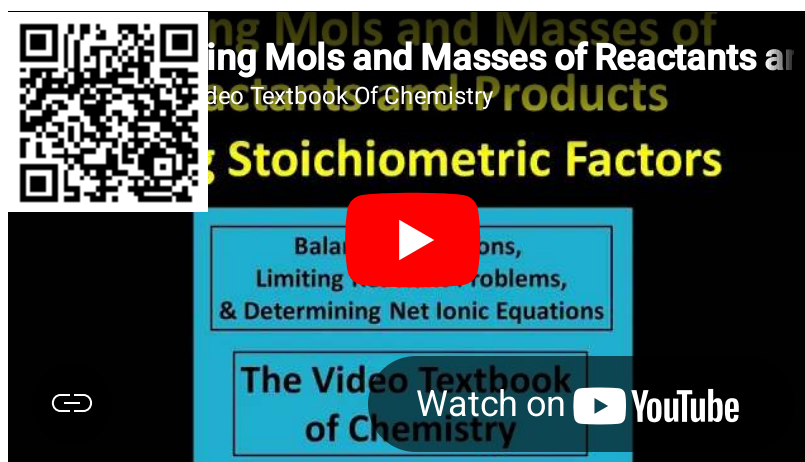
Construct a table showing how to interpret the information in the equation in terms of

- individual molecules and ions.
- moles of reactants and products.
- grams of reactants and products given 2 mol of ammonium nitrate.
- numbers of molecules or formula units of reactants and products given 2 mol of ammonium nitrate.

Answer:

Answer to Exercise 3.1.2

	$2\text{NH}_4\text{NO}_3 (\text{s})$	$\rightarrow$	$2\text{N}_2 (\text{g})$	$+$	$4\text{H}_2\text{O} (\text{g})$	$+$	$\text{O}_2 (\text{g})$
a.	2NH_4^+ ions and 2NO_3^- ions		2 molecules		4 molecules		1 molecule
b.	2 mol		2 mol		4 mol		1 mol
c.	160.0864 g		56.0268 g		72.0608 g		31.9988 g
d.	1.204×10^{24} formula units		1.204×10^{24} molecules		2.409×10^{24} molecules		6.022×10^{23} molecules



Finding Mols and Masses of Reactants and Products Using Stoichiometric Factors (Mol Ratios): [Finding Mols and Masses of Reactants and Products Using Stoichiometric Factors, YouTube\(opens in new window\)](#) [youtu.be]

Summary

A chemical reaction is described by a chemical equation that gives the identities and quantities of the reactants and the products. In a chemical reaction, one or more substances are transformed to new substances. A chemical reaction is described by a **chemical equation**, an expression that gives the identities and quantities of the substances involved in a reaction. A chemical equation shows the starting compound(s)—the **reactants**—on the left and the final compound(s)—the **products**—on the right, separated by an arrow. In a balanced chemical equation, the numbers of atoms of each element and the total charge are the same on both sides of the equation. The number of atoms, molecules, or formula units of a reactant or product in a balanced chemical equation is the **coefficient** of that species. The **mole ratio** of two substances in a chemical reaction is the ratio of their coefficients in the balanced chemical equation.

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CHAPTER OVERVIEW

3: Gases and Climate Change

3.1: The Kinetic-Molecular Theory of Gases

3.2: The Simple Gas Laws

3.3: Earth's Atmosphere- Divisions and Composition

3.4: Outdoor Air Pollution

3.5: Greenhouse Gases

3.6: Stratospheric Ozone- Earth's Vital Shield

3.7: Consequences of Climate Change

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3.1: The Kinetic-Molecular Theory of Gases

Learning Objectives

- State the major concepts behind the kinetic theory of gases.
- Relate the general properties of gases to the kinetic theory.

Gases were among the first substances studied in terms of the modern scientific method, which was developed in the 1600s. It did not take long to recognize that gases all shared certain physical behaviors, suggesting that all gases could be described by one all-encompassing theory. Today, that theory is the **kinetic theory of gases**. It is based on the following statements:

1. Gases consist of tiny particles of matter that are in constant motion.
2. Gas particles are constantly colliding with each other and the walls of a container. These collisions are elastic; that is, there is no net loss of energy from the collisions.
3. Gas particles are separated by large distances, with the size of a gas particle tiny compared to the distances that separate them.
4. There are no interactive forces (i.e., attraction or repulsion) between the particles of a gas.
5. The average speed of gas particles is dependent on the temperature of the gas (in Kelvin).

Figure 3.1.1 shows a representation of how we mentally picture the gas phase.

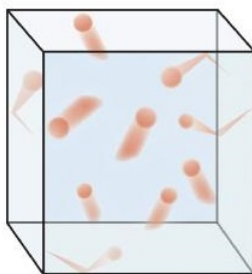


Figure 3.1.1 The kinetic-molecular theory of gases. The kinetic theory of gases describes this state of matter as composed of tiny particles in constant motion with a lot of distance between the particles.

This model of gases explains some of the physical properties of gases. Because most of a gas is empty space, a gas has a low density and can expand or contract under the appropriate influence. The fact that gas particles are in constant motion means that two or more gases will always mix, as the particles from the individual gases move and collide with each other.

An **ideal gas** is a gas that exactly follows the statements of the kinetic theory. Unfortunately, *real gases* are not ideal. Many gases deviate slightly from agreeing perfectly with the kinetic theory of gases. However, most gases adhere to the statements so well that the kinetic theory of gases is well accepted by the scientific community.

- The physical behavior of gases is explained by the kinetic theory of gases.
- An ideal gas adheres exactly to the kinetic theory of gases.

Contributors and Attributions

- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Erin Avram (Cleveland State University)

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3.2: The Simple Gas Laws

Learning Objectives

- Know the different units of pressure.
- Explain the relationship between pressure and volume of a gas and use Boyle's law to calculate the unknown pressure or volume.
- Explain the relationship between temperature and volume of a gas and use Charles's law to calculate the unknown temperature or volume.
- Explain the relationship between pressure and temperature of a gas and use Gay-Lussac's law to calculate the unknown pressure or temperature.

The behavior of gases can be modeled and predicted with gas laws. Boyle's law relates a gas's pressure and volume at constant temperature and amount. Charles's law relates a gas's volume and temperature at constant pressure and amount. In gas laws, temperatures must always be expressed in Kelvin.

Pressure Units

A barometer measures gas pressure by the height of the column of mercury. One unit of gas pressure is the millimeter of mercury (**mm Hg**). An equivalent unit to the **mm Hg** is called the **torr**, in honor of the inventor of the barometer, Evangelista Torricelli. The **pascal (Pa)** is the standard unit of pressure. A pascal is a very small amount of pressure, so the most useful unit for everyday gas pressures is the kilopascal (**kPa**). A kilopascal is equal to 1000 pascals. Another commonly used unit of pressure is the **atmosphere (atm)**. Standard atmospheric pressure is called **1 atm** of pressure and is equal to **760 mm Hg** and **101.3 kPa**. Atmospheric pressure is also often stated as pounds per square inch (**psi**). The atmospheric pressure at sea level is **14.7 psi**.

$$1 \text{ atm} = 760 \text{ mm Hg} = 760 \text{ torr} = 101.3 \text{ kPa} = 14.7 \text{ psi}$$

Boyle's Law: Pressure and Volume

When seventeenth-century scientists began studying the physical properties of gases, they noticed some simple relationships between some of the measurable properties of the gas. Take pressure (P) and volume (V), for example. Scientists noted that for a given amount of a gas (usually expressed in units of moles [n]), if the temperature (T) of the gas was kept constant, pressure and volume were related: As one increases, the other decreases. As one decreases, the other increases. We say that pressure and volume are *inversely related*.

There is more to it, however: pressure and volume of a given amount of gas at constant temperature are *numerically* related. If you take the pressure value and multiply it by the volume value, the product is a constant for a given amount of gas at a constant temperature:

$$P \times V = \text{constant at constant } n \text{ and } T$$

If either volume or pressure changes while amount and temperature stay the same, then the other property must change so that the product of the two properties still equals that same constant. That is, if the original conditions are labeled P_1 and V_1 and the new conditions are labeled P_2 and V_2 , we have

$$P_1 V_1 = \text{constant} = P_2 V_2$$

where the properties are assumed to be multiplied together. Leaving out the middle part, we have simply

$$P_1 V_1 = P_2 V_2 \text{ at constant } n \text{ and } T$$

This equation is an example of a gas law. A **gas law** is a simple mathematical formula that allows you to model, or predict, the behavior of a gas. This particular gas law is called **Boyle's law**, after the English scientist Robert Boyle, who first announced it in

1662. Figure 3.2.1 shows two representations of how Boyle's law works.

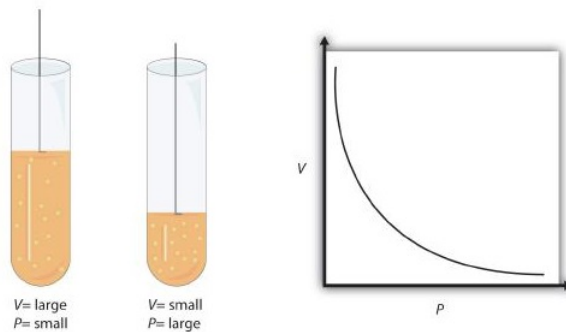


Figure 3.2.1 Boyle's Law. A piston having a certain pressure and volume (left piston) will have half the volume when its pressure is twice as much (right piston). One can also plot P versus V for a given amount of gas at a certain temperature; such a plot will look like the graph on the right.

Boyle's law is an example of a second type of mathematical problem we see in chemistry—one based on a mathematical formula. Tactics for working with mathematical formulas are different from tactics for working with conversion factors. First, most of the questions you will have to answer using formulas are word-type questions, so the first step is to identify what quantities are known and assign them to variables. Second, in most formulas, some mathematical rearrangements (i.e., algebra) must be performed to solve for an unknown variable. The rule is that to find the value of the unknown variable, you must mathematically isolate the unknown variable *by itself and in the numerator* of one side of the equation. Finally, units must be consistent. For example, in Boyle's law there are two pressure variables; they *must* have the same unit. There are also two volume variables; they also must have the same unit. In most cases, it won't matter *what* the unit is, but the unit must be the *same* on both sides of the equation.

✓ Example 3.2.1

A sample of gas has an initial pressure of 2.44 atm and an initial volume of 4.01 L. Its pressure changes to 1.93 atm. What is the new volume if temperature and amount are kept constant?

Solution

Example 3.2.1: Steps for Problem Solving volume and pressure changes.

Steps for Problem Solving	Example 3.2.1
Identify the "given" information and what the problem is asking you to "find."	Given: $P_1 = 2.44$ atm and $V_1 = 4.01$ L $P_2 = 1.93$ atm Find: $V_2 = ?$ L
List other known quantities.	none
Plan the problem.	First, rearrange the equation algebraically to solve for V_2 . $V_2 = \frac{P_1 \times V_1}{P_2}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $V_2 = \frac{2.44 \cancel{\text{ atm}} \times 4.01 \text{ L}}{1.93 \cancel{\text{ atm}}} = 5.07 \text{ L}$
Think about your result.	We know that pressure and volume are inversely related; as one decreases, the other increases. Pressure is decreasing (from 2.44 atm to 1.93 atm), so volume should be increasing to compensate, and it is (from 4.01 L to 5.07 L). So the answer makes sense based on Boyle's law.

? Exercise 3.2.1

If $P_1 = 334$ torr, $V_1 = 37.8$ mL, and $P_2 = 102$ torr, what is V_2 ?

Answer

124 mL

As mentioned, you can use any unit for pressure or volume, but both pressures must be expressed in the same units, and both volumes must be expressed in the same units.

✓ Example 3.2.2:

A sample of gas has an initial pressure of 722 torr and an initial volume of 88.8 mL. Its volume changes to 0.663 L. What is the new pressure?

Solution

Example 3.2.2: Steps for Problem Solving volume and pressure changes

Steps for Problem Solving	Example 3.2.2
Identify the "given" information and what the problem is asking you to "find."	Given: $P_1 = 722$ torr and $V_1 = 88.8$ mL $V_2 = 0.663$ L Find: $P_2 = ?$ torr
List other known quantities.	1 L = 1000 mL to have the same units for volume.
Plan the problem.	1. Perform the conversion of the second volume unit from L to mL. 2. Rearrange the equation algebraically to solve for P_2 . $P_2 = \frac{P_1 \times V_1}{V_2}$
Cancel units and calculate.	1. $0.663 \cancel{\text{L}} \times \frac{1000 \text{ mL}}{1 \cancel{\text{L}}} = 663 \text{ mL}$ 2. Substitute the known quantities into the equation and solve. $P_2 = \frac{722 \text{ torr} \times 88.8 \cancel{\text{mL}}}{663 \cancel{\text{mL}}} = 96.7 \text{ torr}$
Think about your result.	When the volume increased, the pressure decreased, which is as expected for Boyle's law.

? Exercise 3.2.2

If $V_1 = 456$ mL, $P_1 = 308$ torr, and $P_2 = 1.55$ atm, what is V_2 ?

Answer

119 mL

Charles's Law: Temperature and Volume

French physicist Jacques Charles (1746 - 1823) studied the effect of temperature on the volume of a gas at constant pressure. **Charles's Law** states that the volume of a given mass of gas varies directly with the absolute temperature of the gas when pressure is kept constant. The absolute temperature is the temperature measured with the Kelvin scale. The Kelvin scale must be used because zero on the Kelvin scale corresponds to a complete stoppage of molecular motion.

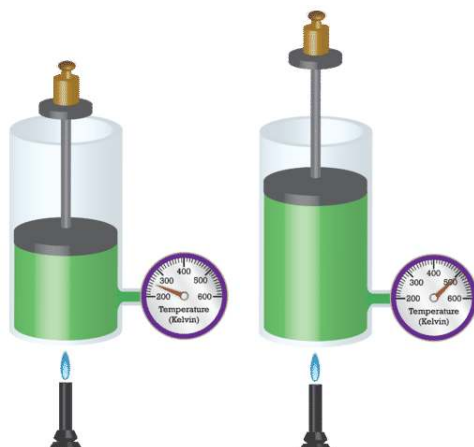


Figure 3.2.2 As a container of confined gas is heated, its molecules increase in kinetic energy and push the movable piston outward, resulting in an increase in volume.

Mathematically, the direct relationship of Charles's Law can be represented by the following equation:

$$\frac{V}{T} = k$$

As with Boyle's Law, k is constant only for a given gas sample. The table below shows temperature and volume data for a set amount of gas at a constant pressure. The third column is the constant for this particular data set and is always equal to the volume divided by the Kelvin temperature.

Table 3.2.1 Temperature-Volume Data

Temperature (K)	Volume (mL)	$\frac{V}{T} = k \left(\frac{\text{mL}}{\text{K}} \right)$
50	20	0.40
100	40	0.40
150	60	0.40
200	80	0.40
300	120	0.40
500	200	0.40
1000	400	0.40

When this data is graphed, the result is a straight line, indicative of a direct relationship, shown in the figure below.

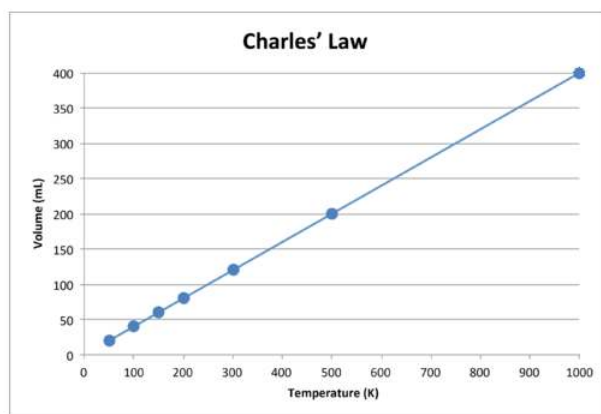


Figure 3.2.3 graph plots how the volume of a gas increases as the Kelvin temperature increases. The increase is linear.

Notice that as temperature increases, volume increases, indicating a direct relationship between the volume and temperature of a gas. We can also see that the line goes exactly toward the origin, meaning that as the absolute temperature of the gas approaches zero, its volume approaches zero. However, when a gas is brought to extremely cold temperatures, its molecules would eventually condense into a liquid state before reaching absolute zero. The temperature at which this change into the liquid state occurs varies for different gases.

Charles's Law can also be used to compare changing conditions for a gas. Now we use V_1 and T_1 to stand for the initial volume and temperature of a gas, while V_2 and T_2 stand for the final volume and temperature. The mathematical relationship of Charles's Law becomes:

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

This equation can be used to calculate any one of the four quantities if the other three are known. The direct relationship will only hold if the temperatures are expressed in Kelvin. Temperatures in Celsius will not work. Recall the relationship that $K = ^\circ C + 273$.

✓ Example 3.2.3:

A balloon is filled to a volume of **2.20 L** at a temperature of **22°C**. The balloon is then heated to a temperature of **71°C**. Find the new volume of the balloon.

Solution

Steps for Problem Solving	Example 3.2.3
Identify the "given" information and what the problem is asking you to "find."	Given: $V_1 = 2.20 \text{ L}$ and $T_1 = 22^\circ\text{C} = 295 \text{ K}$ $T_2 = 71^\circ\text{C} = 344 \text{ K}$ Find: $V_2 = ? \text{ L}$
List other known quantities.	The temperatures have first been converted to Kelvin.
Plan the problem.	First, rearrange the equation algebraically to solve for V_2 . $V_2 = \frac{V_1 \times T_2}{T_1}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $V_2 = \frac{2.20 \text{ L} \times 344 \cancel{\text{K}}}{295 \cancel{\text{K}}} = 2.57 \text{ L}$
Think about your result.	The volume increases as the temperature increases. The result has three significant figures.

? Exercise 3.2.3

If $V_1 = 3.77 \text{ L}$ and $T_1 = 255 \text{ K}$, what is V_2 if $T_2 = 123 \text{ K}$?

Answer

1.82 L

✓ Example 3.2.4:

A sample of a gas has an initial volume of 34.8 L and an initial temperature of -67°C . What must be the temperature of the gas for its volume to be 25.0 L?

Solution

Example 3.2.4: Steps for Problem Solving Volume and Temperature Changes

Steps for Problem Solving

Example 3.2.4

Identify the "given" information and what the problem is asking you to "find."

Given:

Given: $T_1 = -27^\circ\text{C}$ and $V_1 = 34.8\text{ L}$

$V_2 = 25.0\text{ L}$

Find: $T_2 = ?\text{ K}$

List other known quantities.

$\text{K} = -27^\circ\text{C} + 273$

Plan the problem.

1. Convert the initial temperature to Kelvin

2. Rearrange the equation algebraically to solve for T_2 .

$$T_2 = \frac{V_2 \times T_1}{V_1}$$

Cancel units and calculate.

1. $-27^\circ\text{C} + 273 = 246\text{ K}$

2. Substitute the known quantities into the equation and solve.

$$T_2 = \frac{25.0\cancel{\text{ L}} \times 246\text{ K}}{34.8\cancel{\text{ L}}} = 178\text{ K}$$

Think about your result.

This is also equal to -125°C . As temperature decreases, volume decreases, which it does in this example.

? Exercise 3.2.4

If $V_1 = 623\text{ mL}$, $T_1 = 255^\circ\text{C}$, and $V_2 = 277\text{ mL}$, what is T_2 ?

Answer

235 K, or -38°

Gay-Lussac's Law: Temperature and Pressure

When the temperature of a sample of gas in a rigid container is increased, the pressure of the gas increases as well. The increase in kinetic energy results in the molecules of gas striking the walls of the container with more force, resulting in a greater pressure. The French chemist Joseph Gay-Lussac (1778 - 1850) discovered the relationship between the pressure of a gas and its absolute temperature. **Gay-Lussac's Law** states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant. Gay-Lussac's Law is very similar to Charles's Law, with the only difference being the type of container. Whereas the container in a Charles's Law experiment is flexible, it is rigid in a Gay-Lussac's Law experiment.



Figure 3.2.4 Joseph Gay-Lussac.

The mathematical expressions for Gay-Lussac's Law are likewise similar to those of Charles's Law:

$$\frac{P}{T} \quad \text{and} \quad \frac{P_1}{T_1} = \frac{P_2}{T_2}$$

A graph of pressure vs. temperature also illustrates a direct relationship. As a gas is cooled at constant volume its pressure continually decreases until the gas condenses to a liquid.

✓ Example 3.2.5

The gas in an aerosol can is under a pressure of **3.00 atm** at a temperature of **25°C**. It is dangerous to dispose of an aerosol can by incineration. What would the pressure in the aerosol can be at a temperature of **845°C**?

Solution

Example 3.2.5: Steps for Problem Solving Volume and Tempature Changes

Steps for Problem Solving	Example 3.2.5
Identify the "given" information and what the problem is asking you to "find."	Given: $P_1 = 3.00 \text{ atm}$ $T_1 = 25^\circ\text{C} = 298 \text{ K}$ $T_2 = 845^\circ\text{C} = 1118 \text{ K}$ Find: $P_2 = ? \text{ atm}$
List other known quantities	The temperatures have first been converted to Kelvin.
Plan the problem	First, rearrange the equation algebraically to solve for P_2 . $P_2 = \frac{P_1 \times T_2}{T_1}$
Calculate	Now substitute the known quantities into the equation and solve. $P_2 = \frac{3.00 \text{ atm} \times 1118 \text{ K}}{298 \text{ K}} = 11.3 \text{ atm}$
Think about your result.	The pressure increases dramatically due to large increase in temperature.

Molar Volume

It should be obvious by now that some physical properties of gases depend strongly on the conditions. What we need is a set of standard conditions so that properties of gases can be properly compared to each other. **Standard Temperature and Pressure (STP)** is defined as exactly 100 kPa of pressure (0.986 atm) and 273 K (0°C). For simplicity, we will use 1 atm as standard pressure. Defining STP allows us to compare more directly the properties of gases that differ from each other.

One property shared among gases is a molar volume. The **molar volume** is the volume of 1 mol of a gas. At STP, the molar volume of a gas can be easily determined by using the ideal gas law:

$$(1 \text{ atm})V = (1 \text{ mol})\left(0.08205 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273 \text{ K})$$

All the units cancel except for L, the unit of volume. So $V = 22.4 \text{ L}$

Note that we have not specified the identity of the gas; we have specified only that the pressure is 1 atm and the temperature is 273 K. This makes for a very useful approximation: *any gas at STP has a volume of 22.4 L per mole of gas*; that is, the molar volume at STP is 22.4 L/mol (Fig. 6.6.1 Molar Volume). This molar volume makes a useful conversion factor in stoichiometry problems if the conditions are at STP. If the conditions are not at STP, a molar volume of 22.4 L/mol is not applicable. However, if the conditions are not at STP, the combined gas law can be used to calculate what the volume of the gas would be if at STP; then the 22.4 L/mol molar volume can be used.

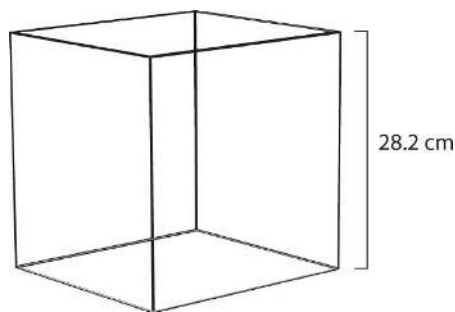


Figure 3.2.5 Molar volume. A mole of gas at STP occupies 22.4 L, the volume of a cube that is 28.2 cm on a side.

✓ Example 3.2.6:

How many moles of Ar are present in 38.7 L at STP?

Solution

We can use the molar volume, 22.4 L/mol, as a conversion factor, but we need to reverse the fraction so that the L units cancel and mol units are introduced. It is a one-step conversion:

$$38.7 \text{ L} \times \frac{1 \text{ mol}}{22.4 \text{ L}} = 1.73 \text{ mol}$$

? Exercise 3.2.5

What volume does 4.87 mol of Kr have at STP?

Figure 3.2.5 Molar volume. A mole of gas at STP occupies 22.4 L, the volume of a cube that is 28.2 cm on a side.

109 L

📌 Note

Chemistry Is Everywhere: Breathing

Breathing (more properly called *respiration*) is the process by which we draw air into our lungs so that our bodies can take up oxygen from the air. Let us apply the gas laws to breathing.

Start by considering pressure. We draw air into our lungs because the diaphragm, a muscle underneath the lungs, moves down to reduce pressure in the lungs, causing external air to rush in to fill the lower-pressure volume. We expel air by the diaphragm pushing against the lungs, increasing pressure inside the lungs and forcing the high-pressure air out. What are the pressure changes involved? A quarter of an atmosphere? A tenth of an atmosphere? Actually, under normal conditions, it's only 1 or 2 torr of pressure difference that makes us breathe in and out.

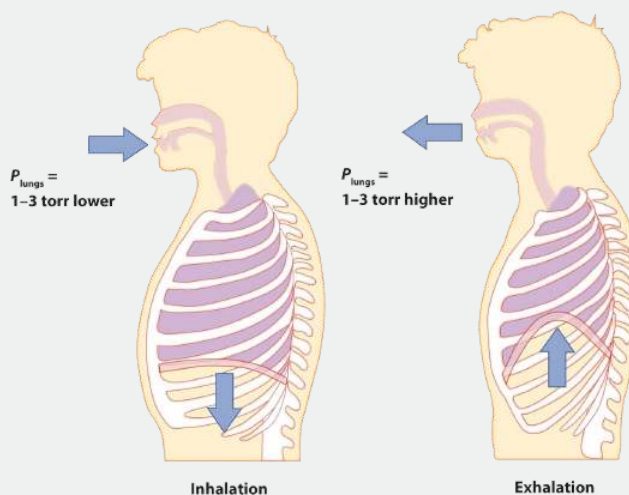


Fig. 6.6.2 Breathing Mechanics. Breathing involves pressure differences between the inside of the lungs and the air outside. The pressure differences are only a few torr.

A normal breath is about 0.50 L. If room temperature is about 22°C, then the air has a temperature of about 295 K. With normal pressure being 1.0 atm, how many moles of air do we take in for every breath? The ideal gas law gives us an answer:

$$(1.0 \text{ atm})(0.50 \text{ L}) = n(0.08205 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(295 \text{ K})$$

Solving for the number of moles, we get

$$n = 0.021 \text{ mol air}$$

This ends up being about 0.6 g of air per breath—not much but enough to keep us alive.

Summary

- The behavior of gases can be modeled with gas laws.
- Boyle's law relates a gas's pressure and volume at constant temperature and amount.
- Charles's law relates a gas's volume and temperature at constant pressure and amount.
- Gay-Lussac's Law states that the pressure of a given mass of gas varies directly with the absolute temperature of the gas, when the volume is kept constant.
- In gas laws, temperatures must always be expressed in Kelvins.

Contributors and Attributions

- CK-12
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Erin Avram (Cleveland State University)

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3.3: Earth's Atmosphere- Divisions and Composition

Learning Objectives

- Describe the different layers of the atmosphere.
- Summarize the current composition of dry air.

The Earth's atmosphere is composed of several layers (Figure 3.3.1). The lowest layer, the **troposphere**, extends from the Earth's surface up to about 6 miles or 10 kilometers (km) in altitude. Virtually all human activities occur in the troposphere. Mt. Everest, the tallest mountain on the planet, is only about 5.6 miles (9 km) high. The next layer, the **stratosphere**, continues from 6 miles (10 km) to about 31 miles (50 km). Most commercial airplanes fly in the lower part of the stratosphere.

Most atmospheric ozone is concentrated in a layer in the stratosphere, about 9 to 18 miles (15 to 30 km) above the Earth's surface (see the figure below). Ozone is a molecule that contains three oxygen atoms. At any given time, ozone molecules are constantly formed and destroyed in the stratosphere. The total amount has remained relatively stable during the decades that it has been measured.

The ozone layer in the stratosphere absorbs a portion of the radiation from the sun, preventing it from reaching the planet's surface. Most importantly, it absorbs the portion of UV light called **UVB**. UVB has been linked to many **harmful effects**, including skin cancers, cataracts, and harm to some crops and marine life.

The mesosphere starts just above the stratosphere and extends to 85 kilometers (53 miles) high. Meteors burn up in this layer. The thermosphere starts just above the mesosphere and extends to 600 kilometers (372 miles) high. Aurora and satellites occur in this layer. The ionosphere is an abundant layer of electrons and ionized atoms and molecules that stretches from about 48 kilometers (30 miles) above the surface to the edge of space at about 965 km (600 mi), overlapping into the mesosphere and thermosphere. This dynamic region grows and shrinks based on solar conditions and divides further into the sub-regions: D, E and F; based on what wavelength of solar vvradiation is absorbed. The ionosphere is a critical link in the chain of Sun-Earth interactions. This region is what makes radio communications possible. The exosphere is the upper limit of the atmosphere. It extends from the top of the thermosphere up to 10,000 km (6,200 mi).

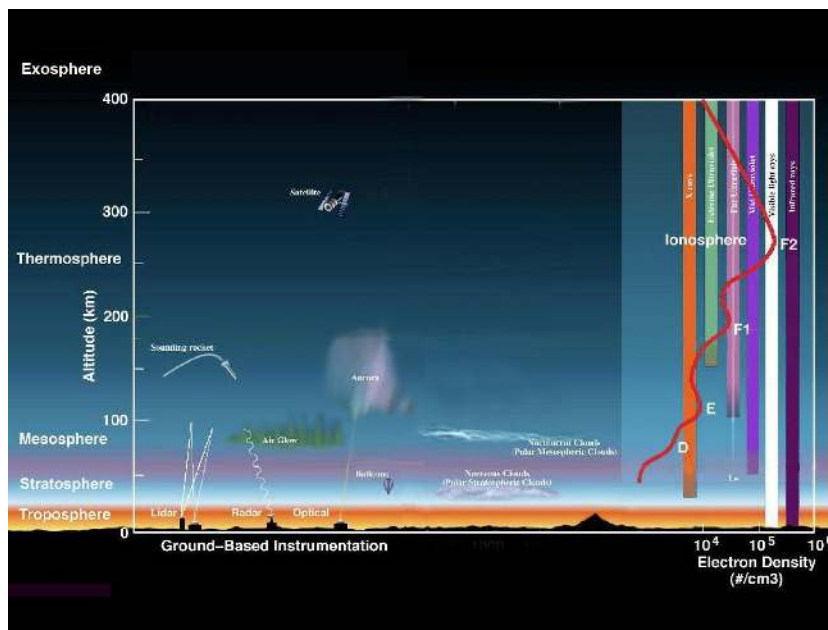


Figure 3.3.1 The Earth's atmosphere *Credit: NASA/Goddard*

Composition of the Atmosphere

Except for water vapor, whose atmospheric abundance varies from practically zero up to 4%, the fractions of the major atmospheric components N_2 , O_2 , and Ar are remarkably uniform below about 100 km (Table 3.3.1). At greater heights, diffusion becomes the principal transport process, and the lighter gases become relatively more abundant. In addition, photochemical processes result in the formation of new species whose high reactivities would preclude their existence in significant concentrations at the higher pressures found at lower elevations.

The atmospheric gases fall into three abundance categories: major, minor, and trace. Nitrogen, the most abundant component, has accumulated over time as a result of its geochemical inertness; a very small fraction of it passes into the other phases as a result of biological activity and natural fixation by lightning. It is believed that denitrifying bacteria in marine sediments may provide the

major route for the return of N_2 to the atmosphere. Oxygen is almost entirely of biological origin, and cycles through the hydrosphere, the biosphere, and sedimentary rocks. Argon consists mainly of ^{40}Ar which is a decay product of ^{40}K in the mantle and crust.

Table 3.3.1 The Major Components of Clean Dry Air in the Lower Atmosphere

nitrogen	N_2	78.08 %
oxygen	O_2	20.95 %
argon	Ar	0.93 %

The most abundant of the minor gases aside from water vapor is carbon dioxide (Table 3.3.2). Next in abundance are neon and helium. Helium is a decay product of radioactive elements in the earth, but neon and the other inert gases are primordial and have probably been present in their current relative abundances since the earth's formation. Two of the minor gases, ozone, and carbon monoxide, have abundances that vary with time and location. A variable abundance implies an imbalance between the rates of formation and removal. In the case of carbon monoxide, whose major source is **anthropogenic** (caused or influenced by people), the variance is probably due largely to localized differences in fuel consumption, particularly in internal combustion engines. The nature of the carbon monoxide sink (removal mechanism) is not entirely clear; it may be partly microbial.

Table 3.3.2 The Minor Components of Air in the Lower Atmosphere.

water	H_2O	0-4 %
carbon dioxide	CO_2	325 ppm
neon	Ne	18 ppm
helium	He	5 ppm
methane	CH_4	2 ppm
krypton	Kr	1 ppm
hydrogen	H_2	0.5 ppm
nitrous oxide	N_2O	0.3 ppm
carbon monoxide	CO	0.05-0.2 ppm
ozone	O_3	0.02 - 10 ppm
xenon	Xe	0.08 ppm

Summary

- The layers of the atmosphere include the troposphere, stratosphere, mesosphere, thermosphere, and exosphere.
- The major components of dry clean air are N_2 , O_2 , and Ar.
- The most abundant of the minor gases are water vapor and CO_2

Contributors and Attributions

- US Environmental Protection Agency (EPA)
- Stephen Lower, Professor Emeritus (Simon Fraser U.) [Chem1 Virtual Textbook](#)
- Marisa Alviar-Agnew (Sacramento City College)
- Erin Avram (Cleveland State University)

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3.4: Outdoor Air Pollution

Learning Objectives

- Define air pollution.
- Know the sources and effects of major outdoor pollutants

Air pollution refers to the introduction, into the atmosphere, of substances that have harmful effects on humans, other living organisms, and the environment either as solid particles, liquid droplets or gases. Air pollution can result from natural processes such as dust storms, forest fires, and volcanic eruptions, or from human activities such as biomass burning, vehicular emissions, mining, agriculture, and industrial processes. Improved technology and government policies have helped reduce most types of outdoor air pollution in many industrialized countries including the United States, in recent decades. However, outdoor air quality is still a problem in less industrialized nations, especially in megacities of rapidly industrializing nations such as China and India.

Outdoor pollutants can come from **stationary (point)** sources or **mobile (nonpoint)** sources (Figure 3.4.1). Stationary sources have a fixed location, for example power plant smokestacks, burning, construction sites, farmlands and surface mines among others. Mobile sources of air pollutants move from place to place while emitting pollutants. Examples of mobile sources include vehicles, aircrafts, ships, and trains.

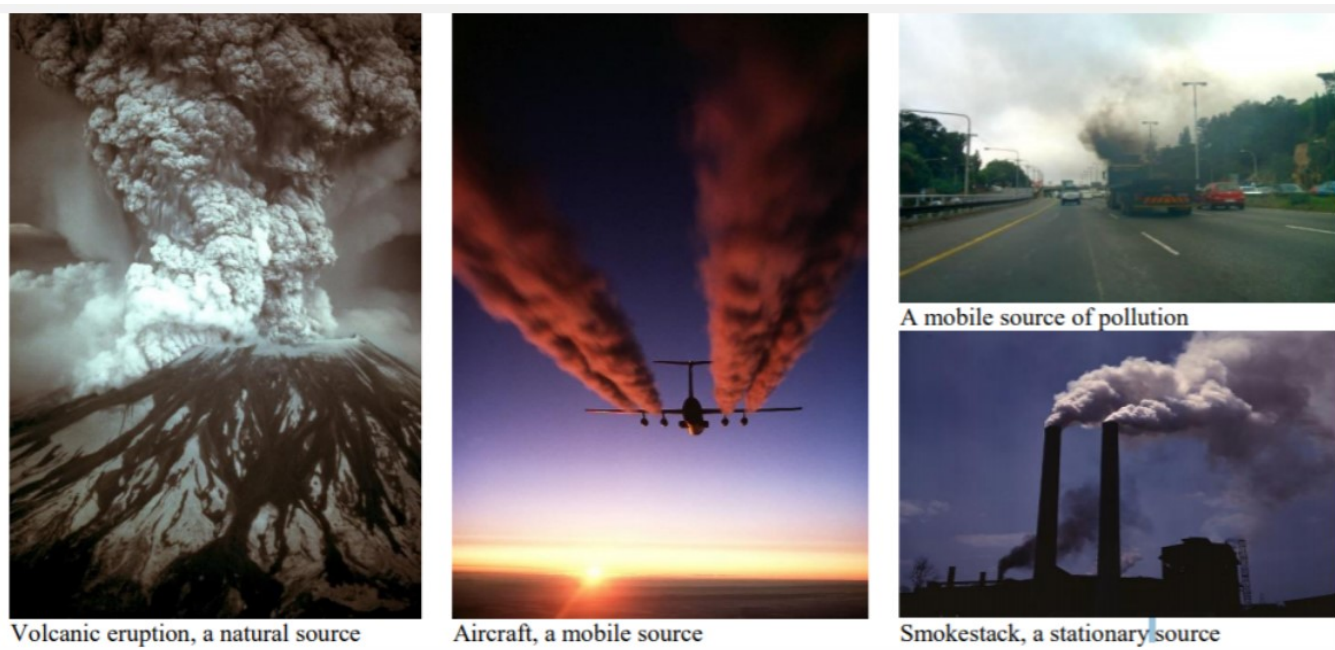


Figure 3.4.1: Images showing various sources of air pollution, including natural and anthropogenic, stationary and mobile. Source: All images obtained from Wiki Commons commons.wikimedia.org/wiki/Air_pollution

Pollutants are categorized into two major types based on how they originated namely primary and secondary pollutants. **Primary pollutants** are those released directly from the source into the air in a harmful form. The primary pollutants that account for nearly all air pollution problems are carbon monoxide (58%), volatile organic compounds (VOCs, 11%), nitrogen oxides (15%), sulfur dioxides (13%), and particulate material (3%). **Secondary pollutants** are produced through reactions between primary pollutants and normal atmospheric compounds. For example, ground-level ozone forms over urban areas through reactions, powered by sunlight, between primary pollutants (oxides of nitrogen) and other atmospheric gases such as VOCs.

Criteria pollutants

Under the Clean Air Act, the Environmental Protection Agency (EPA) establishes air quality standards to protect public health and the environment. EPA has set national air quality standards for six common air pollutants namely: 1) carbon monoxide; 2) ground-level ozone; 3) nitrogen dioxide; 4) Sulfur dioxide; 5) lead; and 6) particulate matter (also known as particle pollution). Of the six pollutants, particle pollution and ground-level ozone are the most widespread health threats. EPA calls these pollutants "criteria" air

pollutants because it regulates them by developing human health-based and/or environmentally-based criteria (science-based guidelines) for setting permissible levels. The set of limits based on human health is called primary standards. Another set of limits intended to prevent environmental and property damage is called secondary standards.

1. **Carbon Monoxide (CO):** is a colorless, odorless gas emitted from combustion processes, specifically, the incomplete combustion of fuel. Nationally and, particularly in urban areas, the majority of CO emissions to ambient air come from mobile sources. CO can cause harmful health effects by reducing oxygen delivery to the body's organs (like the heart and brain) and tissues. At extremely high levels, CO can cause death.
2. **Ground-level ozone (O₃):** is a colorless gas with a slightly sweet odor that is not emitted directly into the air, but is created by the interaction of sunlight, heat, oxides of nitrogen (NO_x) and volatile organic compounds (VOCs). Ozone is likely to reach unhealthy levels on hot sunny days in urban environments. Emissions from industrial facilities and electric utilities, motor vehicle exhaust, gasoline vapors, and chemical solvents are some of the major sources of NO_x and VOCs.
3. **Nitrogen dioxide (NO₂):** is one of a group of highly reactive gasses known as "oxides of nitrogen," or "nitrogen oxides (NO_x)."
Other nitrogen oxides include nitrous acid and nitric acid. NO₂ is a yellowish-brown to reddish-brown foul-smelling gas that is a major contributor to smog and acid rain. Nitrogen oxides result when atmospheric nitrogen and oxygen react at the high temperatures created by combustion engines. Most emissions in the U.S. result from combustion in vehicle engines, electrical utility, and industrial combustion.
4. **Sulfur dioxide (SO₂):** Sulfur dioxide is one of a group of highly reactive gasses known as "oxides of sulfur." The largest sources of SO₂ emissions are from fossil fuel combustion at power plants (73%) and other industrial facilities (20%). Smaller sources of SO₂ emissions include industrial processes such as extracting metals from their ores, and the burning of high sulfur containing fuels by locomotives, large ships, and non-road equipment.
5. **Lead (Pb):** is a metal found naturally in the environment as well as in manufactured products. The major sources of lead emissions have historically been from fuels in motor vehicles (such as cars and trucks) and industrial sources. As a result of EPA's regulatory efforts to remove lead from gasoline, emissions of lead from the transportation sector dramatically declined by 95 percent between 1980 and 1999, and levels of lead in the air decreased by 94 percent during that time period. The major sources of lead emissions today are ore and metal processing and piston-engine aircraft operating on leaded aviation gasoline. Today, the highest levels of lead in air are usually found near lead smelters.
6. **Particulate material (PM),** sometimes known simply as "**particulates**" refers to solid particles and liquid droplets suspended in the air we breathe. Particulate pollution is made up of a variety of components, including acids (nitrates and sulfates), organic chemicals, metals, soil or dust particles, and allergens (pollen and mold spores). The size of the particles is directly linked to their potential for causing health problems. Particles that are 10 micrometers in diameter or smaller generally pass through the throat and nose and enter the lungs. EPA groups these into two types: "**inhalable coarse particles**," with diameters larger than 2.5 micrometers and smaller than 10 micrometers and "**fine particles**," with diameters that are 2.5 micrometers and smaller. How small is 2.5 micrometers? Think about a single hair from your head. The average human hair is about 70 micrometers in diameter – making it 30 times larger than the largest fine particle (Figure 3.4.2). Our respiratory systems are equipped to filter larger particles out of the air once it is inhaled. However, the lungs are vulnerable to both coarse particles (PM₁₀), and fine particles (PM_{2.5}). These can slip past the respiratory system's natural defenses and get deep into the lungs and some may even get into the bloodstream. Coarse particles come from road dust while fine particles come from combustion processes.

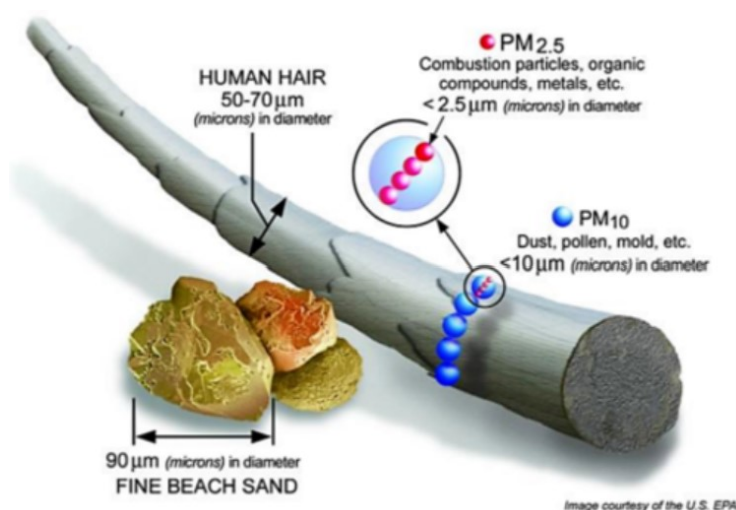


Figure 3.4.2 Graphic showing size comparison of particulate material (PM_{10} and $PM_{2.5}$) compared to fine beach sand and human hair. Source: US EPA <http://www.epa.gov/air/particlepollution/basic.html>

Volatile Organic Compounds

Volatile organic compounds (VOCs) are carbon-containing chemicals emitted as gases from natural and human-made sources. Natural sources of VOCs include plants, the largest source, and bacteria in the guts of termites and ruminant animals. These compounds are generally oxidized to carbon monoxide and carbon dioxide in the atmosphere. VOCs are of great concern because they are precursors for the formation of ozone, a secondary air pollutant.

A large number of synthetic organic chemicals such as benzene, toluene, formaldehyde, vinyl chloride, chloroform, and phenols are widely used as ingredients in countless household products. Paints, paint strippers, varnishes, many cleaning, disinfecting, cosmetic, degreasing, and hobby products all contain VOCs. Fuels are also made up of organic chemicals. All of these products can release organic compounds while you are using them, and, to some degree, when they are stored. The “new car smell” characteristic of new cars is from a complex mix of dozens of VOC. Also, concentrations of many VOCs are consistently higher indoors (up to ten times higher) than outdoors. They are often held responsible for sick building syndrome, an illness resulting from indoor pollution in which the specific cause is not identified.

Smog

Smog is a mixture of air pollutants (sulfur dioxide, nitrogen oxides, ozone, and particulates) that often form over urban areas as a result of fossil fuel combustion. The term was coined from the terms “smoke” and “fog” referring to a brownish haze that pollutes the air, greatly reducing visibility and making it difficult for some people to breathe (Figure 3.4.3 and 3.4.4). There are two main types of smog: industrial and photochemical smog. **Industrial smog** is produced primarily by the burning of fossil fuels which produces carbon dioxide (from complete combustion), carbon monoxides (from partial combustion), sulfur, and mercury. The sulfur reacts with other chemicals in the atmosphere producing several sulfur compounds including sulfur dioxide. These compounds along with particulate material make up industrial smog. **Photochemical smog** is formed when sunlight drives chemical reactions between primary pollutants from automobiles and normal atmospheric compounds. The product is a mix of over 100 different chemicals with the most abundant being ground-level ozone. This type of smog will be discussed in more detail in section 13.5.



Figure 3.4.3: Smog over Almaty city, Kazakhstan. Photo by Igors Jefimovs. Source: Wikicommons commons.wikimedia.org/wiki/File:Smog_over_Almaly.jpg



Figure 3.4.4: Smog over Santiago in Chile. Source: German Wikipedia commons.wikimedia.org/wiki/File:Santiago30std.jpg

Industrial Smog

Industrial smog or London-type smog is mainly a product of burning large amounts of high sulfur coal. Clean air laws passed in 1956 have greatly reduced smog formation in the United Kingdom; however, in other parts of the world London-type smog is still very prevalent. The main constituent of London-type smog is soot; however, these smogs also contain large quantities of fly ash, sulfur dioxide, sodium chloride and calcium sulfate particles. If concentrations are high enough, sulfur dioxide can react with atmospheric hydroxide to produce sulfuric acid, which will precipitate as acid rain.



Concerns about the harmful effects of acid rain have led to strong pressure on industry to minimize the release of SO_2 and NO . For example, coal-burning power plants now use SO_2 “scrubbers,” which trap SO_2 by its reaction with lime (CaO) to produce calcium sulfite dihydrate ($\text{CaSO}_3 \cdot 2\text{H}_2\text{O}$; Figure 3.4.5).

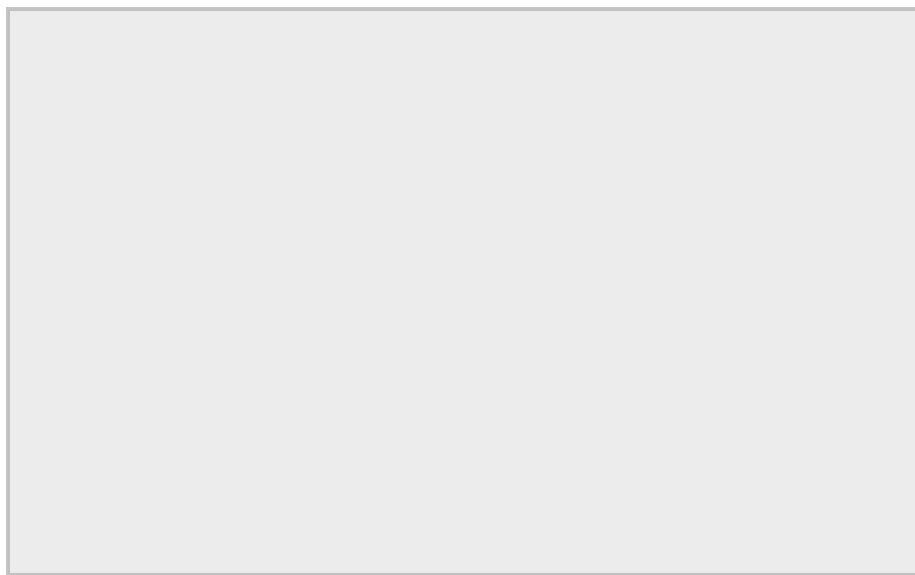


Figure 3.4.5: Schematic Diagram of a Wet Scrubber System. In coal-burning power plants, SO_2 can be removed (“scrubbed”) from exhaust gases by its reaction with a lime (CaO) and water spray to produce calcium sulfite dihydrate ($\text{CaSO}_3 \cdot 2\text{H}_2\text{O}$). Removing SO_2 from the gases prevents its conversion to SO_3 and subsequent reaction with rainwater (acid rain). Scrubbing systems are now commonly used to minimize the environmental effects of large-scale fossil fuel combustion.

Toxic pollutants

Toxic air pollutants, also known as hazardous air pollutants, are those pollutants that are known or suspected to cause cancer or other serious health effects, such as reproductive effects or birth defects, or adverse environmental effects. Examples of toxic air pollutants include **benzene**, which is found in gasoline; **perchloroethylene**, which is emitted from some dry cleaning facilities; **methylene chloride**, which is used as a solvent and paint stripper by a number of industries; and others such as dioxin, asbestos, toluene, and metals such as cadmium, mercury, chromium, and lead compounds.

Most air toxics originate from human-made sources, including mobile sources (e.g., cars, trucks, buses) and stationary sources (e.g., factories, refineries, power plants), as well as indoor sources (e.g., some building materials and cleaning solvents). Some air toxics are also released from natural sources such as volcanic eruptions and forest fires. Exposure to air toxics is mainly through breathing but some toxic air pollutants such as mercury can deposit onto soils or surface waters, where they are taken up by plants and ingested by animals and are eventually magnified up through the food chain. Like humans, animals may experience health problems if exposed to sufficiently high quantities of air toxics over time.

Effects of Air Pollution on Human Health

The World Health Organization (WHO) and other international agencies recognize air pollution as a major threat to human health. Numerous scientific studies have linked air pollution to a variety of health problems (Table 3.4.1) including: aggravation of respiratory and cardiovascular diseases; decreased lung function; increased frequency and severity of respiratory symptoms such as difficulty breathing and coughing; increased susceptibility to respiratory infections; effects on the nervous system, including the brain, such as IQ loss and impacts on learning, memory, and behavior; cancer; and premature death. Immediate effects of air pollution may show up after a single exposure or repeated exposures. Other health effects may show up either years after exposure has occurred or only after long or repeated periods of exposure.

Immediate effects of air pollution include irritation of the eyes, nose, and throat, headaches, dizziness, and fatigue. Such immediate effects are usually short-term and treatable. Sometimes the treatment is simply eliminating the person's exposure to the source of the pollution, if it can be identified. Symptoms of some diseases, including asthma, hypersensitivity pneumonitis, and humidifier fever, may also show up soon after exposure to some indoor air pollutants.

Table 3.4.1: Sources and health effects of criteria pollutants

Pollutant	Sources	Health Effects
Ground-level Ozone (O_3)	Secondary pollutant typically formed by chemical reaction of volatile organic compounds (VOCs) and NO_x in the presence of sunlight.	Decreases lung function and causes respiratory symptoms, such as coughing and shortness of breath; aggravates asthma and other lung diseases leading to increased medication use, hospital admissions, emergency department (ED) visits, and premature mortality.
Particulate Matter (PM)	Emitted or formed through chemical reactions; fuel combustion (e.g., burning coal, wood, diesel); industrial processes; agriculture (plowing, field burning); and unpaved roads.	Short-term exposures can aggravate heart or lung diseases leading to respiratory symptoms, increased medication use, hospital admissions, ED visits, and premature mortality; long-term exposures can lead to the development of heart or lung disease and premature mortality.
Lead	Smelters (metal refineries) and other metal industries; combustion of leaded gasoline in piston engine aircraft; waste incinerators; and battery manufacturing.	Damages the developing nervous system, resulting in IQ loss and impacts on learning, memory, and behavior in children. Cardiovascular and renal effects in adults and early effects related to anemia.
Oxides of Nitrogen (NO_x)	Fuel combustion (e.g., electric utilities, industrial boilers, and vehicles) and wood burning.	Aggravate lung diseases leading to respiratory symptoms, hospital admissions, and ED visits; increased susceptibility to respiratory infection.
Carbon Monoxide (CO)	Fuel combustion (especially vehicles), industrial processes, fires, waste combustion, and residential wood burning.	Reduces the amount of oxygen reaching the body's organs and tissues; aggravates heart disease, resulting in chest pain and other symptoms leading to hospital admissions and ED visits.
Sulfur Dioxide (SO_2)	Fuel combustion (especially high-sulfur coal); electric utilities and industrial processes; and natural sources such as volcanoes.	Aggravates asthma and increased respiratory symptoms. Contributes to particle formation with associated health effects.

Source: www.epa.gov

The likelihood of immediate reactions to air pollutants depends on several factors. Age and preexisting medical conditions are two important influences. Some sensitive individuals appear to be at greater risk for air pollution-related health effects, for example, those with pre-existing heart and lung diseases (e.g., heart failure/ischemic heart disease, asthma, emphysema, and chronic bronchitis), diabetics, older adults, and children. In other cases, whether a person reacts to a pollutant depends on individual sensitivity, which varies tremendously from person to person. Some people can become sensitized to biological pollutants after repeated exposures, and it appears that some people can become sensitized to chemical pollutants as well.

Summary

- Air pollution refers to the introduction, into the atmosphere, of substances that have harmful effects on humans, other living organisms, and the environment either as solid particles, liquid droplets or gases.
- Major sources of air pollution are gas emissions from fossil fueled vehicles and their reaction products (CO , NO_x , Ozone) as well as particulate matter, SO_2 , and VOCs.

- Examples of toxic air pollutants include benzene, which is found in gasoline; perchloroethylene, which is emitted from some dry cleaning facilities; methylene chloride, which is used as a solvent and paint stripper by a number of industries; and others such as dioxin, asbestos, toluene, and metals such as cadmium, mercury, chromium, and lead compounds.
- Smog is a mixture of air pollutants (sulfur dioxide, nitrogen oxides, ozone, and particulates) that often form over urban areas as a result of fossil fuel combustion. The term was coined from the terms “smoke” and “fog” referring to a brownish haze that pollutes the air, greatly reducing visibility and making it difficult for some people to breathe.
- Numerous scientific studies have linked air pollution to a variety of health problems.

Contributors and Attributions

- Libretext: Introduction to Environmental Science (Zendher et al.)
- TextMap: General CHemistry (Averill and Eldredge)
- [Marisa Alviar-Agnew](#) ([Sacramento City College](#))
- Wikipedia

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3.5: Greenhouse Gases

We will be covering the four major categories of greenhouse gases that have been impacted by humans the most. See Table 3.5.1 for a numeric comparison of these greenhouse gases.

- Carbon dioxide, CO₂
- Methane, CH₄
- Nitrous oxide, N₂O
- Synthetic fluorinated gases (including hydrofluorocarbons (HFCs), perfluorocarbons (PFCs), and sulfur hexafluoride (SF₆))

Carbon dioxide (CO₂) is the greenhouse gas responsible for most of the human-caused climate change in our atmosphere. It has the highest concentration in the atmosphere of any of the greenhouse gases that we'll discuss here. Remember that CO₂ is a direct product of both combustion and cellular respiration, causing it to be produced in great quantities both naturally and anthropogenically. Any time biomass or fossil fuels are burned, CO₂ is released. Major anthropogenic sources include: electricity production from coal-fired and natural gas power plants, transportation, and industry (Chapter 4). To get an idea of how CO₂ concentration has changed over time, watch this video compiled by the National Oceanic and Atmospheric Administration (NOAA): <http://www.esrl.noaa.gov/gmd/ccgg/trends/history.html>. This video contains atmospheric CO₂ concentrations measured directly, dating back to 1958, as well as atmospheric CO₂ concentrations measured indirectly from ice core data, dating back to 800,000 BCE. By 1990, a quantity of over seven billion tons of carbon (equivalent to 26 billion tons of carbon dioxide when the weight of the oxygen atoms are also considered) was being emitted into the atmosphere every year, much of it from industrialized nations. Similar to the action of the naturally existing greenhouse gases, any additional greenhouse gases leads to an increase in the surface temperature of the Earth.

While CO₂ is produced by aerobic cellular respiration, gases such as CH₄ and N₂O are often the products of anaerobic metabolisms. Agriculture is a major contributor to CH₄ emissions. In addition to anaerobic bacteria, methane is also a significant component of natural gas, and is commonly emitted through the mining and use of natural gas and petroleum, in addition to coal mining. Finally, landfills contribute significantly to CH₄ emissions, as the waste put into the landfill largely undergoes anaerobic decomposition as it is buried under many layers of trash and soil. Natural sources of CH₄ include swamps and wetlands, and volcanoes.

The vast majority of N₂O production by humans comes from agricultural land management. While some N₂O is naturally emitted to the atmosphere from soil as part of the nitrogen cycle, human changes in land management, largely due to agricultural practices, have greatly increased N₂O emissions. Some N₂O is also emitted from transportation and industry.

Due to their relatively high concentrations in the atmosphere compared to synthetic gases, CO₂, CH₄, and N₂O, are responsible for most of the human-caused global climate change over the past century. Figure 3.5.1 shows the increases in all three gases following the Industrial Revolution. Ice core data shows us that the atmospheric CO₂ concentration never exceeded 300 ppm before the Industrial Revolution. As of early 2015, the current atmospheric CO₂ concentration is 400 ppm.

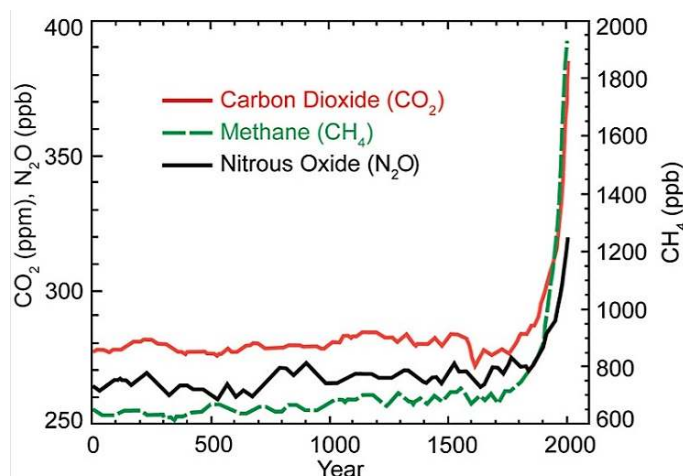


Figure 3.5.1: Increase in greenhouse gas concentrations in the atmosphere over the last 2,000 years. Increases in concentrations of these gases since 1750 are due to human activities in the industrial era. Concentration units are parts per million (ppm) or parts per billion (ppb), indicating the number of molecules of the greenhouse gas per million or billion molecules of air. Source: USGCRP (2009)

One class of greenhouse gas chemicals that has no natural sources is the fluorinated gases. These include HFCs, PFCs, and SF₆, among others. Because these are synthetic chemicals that are only created by humans, these gases were essentially non-existent before the Industrial Revolution. These synthetic gases are used for a wide variety of applications, from refrigerants to semiconductor manufacturing, and propellants to fire retardants. They tend to have a long lifetime in the atmosphere, as seen in Table 3.5.1. Some of these chemicals, as well as the older **chlorofluorocarbons** (CFCs), have been phased out by international environmental legislation under the Montreal Protocol (see Chapter 6). Due to their long lifespan, many of these now banned CFCs remain in the atmosphere. Newer chemical replacements, such as HFCs, provide many of the same industrial applications, but unfortunately have their own environmental consequences.

Just as greenhouse gases differ in their sources and their residence time in the atmosphere, they also differ in their ability to produce the greenhouse effect. This is measured by the **global warming potential**, or GWP, of each greenhouse gas. The GWP of a greenhouse gas is based on its ability to absorb and scatter energy, as well as its lifetime in the atmosphere. Since CO₂ is the most prevalent greenhouse gas, all other greenhouse gases are measured relative to it. As the reference point, CO₂ always has a GWP of 1. Note the very high GWP values of the synthetic fluorinated gases in Table 3.5.1. This is largely due to their very long residence time in the atmosphere. Also note the higher GWP values for CH₄ and N₂O compared to CO₂.

Greenhouse gas	Chemical formula or abbreviation	Lifetime in atmosphere	Global warming potential (100-year)
Carbon dioxide	CO ₂	Variable	1
Methane	CH ₄	12 years	28-36
Nitrous oxide	N ₂ O	114 years	298
Hydrofluorocarbons	Abbreviation: HFCs	1-270 years	12-14,800
Perfluorocarbons	Abbreviation: PFCs	2,600-50,000 years	7,390
Sulfur hexafluoride	SF ₆	3,200 years	22,800

Other climate influencers

In addition to greenhouse gases, other manmade changes may be forcing climate change. Increases in near-surface ozone from internal combustion engines, aerosols such as carbon black, mineral dust and aviation-induced exhaust are acting to raise the surface temperature. This primarily occurs due to a decrease in the **albedo** of light-colored surfaces by the darker-colored carbon black, soot, dust, or particulate matter. It is often considered more comfortable to wear a white shirt on a hot summer day than a

black shirt. Why is this? Because the lighter-colored material bounces more solar radiation back toward space than the darker-colored material does, allowing it to stay cooler. The darker-colored material absorbs more solar radiation, increasing its temperature. Just as the white shirt has a higher albedo than the black shirt, light-colored objects in nature (such as snow) have a higher albedo than dark-colored objects (such as soot or dust). As humans increase the amount of carbon black, soot, dust, and particulates in the atmosphere, we decrease the albedo of light-colored surfaces, causing them to absorb more solar radiation and become warmer than they would without human influence. An example of this can be seen in the snow on Figure 3.5.2.



Figure 3.5.2: A photograph of the extreme dust deposition from the deserts of the Colorado Plateau onto the Colorado Rockies snowpack in 2009. Taken from the high point of the Senator Beck Basin in the San Juan Mountains, it captures the extent of the impact of darkening in which the snow albedo dropped to about 30%, more than doubling the absorption of sunlight. Credit: S. McKenzie Skiles, Snow Optics Laboratory, NASA/JPL

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3.6: Stratospheric Ozone- Earth's Vital Shield

Learning Objectives

- Describe the depletion of the ozone layer.
- Explain how chlorine and bromine atoms react with ozone which leads to the depletion of the ozone layer.

The earth's stratospheric ozone layer plays a critical role in absorbing ultraviolet radiation emitted by the sun. In the last thirty years, it has been discovered that stratospheric ozone is depleting due to anthropogenic pollutants. There are several chemical reactions that can deplete stratospheric ozone; however, some of the most significant of these involve the catalytic destruction of ozone by halogen radicals such as chlorine and bromine.

Introduction

The atmosphere of the Earth is divided into five layers. In order of closest and thinnest to farthest and thickest the layers are listed as follows: troposphere, stratosphere, mesosphere, thermosphere and exosphere. The majority of the ozone in the atmosphere resides in the stratosphere, which extends from six miles above the Earth's surface to 31 miles. Humans rely heavily on the absorption of ultraviolet B rays by the ozone layer because UV-B radiation causes skin cancer and can lead to genetic damage. The ozone layer has historically protected the Earth from harmful UV rays, although in recent decades this protection has diminished due to stratospheric ozone depletion.

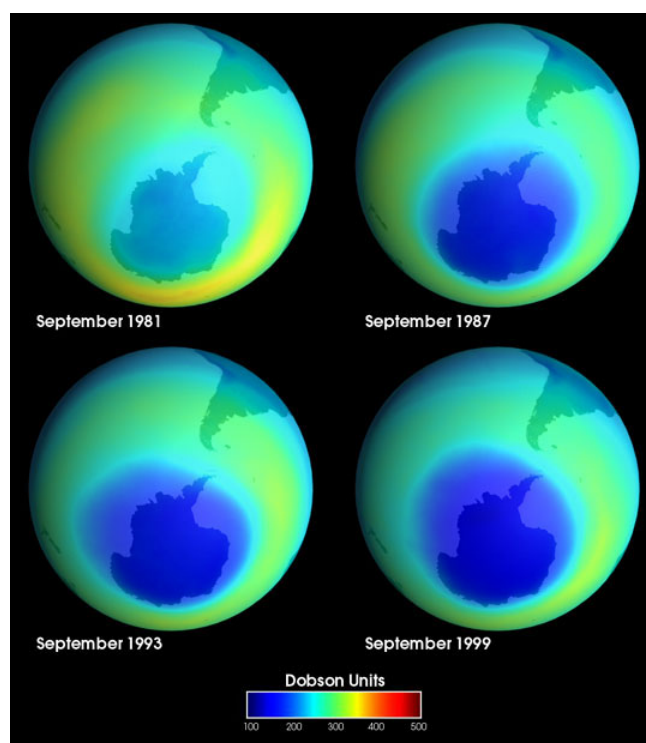


Figure courtesy of NASA.

Ozone depletion is largely a result of man-made substances. Humans have introduced gases and chemicals into the atmosphere that have rapidly depleted the ozone layer in the last century. This depletion makes humans more vulnerable to the UV-B rays which are known to cause skin cancer as well as other genetic deformities. The possibility of ozone depletion was first introduced by scientists in the late 1960's as dreams of supersonic transport began to become a reality. Scientists had long been aware that nitric oxide (NO) can catalytically react with ozone (O_3) to produce O_2 molecules; however, NO molecules produced at ground level have a half-life far too short to make it into the stratosphere. It was not until the advent of commercial supersonic jets (which fly in the stratosphere and at an altitude much higher than conventional jets) that the potential for NO to react with stratospheric ozone

became a possibility. The threat of ozone depletion from commercial supersonic transport was so great that it is often cited as the main reason why the US federal government pulled support for its development in 1971. Fear of ozone depletion was abated until 1974 when Sherwood Rowland and Mario Molina discovered that chlorofluorocarbons could be photolyzed by high-energy photons in the stratosphere. They discovered that this process could release chlorine radicals that would catalytically react with O_3 and destroy the molecule. This process is called the Rowland-Molina theory of O_3 depletion.

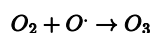
The Chapman Cycle

The ozone (O_3) located in the stratosphere is in a constant cycle with oxygen molecules (O_2) and their interaction with ultraviolet rays. This process is considered a cycle because of its constant conversion between different molecules of oxygen. The ozone layer is created when ultraviolet rays react with oxygen molecules (O_2) to create ozone (O_3) and atomic oxygen (O). This process is called the *Chapman cycle*.

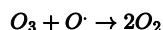
Step 1: An oxygen molecule is photolyzed by solar radiation, creating two oxygen radicals:



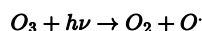
Step 2: Oxygen radicals then react with molecular oxygen to produce ozone:



Step 3: Ozone then reacts with an additional oxygen radical to form molecular oxygen:



Step 4: Ozone can also be recycled into molecular oxygen by reacting with a photon:



It is important to keep in mind that ozone is constantly being created and destroyed by the Chapman cycle and that these reactions are natural processes, which have been taking place for millions of years. Because of this, the thickness of the ozone layer at any particular time can vary greatly. It is also important to know that O_2 is constantly being introduced into the atmosphere through photosynthesis, so the ozone layer has the capability of regenerating itself.

Chemistry of Ozone Depletion

Chlorofluorocarbons (CFCs), made up of chlorine, fluorine, and carbon atoms, are extremely stable. This extreme stability allows CFC's to slowly make their way from the troposphere into the stratosphere (most other less-stable molecules decompose before they can cross into the stratosphere from the troposphere). This prolonged life in the atmosphere allows them to reach great altitudes where photons are more energetic. When the CFC's come into contact with these high-energy photons, their individual components are freed from the whole creating highly reactive free radicals. The following reaction displays how Cl atoms have an ozone-destroying cycle:



Chlorine is able to destroy so much of the ozone because it acts as a catalyst, meaning it is regenerated during this process. Chlorine initiates the breakdown of ozone and combines with a freed oxygen atom to create two oxygen molecules. After each reaction, chlorine begins the destructive cycle again with another ozone molecule. One chlorine atom can thereby destroy thousands of ozone molecules. Because ozone molecules are being broken down they are unable to absorb any ultraviolet light so we experience more intense UV radiation at the Earth's surface.

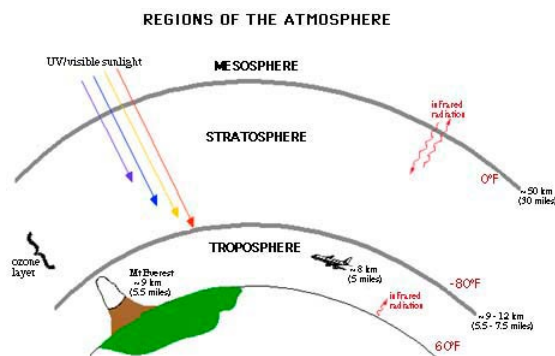


Figure courtesy of NOAA.

From 1985 to 1988, researchers studying atmospheric properties over the south pole continually noticed significantly reduced concentrations of ozone directly over the continent of Antarctica. For three years it was assumed that the ozone data was incorrect and was due to some instrument malfunction. In 1988, researchers finally realized an instrument malfunction had not occurred and the data was correct. They concluded that an enormous hole in the ozone layer had indeed developed over Antarctica. Examination of NASA satellite data later showed that the hole had begun to develop in the mid-1970's.

The ozone hole over Antarctica is formed by a slew of unique atmospheric conditions over the continent that combine to create an ideal environment for ozone destruction.

- Because Antarctica is surrounded by water, winds over the continent blow in a unique clockwise direction creating a so-called "polar vortex" that effectively contains a single static air mass over the continent. As a result, air over Antarctica does not mix with air in the rest of the Earth's atmosphere.
- Antarctica has the coldest winter temperatures on earth, often reaching -110 °F. These chilling temperatures result in the formation of polar stratospheric clouds (PSC's) which are a conglomeration of frozen H₂O and HNO₃. Due to their extremely cold temperatures, PSC's form an electrostatic attraction with CFC molecules as well as other halogenated compounds

As spring comes to Antarctica, the PSC's melt in the stratosphere and release all of the halogenated compounds that were previously absorbed into the cloud. In the antarctic summer, high-energy photons are able to photolyze the halogenated compounds, freeing halogen radicals that then catalytically destroy O₃. Because Antarctica is constantly surrounded by a polar vortex, radical halogens are not able to be diluted over the entire globe. The ozone hole develops as a result of this process.

Recent research suggests that the strength of the polar vortex from any given year is directly correlated to the size of the ozone hole. In years with a strong polar vortex, the ozone hole is seen to expand in diameter, whereas in years with a weaker polar vortex, the ozone hole is noted to shrink

Ozone Depleting Substances

The following substances are listed as ozone-depleting substances under Title VI of the United States Clean Air Act:

Table 3.6.1: Ozone Depleting Substances And Their Ozone-Depletion Potential. Taken directly from the Clean Air Act, as of June 2010.

Substance	Ozone- depletion potential
chlorofluorocarbon-11 (CFC-11)	1.0
chlorofluorocarbon-12 (CFC-12)	1.0
chlorofluorocarbon-13 (CFC-13)	1.0
chlorofluorocarbon-111 (CFC-111)	1.0
chlorofluorocarbon-112 (CFC-112)	1.0
chlorofluorocarbon-113 (CFC-113)	0.8
chlorofluorocarbon-114 (CFC-114)	1.0
chlorofluorocarbon-115 (CFC-115)	0.6
chlorofluorocarbon-211 (CFC-211)	1.0
chlorofluorocarbon-212 (CFC-212)	1.0
chlorofluorocarbon-213 (CFC-213)	1.0
chlorofluorocarbon-214 (CFC-214)	1.0
chlorofluorocarbon-215 (CFC-215)	1.0
chlorofluorocarbon-216 (CFC-216)	1.0
chlorofluorocarbon-217 (CFC-217)	1.0
halon-1211	3.0
halon-1301	10.0
halon-2402	6.0
carbon tetrachloride	1.1
methyl chloroform	0.1
hydrochlorofluorocarbon-22 (HCFC-22)	0.05
hydrochlorofluorocarbon-123 (HCFC-123)	0.02
hydrochlorofluorocarbon-124 (HCFC-124)	0.02
hydrochlorofluorocarbon-141(b) (HCFC-141(b))	0.1
hydrochlorofluorocarbon-142(b) (HCFC-142(b))	0.06

Montreal Protocol

International policy efforts to restrict production of ozone-depleting CFCs culminated in the 1987 treaty known as the **Montreal Protocol** in which signing nations agreed to cut CFC production in half by 1998. At least five follow-up agreements since then helped to deepen the cuts, advanced timetables for compliance, and addressed additional ozone-depleting substances such as halons, methyl chloroform, carbon tetrachloride, and hydrochlorofluorocarbons (HCFCs). Most countries around the world have phased out production of the substances covered by the agreements and industry has been able to shift to safer alternative chemicals. As a result, there's evidence that the Antarctic ozone hole has stopped growing worse, although recovery is not expected anytime soon. Phasing out CFCs and HCFCs is also beneficial in protecting the earth's climate, as these substances are also very damaging greenhouse gases.

As part of the United States' commitment to implementing the Montreal Protocol, the U.S. Congress amended the Clean Air Act (section 6.7), adding provisions for protection of the ozone layer. Most importantly, the amended Act required a gradual end to the production of chemicals that deplete the ozone layer. The Clean Air Act amendments passed by Congress require the Environmental Protection Agency (EPA) to develop and implement regulations for the responsible management of ozone-depleting substances in the United States. Under the Clean Air Act, EPA has created several regulatory programs to address numerous issues, including:

- ending the production of ozone-depleting substances,
- ensuring that refrigerants and halon fire extinguishing agents are recycled properly,
- identifying safe and effective alternatives to ozone-depleting substances,
- banning the release of ozone-depleting refrigerants during the service, maintenance, and disposal of air conditioners and other refrigeration equipment,
- requiring that manufacturers label products either containing or made with the most harmful ozone-depleting substances.

Summary

- Any disruption of the balance that results in a higher rate of ozone destruction than ozone creation would result in the depletion of ozone.
- Some compounds that release chlorine or bromine when they are exposed to intense UV light in the stratosphere contribute to ozone depletion and are called ozone-depleting substances ([ODS](#))
- The Montreal Protocol is an international agreement that committed all parties (signatory nations) to a schedule for phasing out the production and use of CFCs and other substances known to be harmful to the ozone layer.

Contributors and Attributions

- Libretext: Supplemental Module, Physical and Theoretical Chemistry
- [Marisa Alviar-Agnew](#) ([Sacramento City College](#))
- Caralyn Zehnder et al.
- Erin Avram ([Cleveland State University](#))

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3.7: Consequences of Climate Change

In this section, we will discuss the effects of climate change, both those that have already been observed, as well as future predictions based on scientific climate models. Here, the differences between the terms **global warming** and **climate change** become apparent. Global warming refers to the increase in the average temperature of the Earth's atmosphere due to elevated greenhouse gas concentrations, heightening the greenhouse effect. We have also seen, and expect to continue to see, other changes occurring in the **climate** of the Earth, and we expect to continue to observe, changes in other chemical, physical, and biological aspects of the Earth's environment. We will only discuss some of the consequences of climate change in this section, including changes in temperature, precipitation, ocean level, and ocean acidity. There are many more changes that have been seen, and are projected to continue in the future. These include: changes in the amount and distribution of ice and snow, changes in seasonality, ecosystem shift, and habitat changes of plant and animal populations, in addition to others. For more information about these consequences of climate change, visit this site: <http://www.epa.gov/climatechange/sci...ors/index.html>.

Temperature and precipitation

Temperature and **precipitation** are the two most direct impacts on the Earth's climate due to climate change. By now, you should already understand why an increase in greenhouse gas levels in the atmosphere causes an increase in temperature. But why does it also impact precipitation patterns? As the air in the troposphere warms and cools, the amount of water vapor that it holds changes dramatically. The high summer humidity in this region is possible due to the increased capability warm air has to hold water vapor. Simply put, warmer air can hold more water than cooler air. As the air cools, its ability to hold water vapor decreases, and any excess water will leave the air as liquid water. A great example of this is the formation of dew on surfaces overnight. During the day, the temperature is warmer than it is at night, and the air has a relatively high holding capacity for water vapor. When the sun sets, the air cools, decreasing its capacity to hold water vapor. That extra water must go somewhere, and it does that by accumulating on surfaces. Similarly, when warm and cool air fronts collide, the chances for rain and thunderstorms increase. Furthermore, an increase in temperature enhances evaporation occurring at the Earth's surface. This increased evaporation leads to greater concentrations of water vapor in the atmosphere which can lead to increased precipitation.

The change in temperature that we have already seen in the Earth's average atmospheric temperature is relatively small (about 0.6 °C). However, as with many of the aspects of climate change, the potential for greater changes increases dramatically as time progresses in the future. This can be seen in Figure 3.7.1, which displays a model of the predicted temperature increase. Notice that these changes occur relatively rapidly, and are not uniform across the globe. What might be some of the reasons for this?

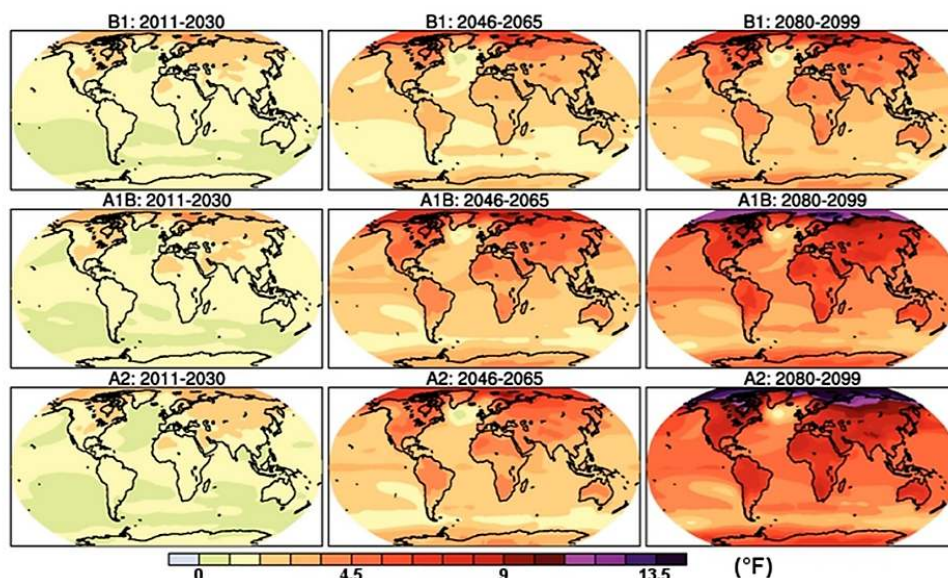


Figure 3.7.1: Projected changes in global average temperatures under three emissions scenarios (rows) for three different time periods (columns). Changes in temperatures are relative to 1961- 1990 averages. The scenarios come from the IPCC Special Report on Emissions Scenarios: B1 is a low emissions scenario, A1B is a medium-high emissions scenario, and A2 is a high emissions scenario. Source: IPCC Working Group I: The Physical Science Basis, 2007.

Changes in precipitation occur due to a variety of factors, including changes in atmospheric water vapor content due to changing temperature, as discussed above. Also at play is the heightened **evaporation** rate of water on Earth's surface under warmer temperatures. More evaporation leads to more precipitation. Finally, shifts in wind patterns impact the distribution of precipitation events. As you can see in Figure 3.7.2, there are some areas of the globe that are expected to have an increase in precipitation, while others are expected to have a dramatic decrease. Some major population centers projected to have a moderate to severe precipitation increase include (population estimates of the metropolitan area given in parentheses): New York, United States (20.1 million); Bogotá, Colombia (12.1 million); and Manila, Philippines (11.9 million). What sort of challenges might these cities face in the future as they deal with this change in their climate?

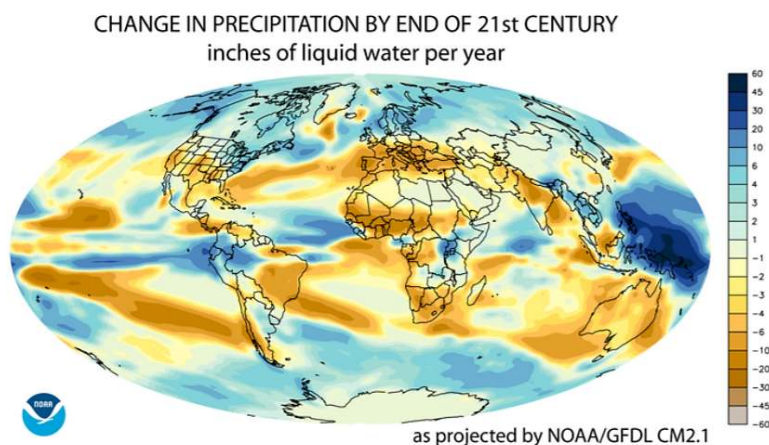


Figure 3.7.2. The plotted precipitation differences were computed as the difference between the 2081 to 2100 20-year averages minus the 1951 to 2000 50-year average. Blue areas project increases in precipitation; brown areas project decreases.

In contrast, many more major metropolitan areas are projected to have a moderate to severe precipitation decrease (droughts) by the end of the 21st century. These include Delhi, India (21.8 m.); Lagos, Nigeria (21 m.); São Paulo, Brazil (20.9 m.); Kolkata, India (14.6 m.); Istanbul, Turkey (14.4 m.); Los Angeles, United States (13.3 m.); Rio de Janeiro, Brazil (12 m.); Paris, France (12 m.); and Lahore, Pakistan (11.3 m.). The largest challenge that these areas are likely to face is a dwindling water supply for drinking and agriculture.

Additional challenges may be felt by all areas of the world regarding changes in the seasonality or timing of precipitation, as well the form in which precipitation falls (e.g., mist or downpour; rain, ice, or snow). All of these factors affect the availability of soil water for plants, the flow of rivers and streams, and the overall accessibility of water worldwide. Furthermore, scientists predict an increase in the number and severity of storms as climate change progresses. For a full discussion of the potential impacts of this, see the assigned article.

Sea level rise

While we know that water continuously cycles around the world (see Chapter 4 for information on the water cycle), and that the overall quantity of water on Earth will not change due to global climate change, the distribution of this water is changing. In particular, oceans are increasing in volume while land ice stores (such as **glaciers**) are decreasing. This contributes to an increase in sea levels worldwide (Figure 3.7.3).

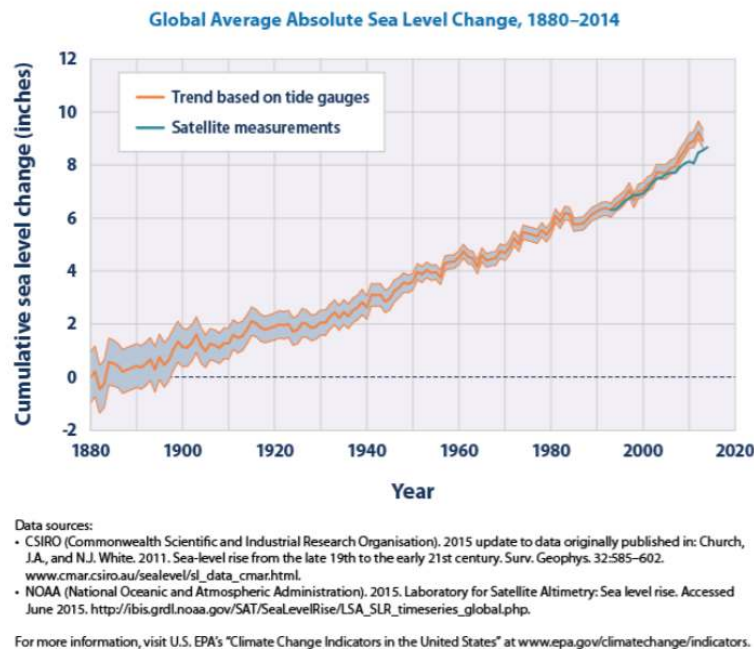


Figure 3.7.3: This graph shows average absolute cumulative changes in sea level for the world's oceans since 1880, based on a combination of tide gauge measurements and recent satellite measurements. The shaded band shows the likely range of values, based on the number of measurements collected and the precision of the methods used.

From the data in Figure 3.7.3, we see that sea level has increased at an average of 0.06 inches (0.15 cm) per year over the time period shown above. Most of this rise, however, has occurred within the most recent decades. The rate of increase has gone up to between 0.11 to 0.14 inches (0.28 to 0.36 cm) per year since 1993. There are two forces causing sea levels to rise, both caused by climate change. First, the increased global temperature has caused increased ice melting in many regions of the globe. Melting **land ice** (such as the glacier shown in Figure 3.7.4) contributes to sea level rise because water that used to be stored in ice sitting on top of land becomes running water which reaches the ocean through **runoff**. We also observe **sea ice** melting (see <http://www.epa.gov/climatechange/science/indicators/index.html> for data and figures). Sea ice, such as the ice that covers the arctic regions of the Northern Hemisphere, has no land underneath it. When it melts, the water stays in the same locations, and the overall sea level does not change.



Figure 3.7.4: Glacier Retreat

The second factor that influences sea level rise is a phenomenon called thermal expansion. Due to the physical properties of water, as water warms, its density decreases. A less dense substance will have fewer molecules in a given area than a more dense substance (see section 1.5 Density). This means that as the overall temperature of the oceans increases due to global climate change, the same amount of water molecules will now occupy a slightly larger volume. This may not seem significant, but considering the 1.3 billion trillion liters (264 billion gallons) of water in the ocean, even a small change in density can have large effects on sea level as a whole.

Scientists have already documented sea level rise in some areas of the world, including one familiar to most of us: the Southeastern United States. Figure 3.7.5 depicts the measured land area lost due to increasing sea level since 1996. Note that the Southeast (defined here as the Atlantic coast of North Carolina south to Florida) is particularly susceptible to land area loss due to the gently sloping nature of our coastline. Moving northward into the Mid-Atlantic States (defined here as Virginia north to Long Island, New York), coastal habitats tend to have a steeper geography, which protects against some losses.

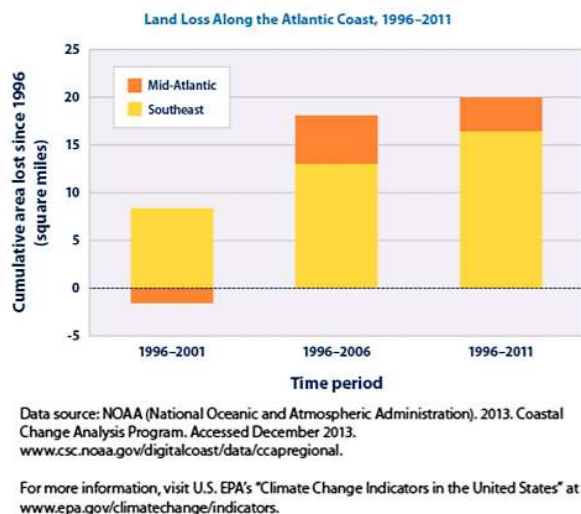


Figure 7.6.5: This graph shows the net amount of land converted to open water along the Atlantic coast during three time periods: 1996–2001, 1996–2006, and 1996–2011. The results are divided into two regions: the Southeast and the Mid-Atlantic. Negative numbers show where land loss is outpaced by the accumulation of new land.

While the ecological effects of sea level rise remain in the United States, we don't project any catastrophic loss of life, property, or livelihood for some time. This is, in part, due to large investments that we have made in infrastructure to protect our cities and farmlands. This is not the case in many areas of the world. For a discussion of the impacts of sea level rise on less industrialized nations of Bangladesh, Maldives, Kiribati, and Fiji.

Ocean acidification

Dissolved CO_2 is essential for many organisms, including shell-building animals and other organisms that form a hard coating on their exterior (e.g., shellfish, corals, Haptophyte algae). This hard coating is built out of aragonite, a mineral form of the molecule calcium carbonate, CaCO_3 . These organisms rely on the formation of carbonate ions (see Chapter 1 supplemental material for information on ions), CO_3^{2-} , from dissolved CO_2 , through a natural, chemical reaction that occurs. This takes place through a chain-reaction equation, where bicarbonate (HCO_3^-) is formed as an intermediate, and hydrogen ions (H^+) are generated (equations 3.7.1 and 3.7.2).



To have a better visualization of this process, follow along with the interactive graphic at: http://www.whoi.edu/home/oceanus_ima...ification.html.

As you can see, both equations 3.7.1 and 3.7.2 each produce one H^+ . This is significant to water chemistry because an increase in H^+ concentration means a decrease in the pH of the water. You can see in Figure 3.7.6 that a lower pH means that the liquid is more acidic. As shown in the interactive graphic, an increase in CO_2 in the atmosphere causes additional CO_2 to be dissolved in the ocean. This means that more CO_2 in the atmosphere leads to more acidic ocean environments.

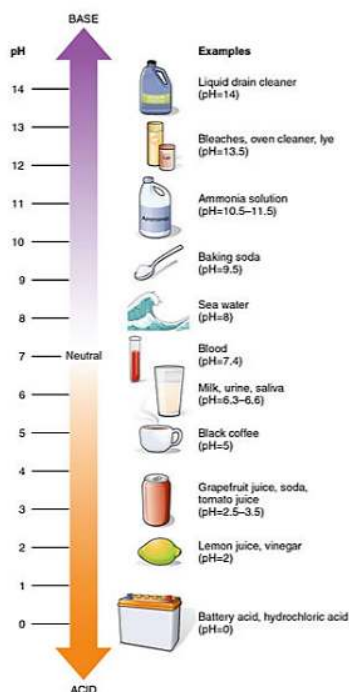


Figure 7.6.6: The pH scale and relative acidity. Illustration from Anatomy & Physiology, Connexions Web site. <http://cnx.org/content/col11496/1.6/>, Jun 19, 2013.

Unfortunately for shell-building animals, the buildup of H^+ in the more acidic ocean environment blocks the absorption of calcium and CO_3^{2-} , and makes the formation of aragonite more difficult. An aragonite deficit is already being documented in many of the world's oceans, as shown in Figure 3.7.7.

The increasing acidity of the world's oceans is resulting in habitat changes across the globe. This is only expected to worsen as atmospheric CO_2 levels continue to increase. Many organisms, including the corals that are the foundation species of the beautiful coral reefs, are very sensitive to changes in ocean pH. Scientists have documented cases of ecosystem destruction through coral bleaching, caused by the effects of climate change including ocean acidification and increased temperature. For more information, visit the NOAA Coral Reef Conservation Program website: coralreef.noaa.gov/threats/climate/.

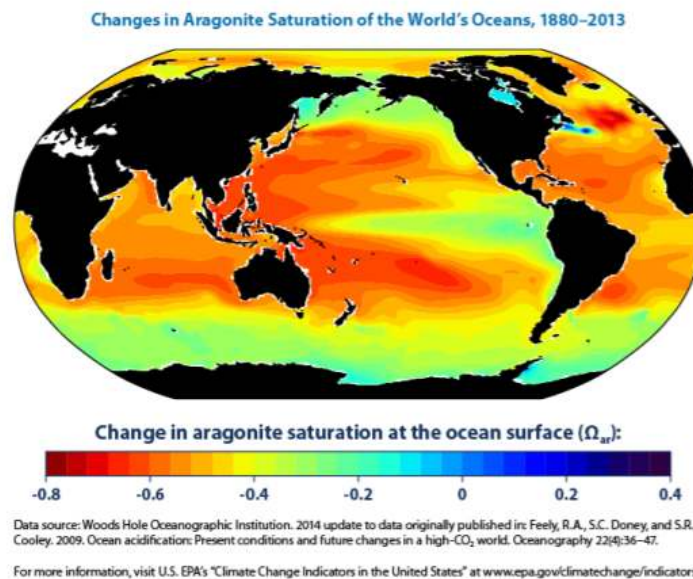


Figure 3.7.7: This map shows changes in the aragonite saturation level of ocean surface waters between the 1880s and the most recent decade (2004–2013). Aragonite is a form of calcium carbonate that many marine animals use to build their skeletons and shells. A negative change represents a decrease in saturation.

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CHAPTER OVERVIEW

4: Chemistry of Water

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4.1: Water Cycle Part 1

Learning Objectives

- Explain the distribution of Earth's water
- Know the processes and phases of the water (hydrologic) cycle.
- Understand hydrologic terms: stream, pond, river, reservoir, wetland, aquifer, and ocean.
- Recall the large consumers of water in the United States and in the world.

Water is composed of one oxygen atom bonded with two hydrogen atoms (H_2O). Despite its simplicity, water has remarkable properties. It expands when it freezes, has high surface tension (because of the polar nature of the molecules that stick together), and has a high boiling point for a liquid. Without water, life might not be able to exist on Earth, and this planet would not have the tremendous complexity and diversity that we see.



Figure 4.1.1: Picture of the Lake Erie coastline and the Fairport Harbor West Breakwater Light. Image courtesy of Erin Avram

Distribution of Earth's Water

Earth's oceans contain 97% of the planet's water, so just 3% is fresh water, water with low concentrations of salts (Figure 4.1.2). The majority of the Earth's water can be classified as being **saline** (or salt-containing). Most freshwater is trapped as ice in the vast glaciers and ice sheets of Greenland. A storage location for water such as an ocean, glacier, pond, or atmosphere is known as a **reservoir**. A water molecule may pass through a reservoir very quickly or may remain for much longer. The amount of time a molecule stays in a reservoir is known as its **residence time**.

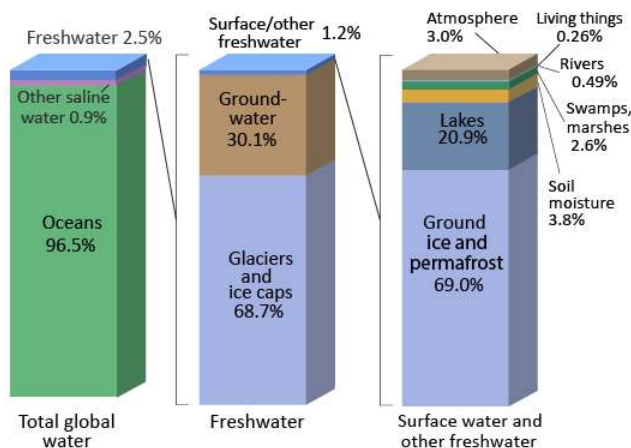


Figure 4.1.2: The distribution of Earth's water. In the first bar, notice how only 2.5% of Earth's water is freshwater - the amount needed for life to survive. The middle bar shows the breakdown of freshwater. Almost all of it is locked up in ice and in the ground. Only a little more than 1.2% of all freshwater is surface water, which serves most of life's needs. The right bar shows the breakdown of surface freshwater. Source: Igor Shiklomanov's chapter "World Freshwater resources" in Peter H. Gleick (editor), 1993 *Water in Crises: A guide to the World's Freshwater Resources*. The numbers are rounded.

The Hydrologic Cycle

Because Earth's water is present in all three states of matter (solid, liquid, and gas), it can get into a variety of environments around the planet. The movement of water around the Earth's surface is the **hydrologic (water) cycle** (Figure 4.1.3).

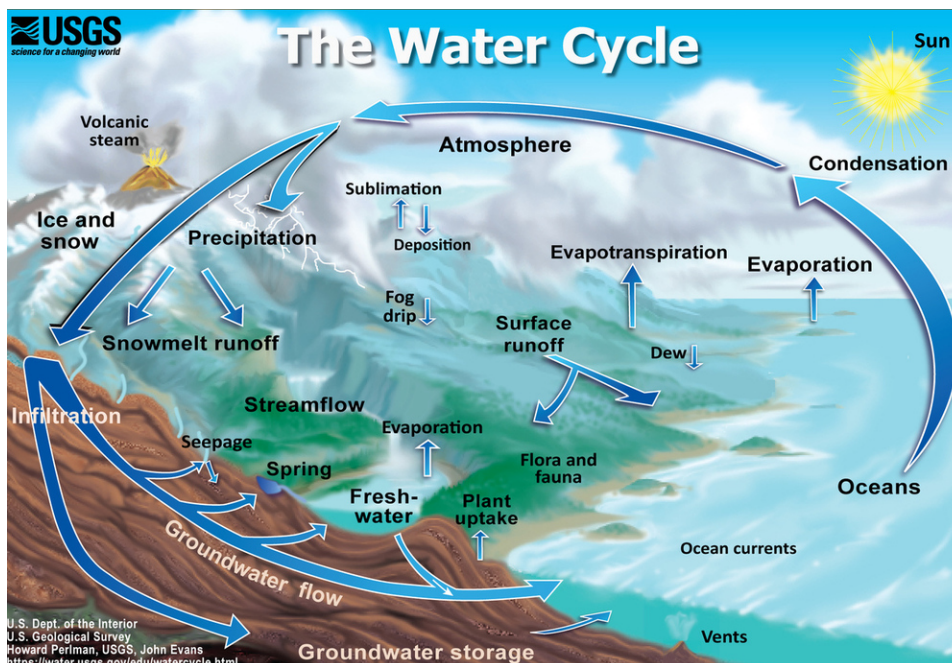


Figure 4.1.3: Because it is a cycle, the water cycle has no beginning and no end. Earth's water is always in movement, and the natural water cycle, also known as the hydrologic cycle, describes the continuous movement of water on, above, and below the surface of the Earth. Water is always changing states between liquid, vapor, and ice, with these processes happening in the blink of an eye and over millions of years.

The Sun, while many millions of kilometers away, provides the energy that drives the water cycle. It impacts the water cycle by supplying the energy needed for evaporation. Most of Earth's water is stored in the oceans where it can remain for hundreds or thousands of years. Water changes from a liquid to a gas by **evaporation** to become water vapor. The Sun's energy can evaporate water from the ocean surface or from lakes, streams, or puddles on land. Only the water molecules evaporate; the salts remain in the ocean or a freshwater reservoir.

The water vapor remains in the atmosphere until it undergoes **condensation** to become tiny droplets of liquid. The droplets gather in clouds, which are blown around the globe by the wind. As the water droplets in the clouds collide and grow, they fall from the sky as **precipitation** (rain, sleet, hail, or snow). Precipitation can fall back into the ocean and/or land surface.



Figure 4.1.4: Snow is one form of precipitation and a key transport mechanism in the water cycle. Image courtesy of Elizabeth R. Gordon

When water falls from the sky as rain it may enter streams and rivers that flow downward to oceans and lakes. Water that falls as snow may sit on a mountain for several months. Snow may become part of the ice in a glacier, where it may remain for hundreds or thousands of years. Snow and ice may go directly back into the air by **sublimation**, the process in which solid changes directly into a gas without first becoming a liquid. Although you probably have not seen water vapor sublime from a glacier, you may have seen dry ice sublime in the air.

Snow and ice slowly melt over time to become liquid water, which provides a steady flow of fresh water to streams, rivers, and lakes below. A water droplet falling as rain could also become part of a stream or a lake. At the surface, the water may eventually evaporate and reenter the atmosphere. A significant amount of water infiltrates into the ground. Soil moisture is an important reservoir for water (Figure 4.1.5). Water may seep through dirt and rock below the soil through pores infiltrating the ground to go into Earth's groundwater system. Groundwater enters **aquifers** that may store fresh water for centuries. Alternatively, the water may come to the surface through springs or find its way back to the oceans.

Water trapped in soil is important for plants to grow. Plants and animals depend on water to live and they also play a role in the water cycle. Plants take up water from the soil and release large amounts of water vapor into the air through their leaves, a process is known as *transpiration*.

It is important to note that water molecules cycle around. If climate cools and glaciers and ice caps grow, there is less water for the oceans and sea level will fall. The reverse can also happen.

Components of the Hydrologic Cycle

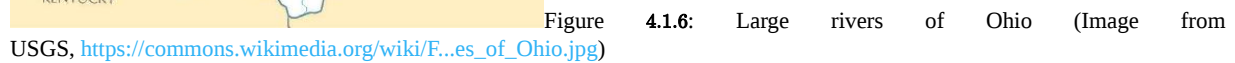
Most precipitation falls in the form of rain, but there are other forms such as snow, hail, and sleet. Once it runs sufficiently, surface water runoff is generated when the ground is saturated or impervious. **Surface water** is a major component of the hydrological cycle and one that we interact with very regularly. It includes lakes, wetlands, stormwater **runoff** (overland flow), ponds, potholes, rivers and streams.

Streams and Rivers

A river forms from water moving from a higher altitude to a lower altitude, under the force of gravity. When rain falls on land, it can seep into the ground or become runoff (water running on the surface). If the water runs on the land surface, it can converge as it moves towards lower elevation. The converging runoff can concentrate into single channels of conveyance called creeks, stream, or rivers. Streams and rivers that join up to form a larger river are called **tributaries**, Figure 4.1.5. The land area drained by a river and all its tributaries is called a **watershed** or river basin.

The area adjacent to a river that floods frequently is called a **flood plain**. These regions temporarily store excess water during storm events and frequently contain fertile soils. Properly functioning flood plains reduce the negative impacts of floods (by reducing the severity of the floods) and they assist in filtering stormwater, which results in protecting the water quality of rivers. Also, flood plains act as areas of recharge for groundwater.





Lakes, reservoirs, and ponds

Figure 4.1.7: Map of the Great Lakes with Lake Erie highlighted in red (CC BY-SA 3.0 <https://upload.wikimedia.org/wikipedia/commons/3/3e/Erie-red.png>.)

A **wetland** is an area which is home to standing water for notable parts of the year, has saturated soils for a large part of the year and has plants that require large amounts of water to survive. Wetlands include swamps, marshes, and bogs. Wetlands are identified using three characteristics: soils (water-saturated soils are present), hydrology (shallow water table) and vegetation (wetland plants that are adapted to areas that are saturated with water for long periods of time). Wetlands are important areas of biological diversity and productivity. These are also important areas where geochemical and biological cycles/ processes are consistently taking place.

For instance, wetlands are considered as areas of significant carbon sequestration (storage), which impacts global climate change. They also act as filters for storm-water runoff before it enters rivers and lakes.



Figure 4.1.8: Image of Mentor Marsh in Mentor, Ohio. Image courtesy of Erin Avram.

Oceans

As you have probably already guessed, oceans are an important component of the hydrologic cycle because they store majority of all water on Earth (about 95%). Most of the major rivers drain into them. The five oceans covering the surface of the Earth are the Atlantic, Indian, Pacific, Arctic and the Southern Ocean (Figure 4.1.8).

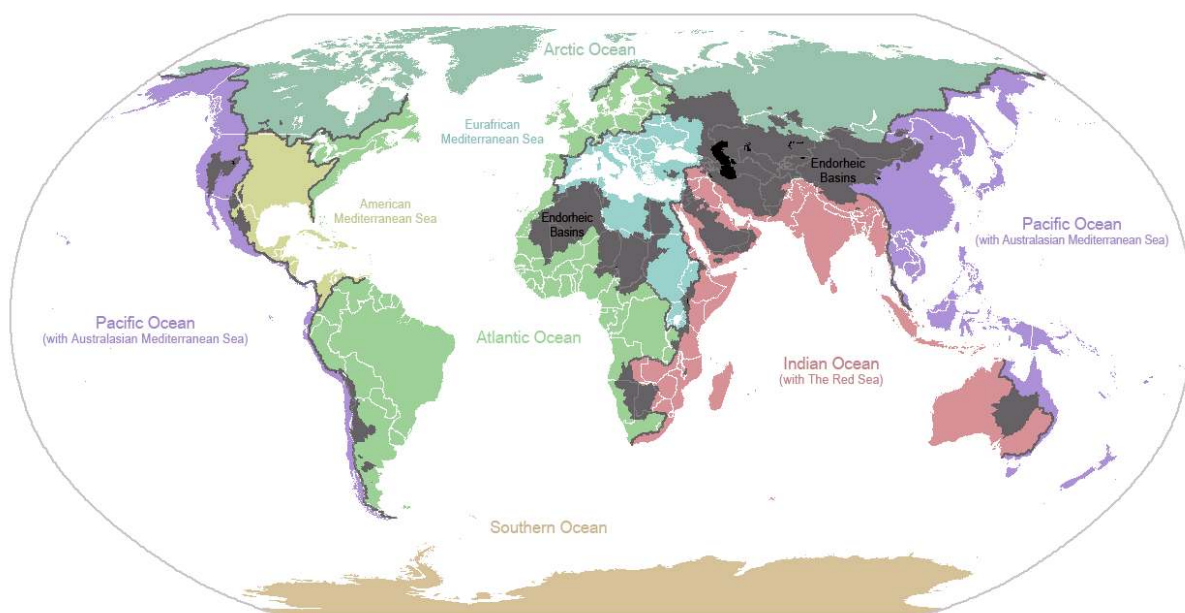


Figure 4.1.9: The five oceans found on planet Earth. The Pacific Ocean is the largest.

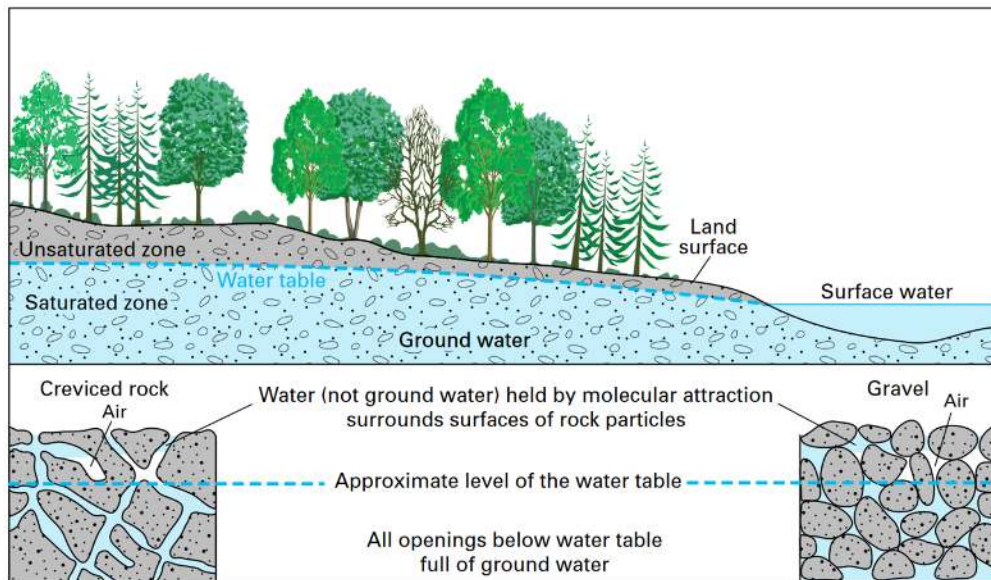
Source: [https://en.m.wikipedia.org/wiki/File...n_drainage.png](https://en.m.wikipedia.org/wiki/File:...n_drainage.png)

Approximately 90 % of the water that is evaporated into the hydrologic cycle comes from the ocean. The ocean contains many forms of life uniquely adapted to survive in this habitat. Unfortunately, humans have degraded the oceans and their life through pollution, excess fishing, carbon dioxide acidification, and resource exploitation.

Storage and Flow

Almost 99% of the available fresh water is found below the surface as groundwater. This resource results from recycling of water in the hydrologic cycle. When precipitation falls, water travels on the surface and some can infiltrate into the ground. Groundwater is replenished when water moves from the surface, through unsaturated rocks or sediment (**unsaturated**), all the way down the **saturated parts** (saturated zone) in a process called infiltration and becomes groundwater (Figure 4.1.10). The top of the saturated portion is called the **water table**, which is the boundary between saturated and unsaturated zone.

Groundwater is found in **aquifers**, which are bodies of rock or sediment that store (and yield) large amounts of usable water in their pores. Aquifer productivity is controlled by **porosity** and **permeability**. Porosity is the percentage of open space in a rock or sediment body. Permeability is the ability of subsurface material to transmit fluids. Groundwater is found in the saturated zone of a rock body where all pores are filled with water. An important concept is that surface water always moves from higher elevation to lower elevation while groundwater always moves from higher energy (hydraulic head) to lower energy.



How ground water occurs in rocks.

of groundwater system showing the different components of an unconfined groundwater system:
<http://water.usgs.gov/edu/earthgwaquifer.html>

Figure 4.1.10: Model

Groundwater will continue to flow until it emerges as a spring, or discharges into surface water bodies on the land or in the ocean. To utilize groundwater, we drill holes (wells) into the ground and pump the water out.

Water Consumption

One of the most important natural resources is water. In some areas of the world, water sources are extremely limited or contaminated. If a geographical area lacks this resource, then agriculture and industry will have difficulty producing goods and foods. If water is contaminated, individuals who reside in these areas will be more likely to develop waterborne illnesses. Lack of water limits a consumer's ability to drink, eat, and bathe.

The table below illustrates who uses the majority of water sources. Be sure to note the differences between the United States and Global percentage applications.

Table 4.1.2: Water Use in the United States and Globally

Use	United States	Global
Agriculture	34%	70%
Domestic (drinking, bathing)	12%	10%
Industry	5%	20%
Power plant cooling	49%	small

Contributors and Attributions

- [Template:ContribCandelaGeo](#)
- [Elizabeth R. Gordon \(Furman University\)](#)
- Erin Avram, Cleveland State University

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4.2: Water Cycle Part 2

Read the information from the [USGS website on the Water Cycle](#) and use it to answer the questions below.

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4.3: EPA's Primary and Secondary Drinking Water Standards

Learning Objectives

- Summarize the accomplishments of the Safe Water Drinking Act (original law and amendments)
- Distinguish between primary and secondary contaminants.
- Classify a contaminant as organic, inorganic, a disinfectant, a disinfectant byproduct, radionuclide, or microbial.
- Distinguish between Maximum Contaminant Level (MCL) and Maximum Contaminant Level Goal (MCLG) drinking water standards.
- Recognize contaminant units: ppm, ppb, picoCuries, MFL, and millirems.
- Distinguish who monitors tap water, well water, and bottled water.

The Safe Drinking Water Act (1974)

Up until 1974, public drinking water supplies in the United States were monitored and regulated by state and local authorities. Lists of contaminants with their various concentrations could vary from state to state. As the chemical industry grew, these same state agencies noted the increased presence of existing and new organic chemicals in public water systems. In order to standardize drinking water across the country, the Environmental Protection Agency (EPA) enacted the Safe Drinking Water Act (SDWA) of 1974.



Figure 4.3.1: AFGE members and environmental activists participated in an event at the Washington, D.C. EPA headquarters to thank EPA workers for their work. (CC BY 2.0; [AFGE](#)).

The 1974 act enabled the EPA to monitor and regulate public water systems that serve over 25 people. Implementation and enforcement of drinking water standards would still be performed by each state. Regarding drinking water sources (surface and ground), the EPA and state agencies protect and monitor these as well. Levels of contaminants would be defined using the concentration terms **Maximum Contaminant Level (MCL)** and **Treatment Technique (TT)**.

The first set of drinking water standards included only 22 chemicals and/or pathogens. EPA established two major types of contaminants: primary and secondary. The first of these types (primary) contaminants are substances (examples could include Hg, As, and U) that can be toxic in small amounts. On the other hand, secondary contaminants are less toxic species (Fe and Zn) but are capable of causing cosmetic issues (color, taste, and odor) in drinking water.

All primary contaminants have enforceable concentration values. For the majority of these pollutants, EPA lists concentration limits by using the term **Maximum Contaminant Level (MCL)**. If a water supplier exceeds a given MCL for a toxin, then the EPA could impose fines and penalties. A few pathogens (*Giardia Lamblia* and *Legionella*) use Treatment Technique (TT) notation rather than numerical MCL concentrations. Water that contains any amount of these pathogens must be sanitized immediately with a standardized EPA procedure.

Table 4.3.1: Exert of the EPA's Primary Drinking Water Contaminant Listing

Primary Contaminant	Classification	MCLG ¹ (mg/L)	MCL or TT ¹ (mg/L)	Potential Health Effects from Long-Term Exposure Above the MCL (unless specified as short-term)	Sources of Contaminant in Drinking Water
<i>Giardia lamblia</i>	Microorganism	zero	TT ³	Gastrointestinal illness (such as diarrhea, vomiting, and cramps)	Human and animal fecal waste
Thallium	Inorganic	0.0005	0.002	Hair loss; changes in blood; kidney, intestine, or liver problems	Leaching from ore-processing sites; discharge from electronics, glass, and drug factories

In Table 4.3.1, a **Maximum Contaminant Level Goal (MCLG)** column is displayed for each of the two drinking water impurities. MCLG values are not enforced by the EPA. Instead, these concentrations are suggested values that water suppliers should strive to meet. Every six years the EPA reviews each primary contaminant with its MCL standard. During these times, they analyze data in regards to health risk assessment. If they choose to lower an MCL (smaller value) of a contaminant, then less health issues should occur. Unfortunately, reducing a concentration requires more technology which will cost the supplier and the consumer more money. By providing MCLG limits, the EPA encourages a water company to gradually work towards lowering a toxin's concentration.

Primary Drinking Water Contaminants

The National Primary Drinking Water Regulations (NPDWR) are legally enforceable primary standards and treatment techniques that apply to public water systems. Primary standards and treatment techniques protect public health by limiting the levels of contaminants in drinking water. A partial list of of standards for different water contaminants is given in the tables below. A more comprehensive and updated list can be found on this [link](#).

These types of toxins are classified into **one of the six** EPA classifications: microorganisms, disinfectants, disinfectant byproducts, inorganics, organics, or radionuclides. Primary contaminants are regulated because they have the capacity to do great harm to humans, plants, and animals. If a water distributor (must serve at least 200 homes) exceeds one of these mandated standards, then the EPA may impose a fine on the company.



Figure 4.3.2: Clark Hill Dam is located on the South Carolina/Georgia border. (CC BY; Elizabeth Gordon).

Concentration Units for Pollutants/Toxins

Typically, concentrations of primary contaminants are listed in parts per million. Metric amounts of a part per million are milligrams of the toxin per liter of the water sample.

$$\text{ppm} = \frac{\text{mg}}{\text{Liters}}$$

For extremely toxic substance, smaller units like parts per billion (ppb) might be used to express concentration. Metric amounts of a part per billion are micrograms of toxin per liter of water sample.

$$\text{ppb} = \frac{\mu\text{g}}{\text{Liters}}$$

The EPA measures a few pollutants using different concentration units. For example, asbestos based products are reported in fibers per liter (MFL). This silica material has been used in the construction and in the automotive industries. The insulating and noncombustible nature of asbestos enabled builders to make homes that were less flammable. In addition, mechanics can install asbestos-lined brakes that would not become flammable. Despite recent declines in use, the previous widespread use of asbestos has led to lung diseases like mesothelioma. In March of 2024, the EPA announced a final rule prohibiting the ongoing use of chrysotile asbestos. For more current information of EPA's stance on asbestos limitations, please click on this [link](#).



Figure 4.3.3: Corrugated asbestos roof. Image courtesy of de.Wikipedia.org/wiki/Benutzer:Hawedi

Other primary contaminants that utilize alternative concentration units are radionuclides and disinfectant byproducts. Radionuclides are materials that naturally emit radioactive particles or waves. Examples of these would include Ra-226 and Uranium. Both of these elements have unstable nuclei that produce ionizing radiation. Exposure to even tiny amounts of radionuclides can increase a person's likelihood of developing cancer. For this reason, the EPA regulates ionizing radiation using small concentrations (ppb). In addition, the EPA uses units like millirem and picocuries for some radionuclides. The unit millirem focuses on biological exposure to radiation. For information regarding the millirem and or Curie units, click on this [OER link](#). As for picocurie, this term defines an isotope's rate of nuclei decomposition.

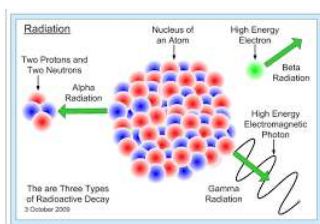


Figure 4.3.4: Types of ionizing radiation produced from a radio nuclide. Image courtesy of: PublicDomainPictures.net

Chemicals used to disinfect drinking water and rid it of pathogens are regulated as primary contaminants. These types of compounds will react with the water to form trace amounts of carcinogen organohalides. For this reason, the EPA regulated disinfectants using Maximum Residual Disinfectant Level Goal (MRDLG) and Maximum Residual Disinfectant Level (MRDL) rather than MCLG and MCL. Once again, the goal portion of this term is a future limit and the level is a fineable standard.

✓ Example 4.3.1

The toxicity of primary contaminants can be understood by looking at the MCL value. The lower the concentration, the more dangerous a contaminant is. Using the primary drinking water standards, determine which organic compound is the most toxic.

Solution

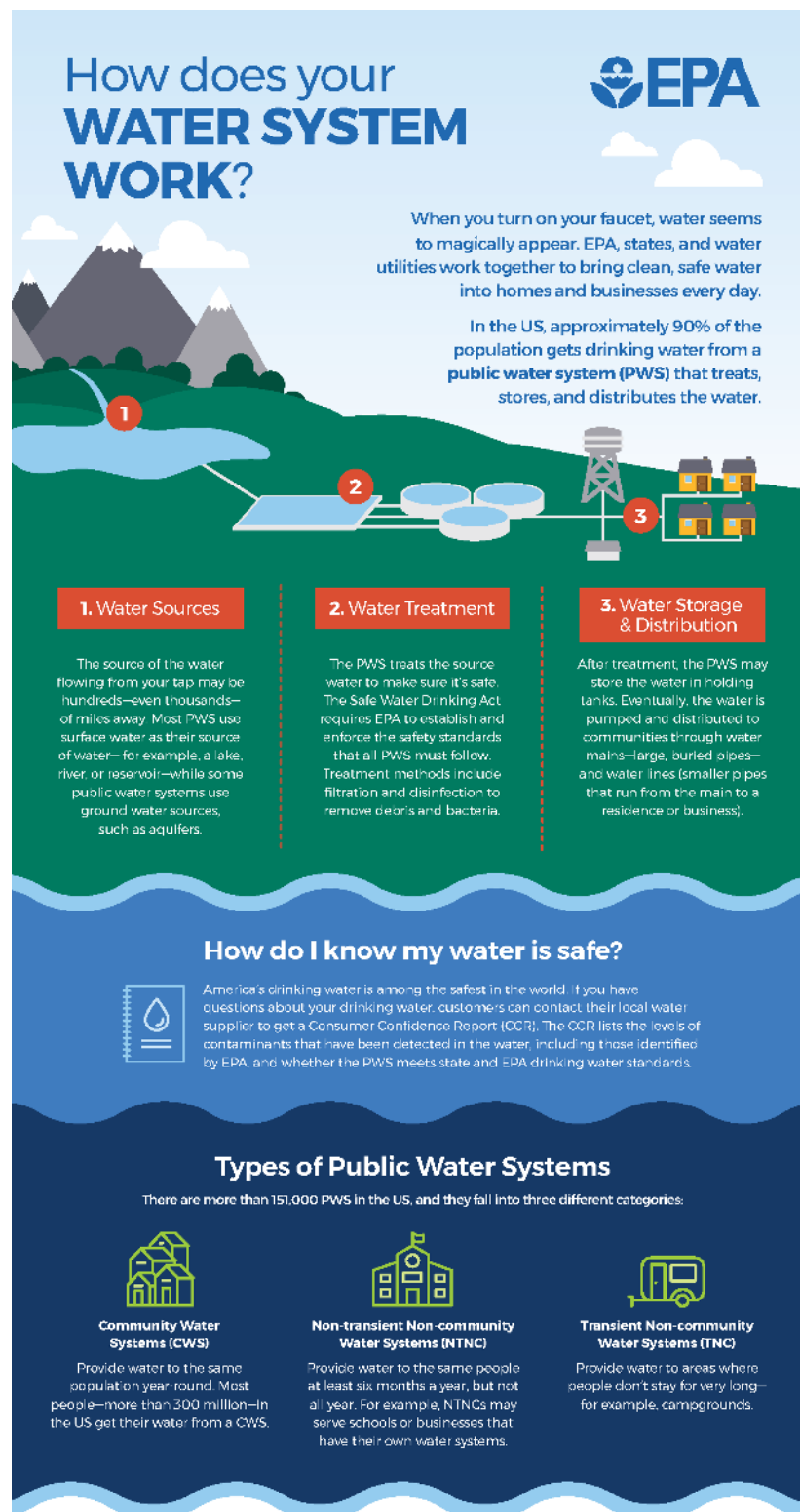
Dioxin is the most toxic organic compound listed on the primary drinking water standards..

Amendment SDWA of 1986

This change to the Safe Drinking Water Act established MCLGs and increased the total of regulated contaminants to 83. The EPA installed more monitoring devices to detect organic contaminants. Research was done to detect pathogens more effectively to reduce disease. Lastly, public notifications from water systems to the consumers would be announced if severe water issues occurred.

Amendment of SDWA of 1996

This particular legislation provided more protection and assessment of water sources (lakes, river, streams). In addition, water companies were required to provide consumers with a water quality report. States could seek federal money for upgrading their water quality processes. Cost-benefit analysis would be done to determine the risks and rewards of lowering the concentration of a contaminant. For small regulated water companies, more financial and technical assistance would be offered to help them maintain drinking water standards.



For more information, visit: [epa.gov/safewater](https://www.epa.gov/safewater)

Figure 4.3.5: Regulating drinking water in the United States. Image taken from: <https://www.google.com/url?sa=i&url=...AAAAAdAAAAABAD>

Non-regulated Water

The EPA does not regulate private wells, water sources that serve less than 25 people, or bottled water in the United States. Well water monitoring is the responsibility of the well owner. As for bottled water, the Food and Drug Administration (FDA) oversees this commodity by using EPA drinking water standards.

Notes (taken from 14.6-Delmar Link)

¹Definitions:

- Maximum Contaminant Level Goal (MCLG) - The level of a contaminant in drinking water below which there is no known or expected risk to health. MCLGs allow for a margin of safety and are non-enforceable public health goals.
- Maximum Contaminant Level (MCL) - The highest level of a contaminant that is allowed in drinking water. MCLs are set as close to MCLGs as feasible using the best available treatment technology and taking cost into consideration. MCLs are enforceable standards.
- Maximum Residual Disinfectant Level Goal (MRDLG) - The level of a drinking water disinfectant below which there is no known or expected risk to health. MRDLGs do not reflect the benefits of the use of disinfectants to control microbial contaminants.
- Treatment Technique (TT) - A required process intended to reduce the level of a contaminant in drinking water.
- Maximum Residual Disinfectant Level (MRDL) - The highest level of a disinfectant allowed in drinking water. There is convincing evidence that addition of a disinfectant is necessary for control of microbial contaminants.

² Units are in milligrams per liter (mg/L) unless otherwise noted. Milligrams per liter are equivalent to parts per million (PPM).

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4.4: Acid Rain and Water Hardness

Learning Objectives

- Explain how combustion processes affect water sources.
- Define the pH and the composition of acid rain (three specific formulas/names).
- Know the precursors of acid rain.
- Describe the environmental and biological effects of acid rain.
- Understand the differences between hard and soft water.

The US Safe Drinking Water Act defines the term "contaminant" as meaning any physical, chemical, biological, or radiological substance or matter in water. Therefore, the "contaminant" definition broadly applies to anything other than water molecules that are present. More information can be found at the [EPA Drinking Water Contaminant Website](https://www.epa.gov/dwqa). Drinking water may reasonably be expected to contain at least small amounts of some contaminants. Some drinking water contaminants may be harmful if consumed at certain levels in drinking water while others may be harmless. The presence of contaminants does not necessarily indicate that the water poses a health risk.

Air Pollutants

Water can become contaminated at any part of the water cycle. Air pollution can affect water vapor and liquid water. Combustion sources like vehicles and power plants generate compounds like NO_x , SO_x , CO , CO_2 , and other various inorganic and organic volatile organic compounds (VOC) species. Some of these compounds can become soluble in water. This could affect pH (or acidity) level of water surface water. Rainwater with a pH below 5.6 is considered to be acid rain.

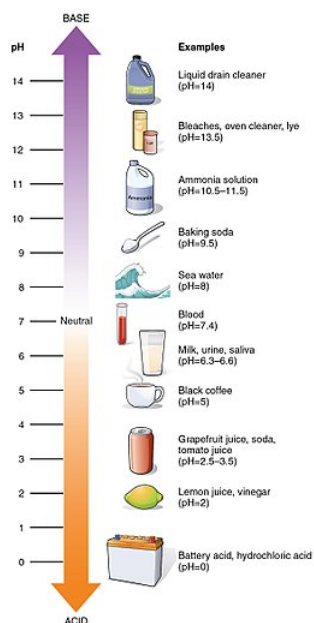


Figure 4.4.1: pH scale. (Copyright; OpenStax College via Anatomy & Physiology, Connexions Web site. <http://cnx.org/content/col11496/1.6/>)

Normally, water is neither acidic or basic, and has a neutral pH (pH= 7). When NO_x , SO_x , CO , CO_2 enter the water cycle, then the pH level is lowered below 7.0. If these gases are absorbed in rain clouds, then acid rain results. Specific acids involved in acid rain are **sulfuric acid (H_2SO_4)**, **nitric acid (HNO_3)**, and **carbonic acid (H_2CO_3)**. This environmental problem affects living organisms and building materials. Acid solutions can corrode metals and make them soluble as well.

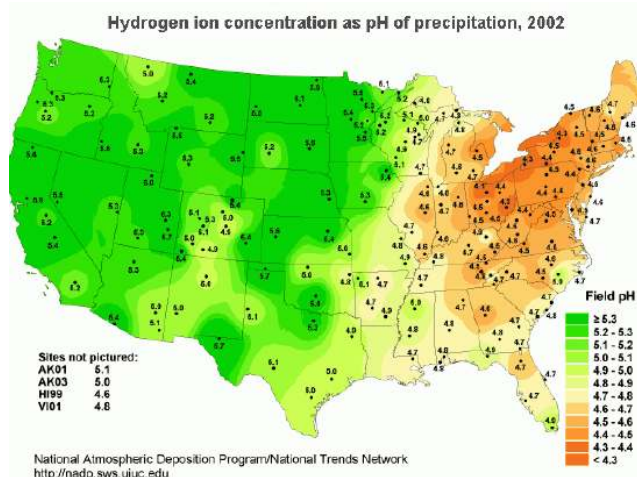
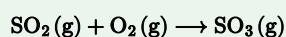


Figure 4.4.2: pH of rainwater over the continental United States.

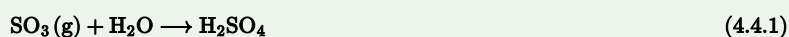
The acidity of rain is measured by collecting samples of rainwater and measuring its pH. To find the distribution of acid rain, weather conditions are monitored and rain samples collected at sites all over the country. The areas of greatest acidity (lowest pH values) are located in the Northeastern United States. This pattern of high acidity is caused by a large number of cities, the dense population, and the concentration of power and industrial plants in the Northeast. In addition, the prevailing wind direction brings storms and pollution to the Northeast from the Midwest, and dust from the soil and rocks in the Northeastern United States is less likely to neutralize acidity in the rain.

Acid rain from sulfuric acid

However, the majority of acid rain is accounted for by the presence of sulfuric acid (H_2SO_4).



Sulfur dioxide reacts with oxygen to form sulfur trioxide. Sulfur trioxide can then react with water to form sulfuric acid



Although sulfuric acid may be produced naturally in small quantities from biological decay and volcanic activity, almost all sulfur dioxide, that is subsequently converted to sulfuric acid, is produced by human activity, especially from the combustion of sulfur-containing fossil fuels in power plants.

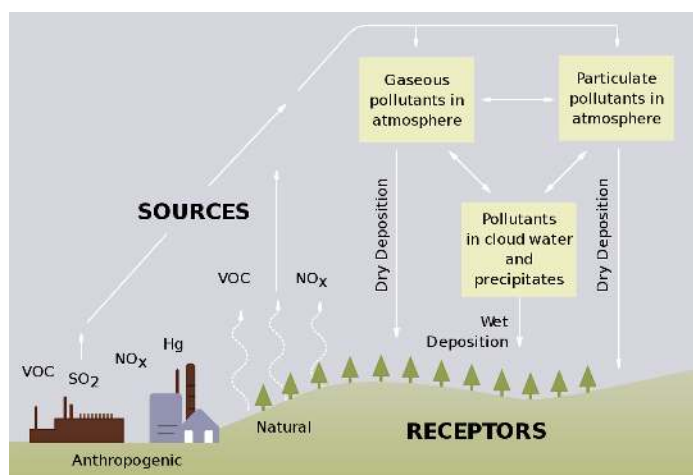
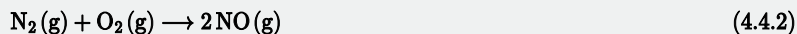


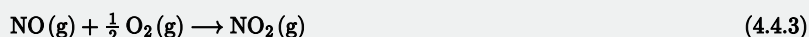
Figure 4.4.3: Processes involved in acid deposition (note that only SO_2 and NO_x play a significant role in acid rain). (public domain; EPA)

📌 Acid rain from nitric acid

Some acid rain is accounted for by nitric acid (HNO_3). This compound can originate from natural processes but is primarily released from high-temperature air combustion, such as the reactions in car engines and power plants. These processes produce large amounts of NO gas. This gas then forms nitric acid via several steps. Nitric oxide (NO) can also be formed during lightning storms by the reaction of nitrogen and oxygen



and NO is oxidized to nitrogen dioxide (NO_2)



$\text{NO}_2(\text{g})$ then reacts with water droplets to generate nitric acid (HNO_3)

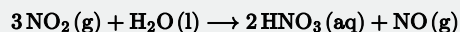


Table 4.4.2: Source of Water contamination in the United States

Combustion products	Sources
CO_2 and CO	Combustion of any material (any fuel or tree)
NO_x (NO_2 and NO)	High-temperature combustion of any fuel (gas, diesel, or coal), a product of lightning
SO_x (SO_2 and SO_3)	Combustion of sulfur-based fuels (diesel and coal), volcanic release
VOC (volatile organic compound)	Combustion of any carbon-based fuel (gas, diesel, or coal), fumes from paints or solvents

Hard Water

When water is described as being hard, this means it has a large concentration of minerals (this typically means metal ions). The **ions** that contribute to this problem are **calcium** (Ca^{2+}), **magnesium** (Mg^{2+}), **carbonate** (CO_3^{2-}), and **iron** ($\text{Fe}^{2+}/\text{Fe}^{3+}$). Hard water is not toxic, but can affect industrial processes. These ions will bond together with nonmetal ions or polyatomic anions to form insoluble salts that are called precipitates. The buildup of these newly formed solids is called scale. Precipitates can collect in boilers, cooling towers, and any equipment that employs water. This buildup can affect how the equipment works by damaging it internally.

📌 Video: Hard Water

Watch this video below to get an idea of what hard water is.



<https://youtu.be/Ts6PNYk7-a0>

1. What chemical species cause hard water?
2. How does hard water affect industry?
3. Hard water is known to contribute to the formation of kidney stones. How else can this type of water affect the body?

Soft Water

Soft water does not contain an excessive amount of alkaline minerals. Instead, **potassium** (K^+) and **sodium** (Na^+) ions are present. Soft water does not form scale but does produce excessive lathers with soap. Although industry and consumers prefer soft water, the sodium content can negatively affect human health if ingested at a high concentrations.

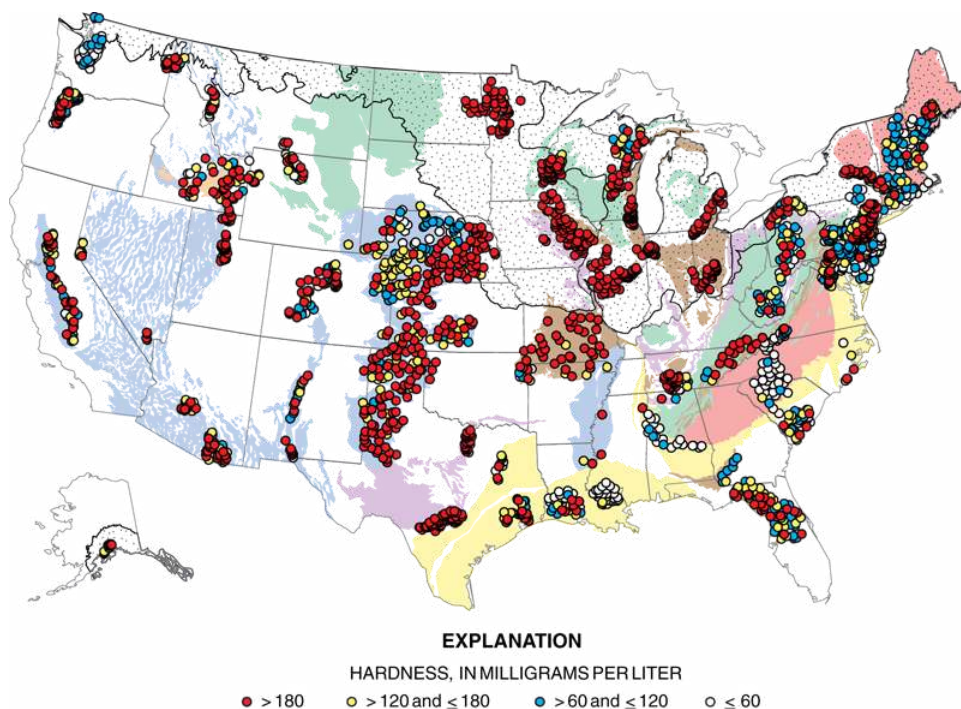


Figure 4.4.4: Hardness of groundwater from domestic wells, a USGS study. A study from the National Water-Quality Assessment (NAWQA) Program assessed water-quality conditions for about 2,100 domestic wells across the United States. (Public Domain; USGS).

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- 8.4: Chemical Contamination of Water by Elizabeth Gordon is licensed CC BY-NC-SA 4.0.

4.5: Sanitation of Drinking Water

Learning Objectives

- Explain why drinking water requires sanitation.
- Describe the three different types of water sanitation methods.
- Distinguish the advantages and disadvantages of each type of sanitation method.
- Define the terms residual protection and disinfection byproducts.

Chlorine-based methods

Chlorination is the process of adding chlorine compounds to drinking water to disinfect it and kill germs. In 1908, New Jersey was the first state to utilize chlorinated compounds for sanitizing drinking water. This method is quite cheap for most countries and rids water of bacteria and most viruses. Regarding protozoa (*Giardia* and *Cryptosporidium*), results can vary depending upon the type of chlorine product that it used and how long it is applied (refer to this CDC table). Chlorine does offer residual protection from treatment facility to consumer and is quite cheap. In the United States, chlorine (Cl_2) has been used to sanitize both drinking and waste waters since the early twentieth century. Unfortunately, gaseous chlorine (Cl_2) is known to be quite corrosive, poisonous, and explosive.



Video 4.5.1: Chlorine and Exposure Concerns. (Copyright; <https://www.youtube.com/@IAQMarketer>). (3:45min)

After watching the video above, answer the questions below:

- 1) Provide the physical and chemical properties of chlorine (Cl_2).
- 2) When was chlorine gas used as a chemical weapon?
- 3) What are some biological reactions to chlorine gas exposure?
- 4) Name the two household cleaners that should never be mixed together.

Other chlorine-based sanitizing agents include chlorine dioxide (ClO_2), sodium hypochlorite (NaClO), calcium hypochlorite ($\text{Ca}(\text{ClO})_2$), and chloramines (e.g., NH_2Cl). The first of these, chlorine dioxide, is a gaseous compound that is capable of destroying viruses and bacteria in drinking water, but requires higher dosages to kill protozoa like *Giardia*. Sanitizing procedures involving chlorine dioxide require longer contact time at higher dosages than those using pure chlorine. Hypochlorite (ClO^-) based compounds are effective at killing viruses and bacteria. Like chlorine dioxide, hypochlorites do not deactivate all protozoa. Deactivating *Cryptosporidium* can be challenging using chlorine based methods. In order to remove this protozoa, water must be flushed through a filter which has a pore size of equal to or less than 1 micron (micrometer). Filters should be labeled as "absolute 1 micron." If reverse osmosis is used to purify water, *Cryptosporidium* and all other disease causing pathogens will be removed as well.



Figure 4.5.1: Image of chlorine dioxide dissolved in water. (Copyright, <https://commons.wikimedia.org/wiki/User:Iridos>)

Chloroamines are a group of chemical compounds that contain chlorine (Cl_2) and ammonia (NH_3). The specific form of chloroamine used in drinking water disinfection is called monochloramine (NH_2Cl). These sanitizers do not kill harmful organisms as fast as chlorine gas and can decompose when exposed to heat, light, and bacteria. Effective chloramine levels range from 1.0 to 4.0 ppm. These compounds provide long-lasting residual protection. For this reason, monochloramine is often utilized as a secondary sanitizer in addition to chlorine or non-chlorine methods of sanitization.

All forms of chlorination do alter the smell and taste of drinking water. In addition, chlorine (particularly chloroamine compounds) can react with other chemicals in tap water to produce organochlorides. The EPA classifies these newly formed compounds as disinfectant byproducts and regulates them as primary drinking water contaminants.

Ozone (O_3) Disinfection

Ozone disinfection, or ozonation, is a powerful oxidizing process which is toxic to most waterborne organisms. This method of disinfection is widely used in Europe. Ozone deactivates cyst formation of protozoa (like *Giardia*). In addition, this chemical effectively kills viruses and bacteria.

Ozone machines were first constructed in the mid-1800's. New devices produce O_3 molecules by passing oxygen through ultraviolet light or a "cold" electrical discharge. To use ozone as a disinfectant, it must be created on-site and then added to the water by bubble contact. Unlike chlorine, the ozonation process does not alter the taste or smell of water.



Video 4.5.2: Demonstration-Effects of Panasonic's "Ozone Water" Device (1.11 min).

Unfortunately, ozone does not provide any residual protection once water has left the treatment facility. For water that must travel far distances to consumers, chlorine or chloramine is added throughout a distribution system to remove any potential pathogens.

during the transportation process.

Although fewer by-products are formed by ozonation, it has been discovered that ozone reacts with bromide ions in water to produce concentrations of a suspected carcinogen called bromate (BrO_3^-). Bromide (Br^-) can be found in fresh water supplies in sufficient concentrations to produce (after ozonation) more than 10 parts per billion (ppb) of bromate — the maximum contaminant level established by the EPA. Other disadvantages to using ozonation methods are energy requirements and cost. Of the three types of sanitization processes (chlorination, ozonation, and UV irradiation), this technique is the most expensive.

Ultraviolet Disinfection (UV)

Ultraviolet radiation (namely UVC) can be used to eradicate viruses, bacteria, and parasites from drinking water. This method of sanitation does not produce byproducts or alter the physical properties (namely taste or smell) of the water.



Video 4.5.3: Understanding Ultraviolet UV Radiation and its Effects. (Copyright, Larson Electronics). (1:54 min).

However, ionizing radiation has difficulty inactivating parasitic cysts in turbid (unclear) water. Like ozonation, UV irradiation offers no residual protection and is pricier than chlorination. Smaller ultraviolet units do supply over 2 million people in 28 developing countries with clean water for drinking, eating, and cooking.



Figure 4.5.2: Ultraviolet lamp sanitizing water. Image courtesy of: <https://www.nab.usace.army.mil/Media/News-Stories/Article/946068/army-water-treatment-plant-brings-a-unique-first/>

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- 1) <https://www.oxidationtech.com/ozone/history.html>
- 2) 14.6: Making Water Fit to Drink - Chemistry LibreTexts
- 3) https://r.search.yahoo.com/_ylt=Awri...6JpE5SOBYtIBc-

4) cdc.gov

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4.6: Flint Water Crisis

Read the article: [Flint, Michigan: Lethal Water](#) and answer the following questions.

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CHAPTER OVERVIEW

5: Nuclear Chemistry

- 5.1: The Discovery of Radiation
- 5.2: Types of Radiation
- 5.3: Ionizing Radiation and Non-ionizing Radiation
- 5.4: The Basics of Skin Cancer
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- 5.6: Applications of Nuclear Chemistry
 - 5.6.1: Using Nuclear Science to Diagnose Disease
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 - 5.6.4: Uses and Consequences of Nuclear Energy
 - 5.6.5: Three Mile Island Nuclear Accident
 - 5.6.6: Chernobyl Nuclear Disaster
 - 5.6.7: Fukushima Nuclear Disaster
 - 5.6.8: Nuclear Bombs

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5.1: The Discovery of Radiation

Learning Objectives

- Understand the difference between induced and natural radioactivity
- Know which elements have unstable nuclei.
- Correlate the relationship between binding energy and stability.

In [section 2.1](#) of this text, we learned how J. J. Thomson discovered the electron using cathode ray tubes (CRT). In November 1895, a German physicist named Wilhelm Roentgen used this same technology to discover the X-ray. Roentgen's ray originated from the CRT and reacted with a barium platinocyanide screen to fluoresce. After seeing this, Roentgen placed objects between this ray and the screen. He noted that the ray would still penetrate substances and leave an image on photographic film.

When naming this type of radiation, Roentgen reflected on his experiments. He could not confidently explain his observations. For this reason, he decided to call this new type of radiation the X-ray. Throughout his lifetime, he continued to explore this technology and even used his wife as a test subject (refer to Chapter 1 for her X-ray image). By never patenting this invention, Roentgen freely shared the X-ray with other researchers and medical professionals. In 1901, he was awarded the Nobel Prize in Physics for his work with the x-ray.

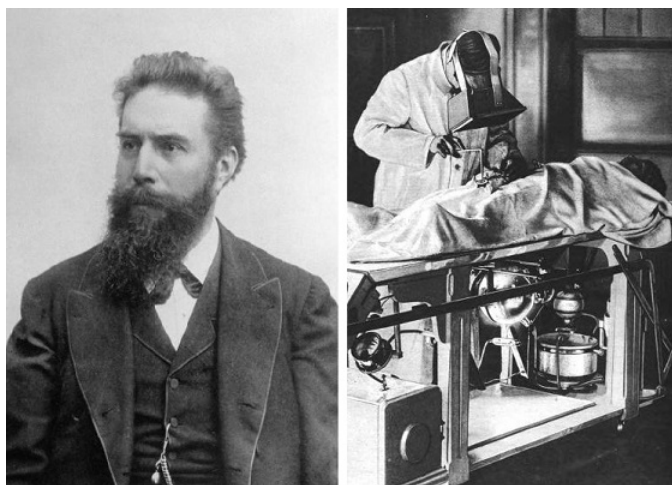


Figure 5.1.1: (left) Photo of Wilhelm Conrad Roentgen, the discoverer of x-rays. (right) Surgical operation during World War I using a fluoroscope that used X-rays to find embedded bullets

As with any new technology, people looked for ways to apply the X-ray to day-to-day life. Thomas Edison, the inventor of the light bulb, thought the average person should have an X-ray machine in their home. He designed X-ray machines to be smaller and more portable. Unfortunately, many of his technicians died from radiation poisoning.

During Edison's time, people would host X-ray parties in their homes. The host of these gatherings would allow guests to x-ray different parts of their body. People would leave these events with their framed souvenirs. X-rays were even used to measure foot size. Devices like the fluoroscope were placed in shoe stores. Technicians, consumers, and observers could look in oculars while the foot was being x-rayed. Unnecessary exposure to radiation was not considered a hazard during these times. In this era, most believed this was the most accurate way to measure the foot.

Unfortunately, the dangers of exposure to X-rays were not known at the time. One of Thomas Edison's colleagues, Clarence Madison Dally, unknowingly gave his life to assist in the testing of X-rays. From repeated exposure to X-rays through his work with Edison, Dally developed lesions on his hands, requiring the amputation of his fingers and then eventually of his arms. After 8 years of experimenting with X-rays, Dally died due to the damage the exposure inflicted on his body. More information about this work and Dally's life can be found [in this article](#) in the Smithsonian Magazine.



Video 5.1.1: The shoe fitting fluoroscope was a common fixture in shoe stores during the 1930s, 1940s and 1950s. The first fluoroscopic device for x-raying feet may have been created during World War I to eliminating the need for patients to remove their boots, to speed up the processing of the large number of injured military personnel who were seeking help. After the war, the device was modified for shoe-fitting and showcased at a shoe retailers convention in Boston in 1920.

Radioactivity

When Becquerel heard about Roentgen's discovery, he wondered if his fluorescent minerals would give the same X-rays. Becquerel placed some of his rock crystals on top of a well-covered photographic plate and sat them in the sunlight. The sunlight made the crystals glow with a bright fluorescent light, but when Becquerel developed the film he was very disappointed. He found that only one of his minerals, a uranium salt, had fogged the photographic plate. He decided to try again, and this time, to leave them out in the sun for a longer period of time. Fortunately, the weather did not cooperate and Becquerel had to leave the crystals and film stored in a drawer for several cloudy days. Before continuing his experiments, Becquerel decided to check one of the photographic plates to make sure the chemicals were still good. To his amazement, he found that the plate had been exposed in spots where it had been near the uranium-containing rocks and some of these rocks had not been exposed to sunlight at all. In later experiments, Becquerel confirmed that the radiation from the uranium had no connection with light or fluorescence, but the amount of radiation was directly proportional to the concentration of uranium in the rock. Becquerel had discovered **radioactivity**.



Figure 5.1.3: Image of Becquerel's photographic plate which has been fogged by exposure to radiation from a uranium salt. The shadow of a metal Maltese Cross placed between the plate and the uranium salt is clearly visible (Public Domain).

The Curies and Radium

One of Becquerel's assistants, a young Polish scientist named Maria Sklowdowska (to become Marie Curie after she married Pierre Curie), became interested in the phenomenon of radioactivity. With her husband, she decided to find out if chemicals other than uranium were radioactive. The Austrian government was happy to send the Curies a ton of pitchblende from the mining region of Joachimsthal because it was waste material that had to be disposed of anyway. The Curies wanted the pitchblende because it was the residue of uranium mining. From the ton of pitchblende, the Curies separated **0.10 g** of a previously unknown element, radium, in the form of the compound radium chloride. This radium was many times more radioactive than uranium.

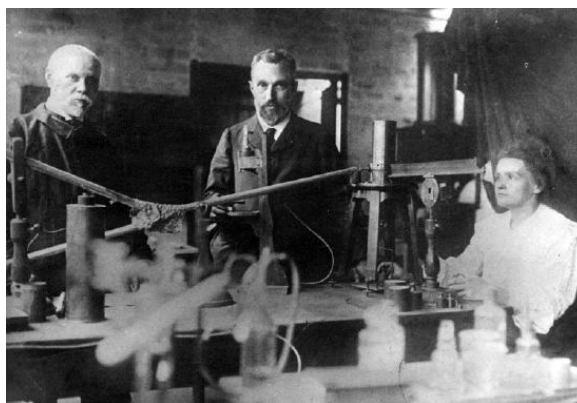


Figure 5.1.4: Marie (right) and Pierre Curie (middle) with Henri Becquerel (left) shared the 1903 Nobel Prize

By 1902, the world was aware of a new phenomenon called radioactivity and of new elements that exhibited natural radioactivity. For this work, Becquerel and the Curies shared the 1903 Nobel Prize in Physics. These three researchers continued to work with hazardous radioactive materials. They experienced various ailments including burns and weakness by touching these substances. In 1906, Pierre Curie accidentally fell under a horse-drawn cart. He immediately died from his wounds and left Marie with two small children (Irene, 9 years old and Eve, 2 years old). After his death, Marie continued her research by identifying a new element, Polonium. In 1911, she was awarded the Nobel Prize in Chemistry for the discoveries of radium and polonium. As of today, Marie Curie is one of only four people to have received two Nobel Prizes in the sciences. To learn more about this famous female scientist, click on this [link](#) to watch a short video biography of her life.



Figure 5.1.2: Everyone knows that winning the Nobel Prize is a big deal, but why do we even have a Nobel Prize? And why does it matter?

Further experiments provided information about the characteristics of the penetrating emissions from radioactive substances. It was soon discovered that there were three common types of radioactive emissions. Some of the radiation could pass easily through aluminum foil while the foil stopped some of the radiation. Some radiation could even pass through foil up to a centimeter thick. The three basic types of radiation were named alpha, beta, and gamma radiation.

Eventually, scientists were able to demonstrate experimentally that the alpha particle, α , was a helium nucleus (a particle containing two protons and two neutrons), a beta particle, β , was a high-speed electron, and gamma rays, γ , were a very high energy form of light (even higher energy than x-rays).

Unstable Nuclei May Disintegrate

A nucleus (with one exception, hydrogen-1) consists of some number of protons and neutrons pulled together in an extremely tiny volume. Since protons are positively charged and like charges repel, it is clear that protons cannot remain together in the nucleus unless there is a powerful force holding them there. The force that holds the nucleus together is generated by **nuclear-binding energy**.

A nucleus with a large amount of binding energy per nucleon (proton or neutron) will be held together tightly and is referred to as stable. These nuclei do not break apart. When there is too little binding energy per nucleon, the nucleus will be less stable and may disintegrate (come apart). Such disintegrations are referred to as **natural radioactivity**. It is also possible for scientists to smash nuclear particles together and cause nuclear reactions between normally stable nuclei. These disintegrations are referred to as **artificial radioactivity**. None of the elements above #92 on the periodic table occur on earth naturally; they are all products of artificial radioactivity (man-made).

When nuclei come apart, they come apart violently accompanied by a tremendous release of energy in the form of heat, light, and radiation. This energy comes from some of the nuclear binding energy. In nuclear changes, the energy involved comes from the nuclear binding energy. However, in chemical reactions, the energy comes from electrons moving energy levels. A typical nuclear change (such as fission) may involve millions of times more energy per atom changing compared to a chemical change (such as burning)!

Contributors and Attributions

- Wikibooks
- CK-12
- Elizabeth Gordon (Furman University)
- Emma Gibney (Furman University)
- Erin Avram (Cleveland State University)

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5.2: Types of Radiation

Learning Objectives

- Use A/Z formats to depict alpha, beta, positron, neutron, gamma, and the three isotopes of hydrogen.
- Express the changes in the atomic number and mass number of radioactive nuclei when a particle or ray is emitted.
- Convert symbol-mass format of elements to A/Z format and then balance a nuclear reaction.
- Generate a decay reaction if provided a symbol-mass format of an element.

Many nuclei are radioactive; that is, they decompose by emitting particles or rays and in doing so, become a different nucleus. In our discussions up to this point, atoms in a molecule could rearrange, but could not be converted to a different element. This is because we have focused on physical and chemical changes, where the change involves the state of matter (physical) or electron configuration (chemical). Now we will introduce situations that involve changes to the atomic nucleus, reactions that involve changes to the number of protons or neutrons (collectively called nucleons). All nuclei with 84 or more protons are radioactive and elements with less than 84 protons have both stable and unstable isotopes. All of these elements can go through nuclear changes and turning into different elements.

In natural radioactive decay, three common emissions occur. When these emissions were originally observed, scientists were unable to identify them as some already known particles and so named them

- alpha particles (α),
- beta particles, (β), and
- gamma rays (γ)

using the first three letters of the Greek alphabet. At some later time, alpha particles were identified as helium-4 nuclei, beta particles were identified as electrons, and gamma rays as a form of electromagnetic radiation like X-rays except much higher in energy and even more dangerous to living systems.

Name	Symbol(s)	Representation	Description
Alpha particle	${}^4_2\text{He}$ or ${}^4_2\alpha$		(High-energy) helium nuclei consisting of two protons and two neutrons
Beta particle	${}^0_{-1}\text{e}$ or ${}^0_{-1}\beta$		(High-energy) electrons
Positron	${}^0_{+1}\text{e}$ or ${}^0_{+1}\beta$		Particles with the same mass as an electron but with 1 unit of positive charge
Proton	${}^1_1\text{H}$ or ${}^1_1\text{p}$		Nuclei of hydrogen atoms
Neutron	${}^1_0\text{n}$		Particles with a mass approximately equal to that of a proton but with no charge
Gamma ray	γ		Very high-energy electromagnetic radiation

Figure 5.2.1: Although many species are encountered in nuclear reactions, this table summarizes the names, symbols, representations, and descriptions of the most common of these. (CC BY 4.0; OpenStax)

Alpha Decay

The nuclear disintegration process that emits alpha particles is called alpha decay. An example of a nucleus that undergoes alpha decay is uranium-238. The alpha decay of U-238 is



In this nuclear change, the uranium atom (${}^{238}_{92}\text{U}$) transmuted into an atom of thorium (${}^{234}_{90}\text{Th}$) and, in the process, gave off an alpha particle. Look at the symbol for the alpha particle: ${}^4_2\text{He}$. Where does an alpha particle get this symbol? The bottom number in a

nuclear symbol is the number of protons. That means that the alpha particle has two protons that were lost by the uranium atom. The two protons also have a charge of **+2**. The top number, 4, is the mass number or the total of the protons and neutrons in the particle. Because it has 2 protons, and a total of 4 protons and neutrons, alpha particles must also have two neutrons. Alpha particles always have this same composition: two protons and two neutrons.

Another alpha particle producer is thorium-230.

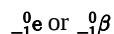


Alpha decays occur with radioactive isotopes of radium, radon, uranium, and thorium.

Beta Decay

Another common decay process is beta particle emission, or beta decay. A beta particle is simply a high-energy electron that is emitted from the nucleus. It may occur to you that we have a logically difficult situation here. Nuclei do not contain electrons, and yet during beta decay, an electron is emitted from a nucleus. At the same time that the electron is being ejected from the nucleus, a neutron is becoming a proton. It is tempting to picture this as a neutron breaking into two pieces with the pieces being a proton and an electron. That would be convenient for simplicity, but unfortunately that is not what happens. For convenience sake, though, we will treat beta decay as a neutron splitting into a proton and an electron. The proton stays in the nucleus, increasing the atomic number of the atom by one. The electron is ejected from the nucleus and is the particle of radiation called beta.

To insert an electron into a nuclear equation and have the numbers add up properly, an atomic number and a mass number had to be assigned to an electron. The mass number assigned to an electron is zero (0) which is reasonable since the mass number is the number of protons plus neutrons and an electron contains no protons and no neutrons. The atomic number assigned to an electron is negative one (-1), because that allows a nuclear equation containing an electron to balance atomic numbers. Therefore, the nuclear symbol representing an electron (beta particle) is



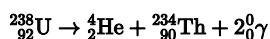
Thorium-234 is a nucleus that undergoes beta decay. Here is the nuclear equation for this beta decay.



Beta decays are common with Sr-90, C-14, H-3 and S-35.

Gamma Radiation

Frequently, gamma ray production accompanies nuclear reactions of all types. In the alpha decay of U-238, two gamma rays of different energies are emitted in addition to the alpha particle.



Virtually all of the nuclear reactions in this chapter also emit gamma rays, but for simplicity the gamma rays are generally not shown. Nuclear reactions produce a great deal more energy than chemical reactions. Chemical reactions release the difference between the chemical bond energy of the reactants and products, and the energies released have an order of magnitude of 1×10^3 **kJ/mol**. Nuclear reactions release some of the binding energy and may convert tiny amounts of matter into energy. The energy released in a nuclear reaction has an order of magnitude of 1×10^{18} **kJ/mol**. That means that nuclear changes involve almost **a million times more energy** per atom than chemical changes!

Common gamma emitters would include I-131, Cs-137, Co-60, and Tc-99.

Virtually all of the nuclear reactions in this chapter also emit gamma rays, but for simplicity the gamma rays are generally not shown.

The essential features of each reaction are shown in Figure 5.2.2.

Decay Type	Radiation Emitted	Generic Equation	Model
Alpha decay	${}^4_2\alpha$	${}_Z^AX \longrightarrow {}_{Z-2}^{A-4}X' + {}^4_2\alpha$	 Parent → Daughter + Alpha Particle
Beta decay	${}^0_{-1}\beta$	${}_Z^AX \longrightarrow {}_{Z+1}^AX' + {}^0_{-1}\beta$	 Parent → Daughter + Beta Particle
Gamma emission	${}^0_0\gamma$	${}_Z^AX^* \xrightarrow{\text{Relaxation}} {}_Z^AX' + {}^0_0\gamma$	 Parent (excited nuclear state) → Daughter + Gamma ray

Figure 5.2.2: Three most common modes of nuclear decay

"Nuclear Accounting"

When writing nuclear equations, there are some general rules that will help you:

- The sum of the mass numbers (top numbers) on the reactant side equal the sum of the mass numbers on the product side.
- The atomic numbers (bottom numbers) on the two sides of the reaction will also be equal.

In the alpha decay of ${}^{238}\text{U}$ (Equation 5.2.1), both atomic and mass numbers are conserved:

- mass number: $238 = 4 + 234$
- atomic number: $92 = 2 + 90$

Confirm that this equation is correctly balanced by adding up the reactants' and products' atomic and mass numbers. Also, note that because this was an alpha reaction, one of the products is the alpha particle, ${}^4_2\text{He}$.

Note that both the mass numbers and the atomic numbers add up properly for the beta decay of Thorium-234 (Equation 5.2.3):

- mass number: $234 = 0 + 234$
- atomic number: $90 = -1 + 91$

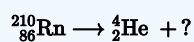
The mass numbers of the original nucleus and the new nucleus are the same because a neutron has been lost, but a proton has been gained and so the sum of protons plus neutrons remains the same. The atomic number in the process has been increased by one since the new nucleus has one more proton than the original nucleus. In this beta decay, a thorium-234 nucleus has one more proton than the original nucleus. In this beta decay, a thorium-234 nucleus has become a protactinium-234 nucleus. Protactinium-234 is also a beta emitter and produces uranium-234.



Once again, the atomic number increases by one and the mass number remains the same; confirm that the equation is correctly balanced. Before starting the problems, below preview this video of your instructor balancing nuclear reactions.

✓ Example 5.2.1

Complete the following nuclear reaction by filling in the missing particle.



Solution

This reaction is an alpha decay. We can solve this problem one of two ways:

Solution 1: When an atom gives off an alpha particle, its atomic number drops by 2 and its mass number drops by 4 leaving: $^{206}_{84}\text{Po}$. We know the symbol is **Po**, for polonium, because this is the element with 84 protons on the periodic table.

Solution 2: Remember that the mass numbers on each side must total up to the same amount. The same is true of the atomic numbers.

- Mass numbers: $210 = 4 + ?$
- Atomic numbers: $86 = 2 + ?$

We are left with $^{206}_{84}\text{Po}$

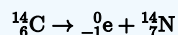
✓ Example 5.2.2

Write each of the following nuclear reactions.

- Carbon-14, used in carbon dating, decays by beta emission.
- Uranium-238 decays by alpha emission.

Solution

a) Beta particles have the symbol $^0_{-1}\text{e}$. Emitting a beta particle causes the atomic number to increase by 1 and the mass number to not change. We get atomic numbers and symbols for elements using our periodic table. We are left with the following reaction:



b) Alpha particles have the symbol ^4_2He . Emitting an alpha particle causes the atomic number to decrease by 2 and the mass number to decrease by 4. We are left with:



Decay Series

The decay of a radioactive nucleus is a move toward becoming stable. Often, a radioactive nucleus cannot reach a stable state through a single decay. In such cases, a series of decays will occur until a stable nucleus is formed. The decay of **U-238** is an example of this. The **U-238** decay series starts with **U-238** and goes through fourteen separate decays to finally reach a stable nucleus, **Pb-206** (Figure 5.2.3). There are similar decay series for **U-235** and **Th-232**. The **U-235** series ends with **Pb-207** and the **Th-232** series ends with **Pb-208**.

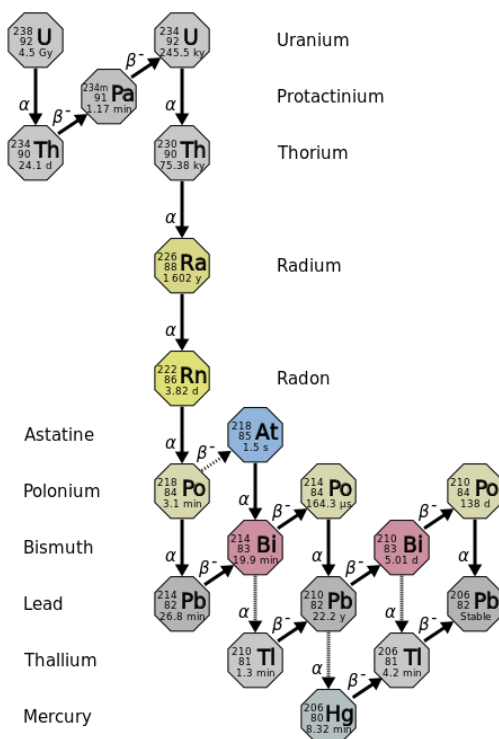


Figure 5.2.3: Uranium-238 Decay chain. (CC-BY-3.0 Tosaka)

Several of the radioactive nuclei that are found in nature are present there because they are produced in one of the radioactive decay series. For example, there may have been radon on the earth at the time of its formation, but that original radon would have all decayed by this time. The radon that is present now is present because it was formed in a decay series (mostly by U-238).

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5.3: Ionizing Radiation and Non-ionizing Radiation

Learning Objectives

- Qualitatively compare the ionizing and penetration power of alpha particles (α), beta particles (β), and gamma rays (γ).
- Describe the biological impact of ionizing radiation.
- Know the most common source of background radiation and how to minimize exposure to this source.

The Ionizing and Penetration Power of Radiation

The increased use of radioisotopes has led to increased concerns over the effects of these materials on biological systems (such as humans). All radioactive nuclides emit high-energy particles or electromagnetic waves. When this radiation encounters living cells, it can cause increase heat, break chemical bonds, or ionize molecules. The most serious biological damage results when these radioactive emissions fragment or ionize molecules. For example, alpha and beta particles emitted from nuclear decay reactions possess much higher energies than ordinary chemical bond energies. When these particles strike and penetrate matter, they produce ions and molecular fragments that are extremely reactive. The damage this does to biomolecules in living organisms can cause serious malfunctions in normal cell processes, taxing the organism's repair mechanisms and possibly causing illness or even death (Figure 5.3.1).

The ability of radiation to damage molecules is analyzed in terms of **ionizing power**. When a radiation particle interacts with atoms, the interaction can cause the atom to lose electrons and thus become ionized. The greater the likelihood that damage will occur by an interaction means the radiation has greater ionizing power. Ionizing radiation could affect either the whole body (somatic damage) and/or eggs and sperm (genetic damage).

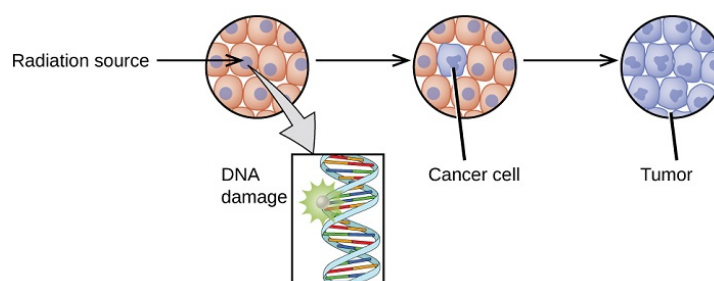


Figure 5.3.1: Radiation can harm biological systems by damaging the DNA of cells. If this damage is not properly repaired, the cells may divide in an uncontrolled manner and cause cancer.

Much of the threat from radiation relates to the ability to protect oneself from the emitted radiation. The ability of each type of radiation to pass through matter is expressed in terms of **penetration power**. The more material the radiation can pass through and the easier the radiation passes through, the greater the penetration power and the more dangerous they are. In general, the greater the mass of the particle, the greater the ionizing power but the lower the penetration power.

When comparing the most common forms of ionizing radiation (alpha, beta, and gamma), alpha particles have the greatest mass. Alpha particles have approximately four times the mass of a proton or neutron and approximately ~8,000 times the mass of a beta particle (Figure 5.3.1). Because of the large mass of the alpha particle, it has the highest ionizing power and the greatest ability to damage tissue. That same large size of alpha particles, however, makes them less able to penetrate matter. They collide with molecules very quickly when striking matter. This collision often results in the addition of two electrons, causing the alpha particle to become a harmless helium atom. Due to their relatively heavy mass, alpha particles have the least penetration power and can be stopped by a thick sheet of paper or even a layer of clothes. They are also stopped by the outer layer of dead skin on people. This may seem to remove the threat from alpha particles but only from external sources. In a situation like a nuclear explosion or a nuclear accident where radioactive emitters are spread around in the environment, the emitters can be inhaled or taken in with food or water. If the alpha emitter enters the body, you no longer have protection from the particles and bodily harm can result.

Figure 5.3.1: Properties of the three Primary Radiation Types

Particle	Symbol	Mass	Penetrating Power	Ionizing Power	Shielding
alpha	α	4 amu	Very Low	Very High	Paper and Skin
beta	β	1/1837 amu	Intermediate	Intermediate	Aluminum
gamma	γ	0 (energy only)	Very High	Very Low	2 inches lead

Beta particles are much smaller than alpha particles, and therefore, they have much less ionizing power (less ability to damage tissue), but their small size gives them much greater penetration power. Most resources say that beta particles can be stopped by a one-quarter inch thick sheet of aluminum. Once again, however, the greatest danger occurs when the beta-emitting source gets inside your body.

Gamma rays are not particles, but a high energy form of electromagnetic radiation (like X-rays, except more powerful). Gamma rays are energy with no mass or charge. Gamma rays have tremendous penetration power and require several inches of dense material (like lead) to shield them. It is possible for gamma rays to pass completely through a human body without striking anything. They are considered to have the least ionizing power but the greatest penetration power.

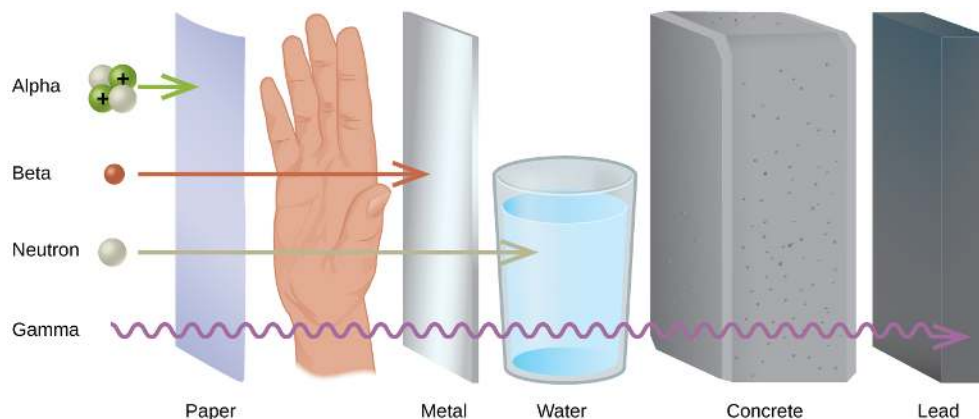


Figure 5.3.2: The ability of different types of radiation to pass through material is shown. From least to most penetrating, they are alpha < beta < neutron < gamma. (CC BY-SA, OpenStax).

The safest amount of radiation to the human body is zero. Unfortunately, it is not possible to completely avoid exposure to ionizing radiation. The next best goal is to minimize your exposure by limiting the length of time of your exposure and by increasing your distance from the source. The image below summarizes the key concepts of ionization and penetration abilities of alpha, beta, and gamma radiation.

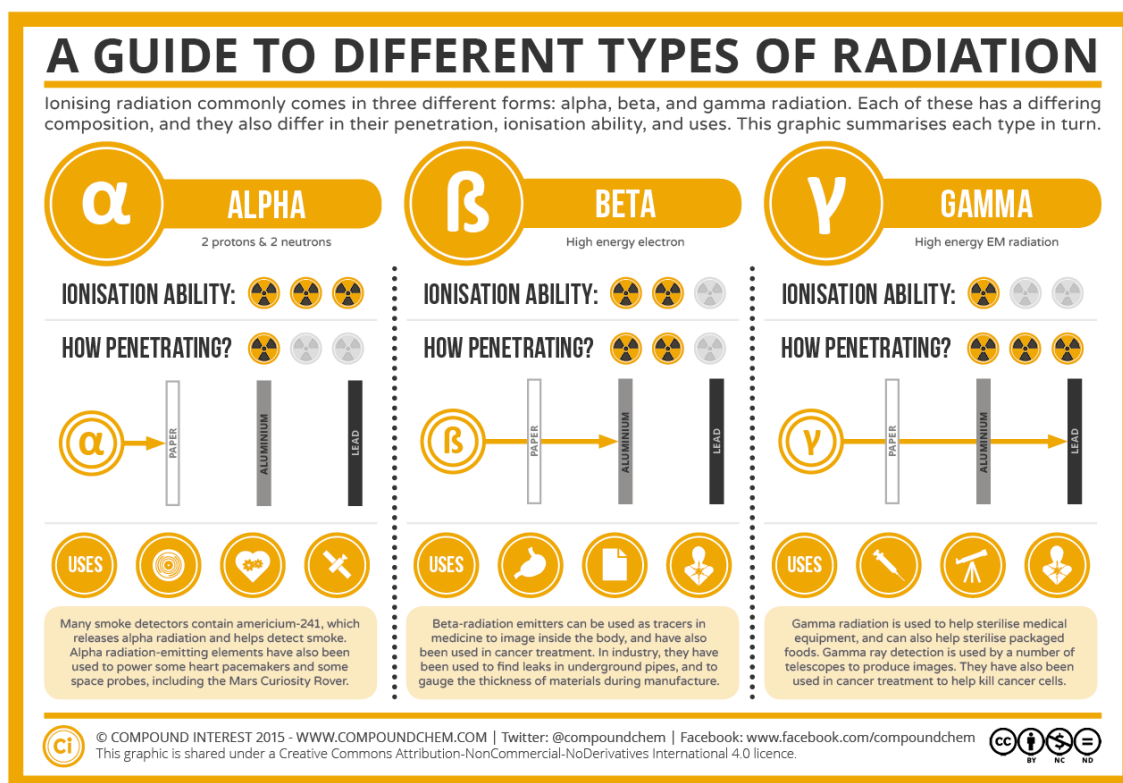


Figure 5.3.3: Ionization and penetrations abilities of alpha, beta, and gamma radiation. (CC BY-NC-ND, CompoundChem.com).

Nonionizing Radiation

There is a large difference in the magnitude of the biological effects of nonionizing radiation (for example, light and microwaves) and ionizing radiation, emissions energetic enough to knock electrons out of molecules, for example, α and β particles, γ rays, X-rays, and high-energy ultraviolet radiation (Figure 5.3.2).

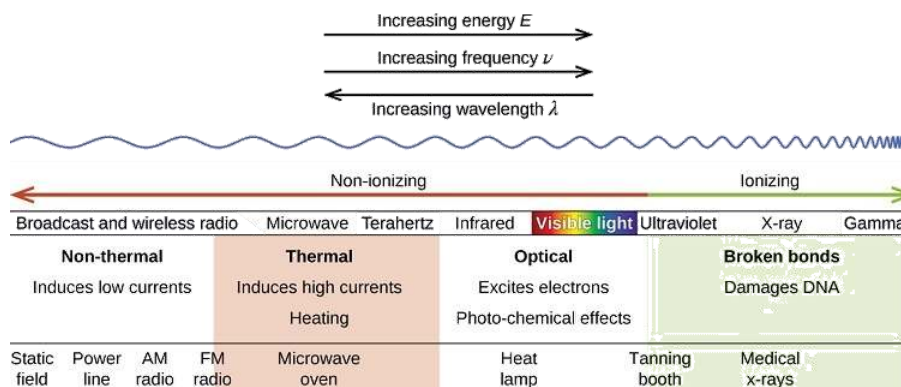


Figure 5.3.4: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing. (CC BY-SA, OpenStax).

Energy absorbed from nonionizing radiation speeds up the movement of atoms and molecules, which is equivalent to heating the sample. Although biological systems are sensitive to heat (as we might know from touching a hot stove or spending a day outside on a hot day), a large amount of nonionizing radiation is tolerated before negative effects begin. Forms of nonionizing radiation include wave-like radiation shown on the left side of the image. This type of radiation would include the visible spectrum through radio waves.

Everyday Exposure to Radiation

Natural radiation provides the majority of exposure to the average person. Looking at the pie chart below, the largest sources of exposure to radiation is from radon gas (Rn-222). This isotope is an alpha emitter with a half-life of 3.82 days. Radon is produced through the radioactive decay of U-238, which is found in trace amounts in soil and rocks. In the environment, radon concentrations can vary depending on geographical location. Once the soil of a particular region is disturbed, this element can escape and cause serious health issues. Please be aware that radon, not nuclear reactors, affects more people in the United States.

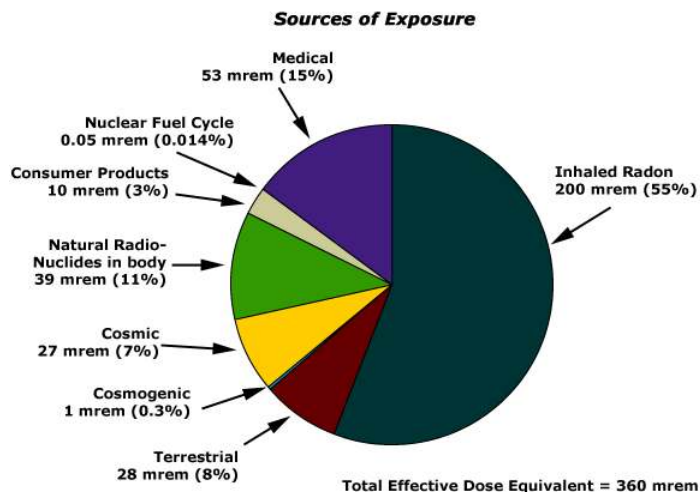


Figure 5.3.5: Sources of exposure to radiation. from [Nuclear ScienceNational](#) of Lawrence Berkeley National Lab and with [CPEP](#) (the Contemporary Physics Education Project).

Radon gas escapes from the ground and gradually seeps into homes and other structures situated above. Since radon gas is about eight times more dense than air, it accumulates in basements and lower floors, slowly diffusing throughout buildings (Figure 5.3.5). Once airborne, the radon enters the body through inhalation or ingestion. Through alpha emission, Rn-222 decays to produce large particles. These alpha particles then travel to the respiratory tract where they can ionize lung tissue. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

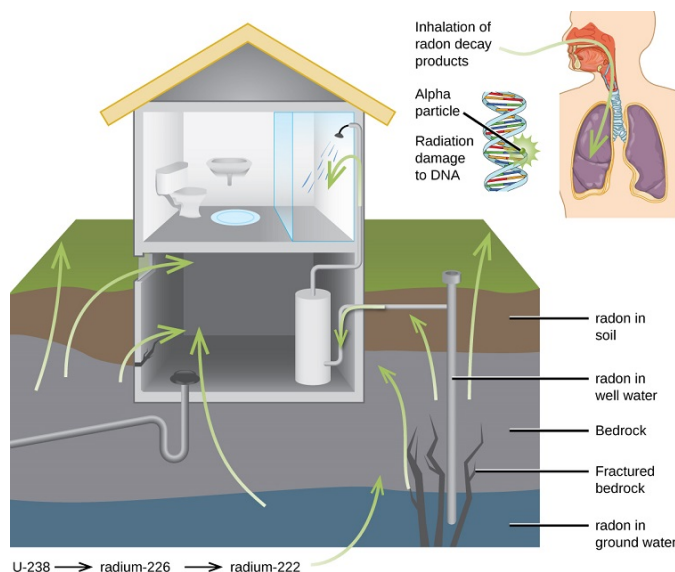


Figure 5.3.6: Radon-222 seeps into houses and other buildings from rocks that contain uranium-238, a radon emitter. The radon enters through cracks in concrete foundations and basement floors, stone or porous cinder block foundations, and openings for water and gas pipes.

Radon is found in buildings across the country, with amounts varying depending on where you live. The average concentration of radon inside houses in the United States (1.25 pCi/L) is about three times the levels found in outside air. About one in six houses have radon levels high enough to warrant remediation efforts to reduce the radon concentration. The Environmental Protection Agency (EPA) sets regulations and monitors for radon throughout the United States. Regions are classified into three different zones using color-coded keys (Zone 1 (red): highest potential for radon; Zone 2 (orange): moderate potential for radon; Zone 3 (yellow): low potential for radon). Based on the zone for your region, an individual can assess their probability of being exposed to this deadly isotope and determine if they should have their home tested for radon. If the radon level in a home is greater than 4 pCi/L (picocuries per liter), the EPA recommends action is taken to reduce the radon in the home. Click on this [link](#) to see your state and county's radon-risk level.

Radon tests can be purchased through the EPA or at most hardware stores. If a consumer finds excessive radon in their home, then alterations to the existing construction can be made. Some of these changes include installing barriers between the soil and the home and placing ventilation on the ground floor to release radon outside. For more information regarding these types of alterations, click on this [link](#) and select "how to fix your home."

Other sources of radiation in consumer products include smoke detectors, antique watches/clocks, and older ceramics/glass. Smoke detectors emit alpha particles of americium-241. Antique watches and clocks use hydrogen-3 (tritium), promethium-147, or radium-226 as a fluorescent light source. Antique ceramics (dating before 1970) could contain uranium-238 if the color of the ceramic is an orangey-red. If antique glassware is yellow or green in color, it could contain uranium-238 as well. Glassware with this isotope will glow under a black light.



Figure 5.3.7: Vaseline glass contains natural or depleted uranium. Image courtesy of Elizabeth R. Gordon

Concrete, fertilizers, kitty litter, and even food (bananas and salt substitutes) can contain trace amounts of different types of radioactivity. Potassium-40 is present in all of these substances. You can estimate your yearly radiation exposure using the EPA's "[Calculate Your Radiation Dose](#)."

References

1. <https://www.epa.gov/radon>
2. <https://www.ornl.gov/ptp/collection/consumer%20products/consumer.htm>

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- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a [Creative Commons Attribution License 4.0](#) license. Download for free at <http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110>.
- Elizabeth R. Gordon (Furman University)

- Emma Gibney (Furman University)
- Erin Avram (Cleveland State University)

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5.4: The Basics of Skin Cancer

Learning Objectives

- Distinguish the different types of ultraviolet radiation
- Summarize the ABCDEs of melanoma and protective measures that can be taken
- Identify ADA-recommended skin care products (i.e.- SPF, Broad Spectrum terminology, UPF)

Exposure to Ultraviolet Radiation

Of all the types of ionizing radiation, people come into more contact with ultraviolet radiation. Sources of this radiation would include the sun, ultraviolet lights, and tanning beds. The sun produces three different forms of ultraviolet rays (UVA, UVB, and UVC). Fortunately, the deadliest of these three rays (UVC) never reaches the earth's surface and is absorbed by the Earth's ozone layer. However, both UVA and UVB reach the Earth's surface and impact living systems.

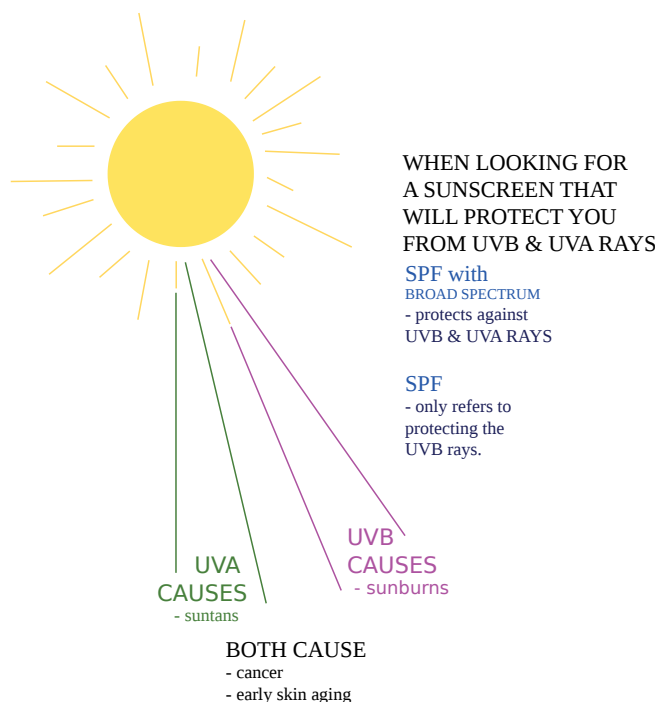


Figure 5.4.1: Effect of Light on skin

Both UVA and UVB are produced naturally by the sun and artificially by lamps and tanning beds. For years, sunbathers believed that UVA provided only a youthful glow, whereas UVB are responsible for sunburns. Now, the American Academy of Dermatologists has condemned the use of tanning beds and exposure to both UVA and UVB. Scientific research has correlated certain types of skin cancers (mainly basal and squamous cell carcinoma) with exposure to UVA and UVB, particularly from tanning beds. Besides skin cancer, tanning beds cause premature aging and damage to your eyes. If a tanning bed is not properly sanitized between uses, skin diseases can be transmitted from one consumer to another. UVA rays make up 95% of sunlight and are able to penetrate the skin more deeply. UVB makes up about 5% of the UV rays from the sun, doesn't penetrate the skin as far, but has a lot more energy.

Recognizing Suspicious Skin Changes

According to the American Academy of Dermatologists, skin cancer, is the most common form of cancer. Exposure to ultraviolet rays or arsenic can cause skin diseases. Additionally, some individuals may be genetically predisposed to developing this illness

(check out your family tree). Approximately 10,000 Americans a day will be diagnosed with one of the three different forms of skin cancer (basal cell carcinoma, squamous cell carcinoma, and melanoma). These types of cancer are named based on the type of skin cells that are affected. Millions of Americans have already been diagnosed with the most common form of skin cancer, basal cell carcinoma. This type of cancer grows very slowly and can spread to other parts of the body. Basal cell carcinoma frequently appears on the head and neck of a patient.

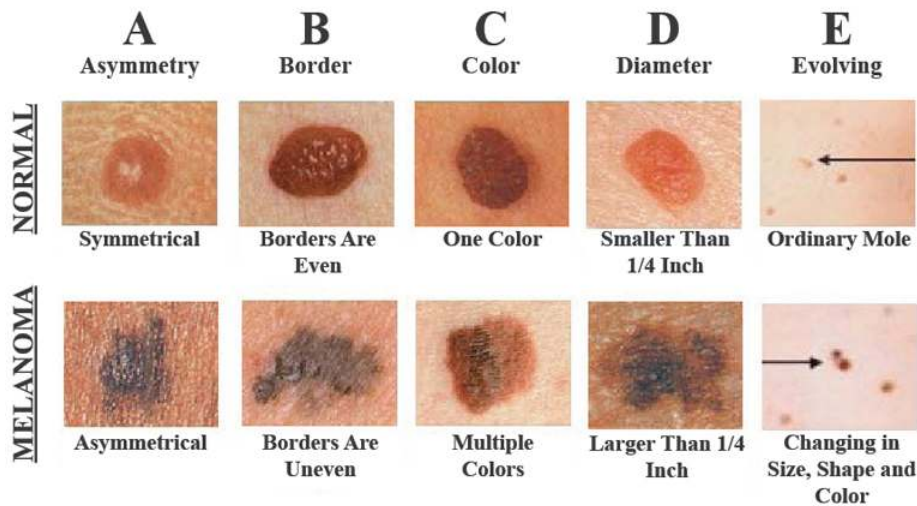


Figure 5.4.2: ABCDEs of Skin Cancer. A brown spot on the skin is likely to be a harmless mole, but it could be a sign of skin cancer. Unlike moles, skin cancers are generally asymmetrical, have irregular borders, may be very dark in color, and may have a relatively great diameter.

Squamous cell carcinoma appears more commonly on the face, neck, lips, and hands of the person. For females, squamous cell carcinoma typically presents on the lower leg portion. This form of skin cancer is more likely to metastasize (spread) than basal cell carcinoma. If caught early and treated, patients can easily survive both basal and squamous cell carcinomas.

Melanoma, the most serious form of skin cancer, is the least common of all skin cancers. This type of cancer usually manifests on the trunks of males and the legs of females. It can spread easily to the brain, liver, bones, abdomen, or lymph nodes. If detected early, survival rates are good and life expectancy will be normal.

Figure 5.4.2 compares normal pigmentations to moles that have changed into melanoma. Monthly reviews of skin lesions and seeking medical advice is key to surviving this disease. Be sure to note these visual characteristics (the ABCDEs) of skin cancer for future evaluations (life and class).

Sunblocks and Sunscreens

UVB radiation in sunlight allows the skin to produce **vitamin D**. This vitamin prevents bone disorders like rickets and osteoporosis (brittle bone disease). The American Academy of Dermatology suggests vitamin D be obtained through foods or nutritional supplements. Excessive exposure to UV can be damaging and the pigment **melanin**, deposited in cells at the base of the epidermis, helps to protect the underlying layers of the skin from this damage. Melanin also colors the skin, variations in the amount of melanin produce colors from pale yellow to black. The darker the skin tone, the more melanin one has, and the less likely skin cancer will occur.



Figure 5.4.3: The Dr. Elizabeth Gordon's great, great-grandmother was diagnosed with skin cancer on her nose during the 1950's.

Sunscreens and sunblocks are designed to protect skin from ultraviolet rays. *Sunblocks* contain inorganic ingredients like zinc oxide or titanium dioxide. These chemicals act as UV filters by reflecting the sun's UV rays. Sunblocks can have grainy textures due to the inorganic components. The thick nature of a sunblock can make it difficult to spread evenly on the skin. *Sunscreens* contain organic compounds like oxybenzone, avobenzone, homosalate, and octinoxate (Figure 5.4.4). By absorbing ultraviolet rays, these compounds decompose and give off heat. Sunscreens apply smoother than sunblocks. Often, manufacturers will combine sunscreen and sunblock ingredients to make their products.

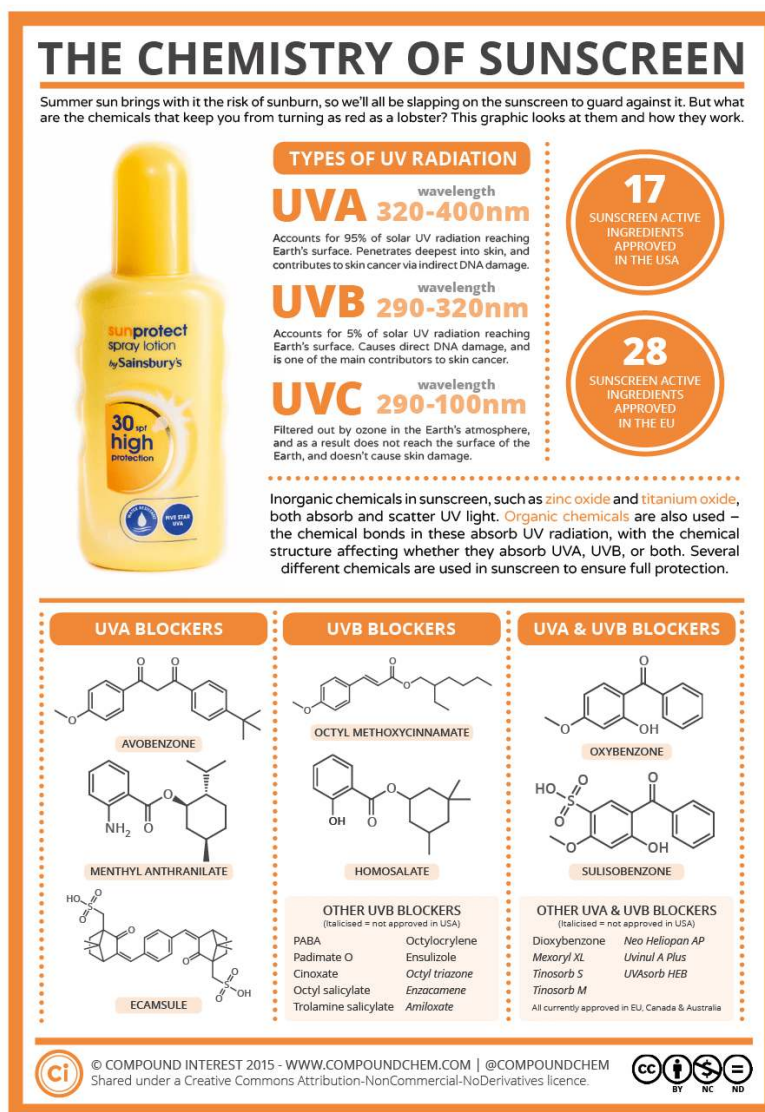


Figure 5.4.4: Effect of Light on skin

Sun protection factor (SPF) measures a product's protection from UVB rays. SPF does not quantify protection from UVA radiation. The American Academy of Dermatology recommends selection of a sunscreen or sunblock with a minimum SPF of 30. This SPF value means a lotion can filter out 97% of UVB rays. A product with SPF 50 will only filter out 1 more percentage of UVB rays. Increasing an SPF factor past 30 does very little in shielding skin from UVB radiation.

Products that protect skin from UVA must be labeled as "broad spectrum." Sunblocks provide UVA and UVB protection, but sunscreens can vary on what they are effective against. Every two hours, sunscreens/sunblocks should be reapplied to the skin. Sweating and swimming can remove sunscreen/sunblock products. No sun products are waterproof, but some are labeled as being water-resistant. Consumers are encouraged to reapply these products every two hours as well.

Clothing and Other Garments

Though most people are unaware, the clothing they wear every day can serve as a barrier against harmful ultraviolet rays, although most garments offer only a little protection. The garment Ultraviolet Protection Factor (UPF) is measured by how much UVA and UVB rays come in contact with the skin. Garments that receive a good rating by the Skin Cancer Foundation's Seal of Recommendation have a UPF rating of at least 30. Fabrics that receive excellent ratings have UPF values of 50 and can block 98% of the sun's rays, allowing only 2% to penetrate the fabric to the skin.



Video 5.4.1: Labor Day skincare tips: UPF clothing. Whether you're heading to a football game or having a cookout by the pool, it's important to protect your skin from the sun, and sunscreen isn't the only way you can do that. Dr. Kellie Reed with Sanova Dermatology discusses UPF clothing.

Clothing does not protect equally. There are seven specific factors that make clothing more effective: UPF, color, construction, content, fit, coverage, and activity. Clothing that is dark or brightly colored is more likely to absorb UV rays rather than letting them penetrate to your skin. Densely woven clothing, like denim, wool, or synthetic fibers is more beneficial than tight clothing. If clothing becomes stretched, thinner, or wet, UV rays can seep through the fabric. Therefore, the thicker and looser the clothing, the more protection it will provide for the skin. Shiny polyester and lightweight silk fabrics are highly protective because they reflect radiation. Other fabrics like unbleached cotton and high-tech fabrics treated with chemical UV absorbers or dyes are also effective in shielding skin from UV exposure.

When spending large amounts of time outside, wearing long-sleeved shirts and pants offers more protection for your skin. Topping off your outfit with a wide-brimmed hat maximizes protection from UV rays and helps prevent basal and squamous cancers that often appear on the head and neck (locations that account for 90% of skin cancers). Lastly, select sunglasses certified to protect your eyes from UV rays, a good pair of sunglasses will reduce your risk of developing cataracts later in life.

Contributors and Attributions

- [Elizabeth R. Gordon \(Furman University\)](#)
- [Template:ContribLawson](#)
- Emma Gibney, Jordan Radoll, Hannah Perry, and McKenna White (Furman University)
- Erin Avram (Cleveland State University)

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5.5: Calculating Half-Life

Learning Objectives

- Define half-life and explain the factors that affect it.
- Calculate the amount of radioactive material that will remain after an integral number of half-lives.
- Calculate the age of a material based on its half-life.
- Describe how carbon-14 is used to determine the age of carbon-containing objects.

Rate of Radioactive Decay

During natural radioactive decay, not all atoms of an element are instantaneously changed to atoms of another element. The decay process takes time and there is value in being able to express the rate at which a process occurs. A useful concept is **half-life** (symbol is $t_{1/2}$), which is the time required for half of the starting material to change or decay. Half-lives can be calculated from measurements of the change in mass of a nuclide and the time it takes to occur. Although chemical changes can be sped up or slowed down by changing factors such as temperature, concentration, etc, these factors have no impact on half-life. Each radioactive isotope has its own, unique half-life that is independent of any of these factors.

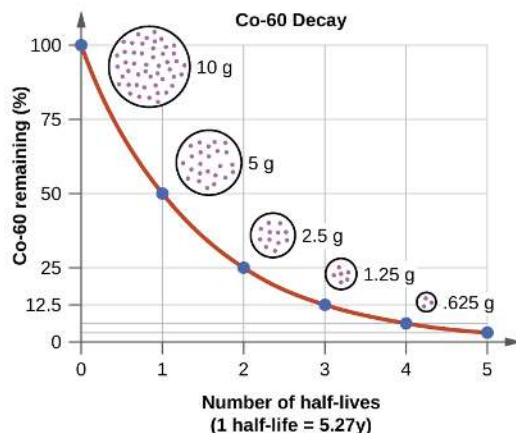


Figure 5.5.1: For cobalt-60, which has a half-life of 5.27 years, 50% remains after 5.27 years (one half-life), 25% remains after 10.54 years (two half-lives), 12.5% remains after 15.81 years (three half-lives), and so on. (CC-BY 4.0; OpenStax).

The half-lives of many radioactive isotopes have been determined and have been found to range from extremely short half-lives of fractions of a second to extremely long half-lives of 10 billion years. The table below illustrates half-lives for selected elements. In addition, the final elemental product is listed after the decay process.

Table 5.5.1: Table of Selected Half-lives

Element	Mass Number (A)	Half-life	Element	Mass Number (A)	Half-Life
Uranium	238	4.5 Billion years	Californium	251	800 years
Neptunium	240	1 hour	Nobelium	254	3 seconds
Plutonium	243	5 hours	Carbon	14	5730 years
Americium	245	25 minutes	Carbon	16	740 milliseconds

The quantity of radioactive nuclei at any given time can be calculated based on the length of 1 half-life. For each half-life, the mass of the sample will decrease by one-half. For example, if there were **100 g** of **Cf-251** in a sample at some time, after 800 years (1 half-life), there would be **50 g** of **Cf-251** remaining. After another 800 years (1600 years total, or 2 half-lives), there would only be **25 g** remaining.

Remember, the half-life is the time it takes for half of your sample, no matter how much you have, to remain. Each half-life will follow the same general pattern as **Cf-251**. The only difference is the length of time it takes for half of a sample to decay.

📌 Interactive Simulation: Visualizing Half-Life

Click on this [interactive simulation](#) to visualize what happens to a radioisotope when it decays, and learn about different types of radiometric dating, such as carbon dating. Understand how decay and half life work to enable radiometric dating. Play a game that tests your ability to match the percentage of the dating element that remains to the age of the object.

There are two types of half-life problems we will perform. One format involves calculating a mass amount of the original isotope. Using the equation below, we can determine how much of the original isotope remains after a certain interval of time.

$$\text{how much mass remains} = \frac{1}{2^n} (\text{original mass})$$

n is the number of half-lives.

✓ Example 5.5.1

If there are 60 grams of **Np-240** present, how much **Np-240** will remain after 4 hours? (**Np-240** has a half-life of 1 hour)

Solution

$$\text{how much mass remains} = \frac{1}{2^4} (60 \text{ grams})$$

After 4 hours, only **3.75 g** of our original **60 g** sample would remain the radioactive isotope **Np-240**.

✓ Example 5.5.2

A sample of **Ac-225** originally contained **8.0 ug**. After 720 hours, how much of the original **Ac-225** remains? The half-life of this isotope is 10 days.

Solution

To determine the number of half-lives (n), both time units must be the same.

$$\cancel{720 \text{ hours}} \times \frac{1 \text{ day}}{\cancel{24 \text{ hours}}} = 30 \text{ days}$$

$$n = 3 = \frac{30 \text{ days}}{10 \text{ days}}$$

$$\text{how much mass remains} = \frac{1}{2^3} (8.0 \text{ ug})$$

After 720 hours, **1.0 ug** of the material remains as **Ac-225**

Radioactive Dating

Radioactive dating is a process by which the approximate age of an object is determined through the use of certain radioactive nuclides. For example, carbon-14 has a half-life of 5,730 years and is used to measure the age of organic material. The ratio of carbon-14 to carbon-12 in living things remains constant while the organism is alive because fresh carbon-14 is entering the organism whenever it consumes nutrients. When the organism dies, this consumption stops, and no new carbon-14 is added to the organism. As time goes by, the ratio of carbon-14 to carbon-12 in the organism gradually declines because carbon-14 radioactively decays while carbon-12 is stable. Analysis of this ratio allows archaeologists to estimate the age of a sample.

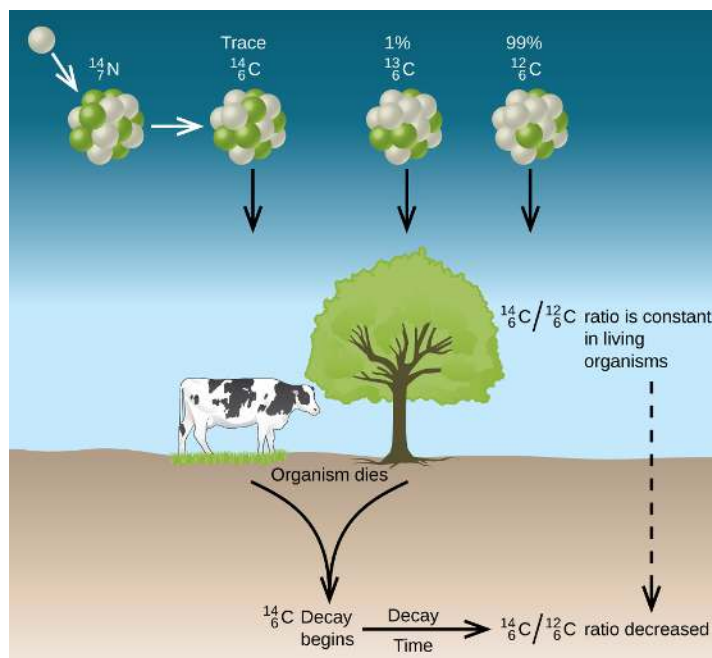


Figure 5.5.2: Along with stable carbon-12, radioactive carbon-14 is taken in by plants and animals, and remains at a constant level within them while they are alive. After death, the C-14 decays and the C-14:C-12 ratio in the remains decreases. Comparing this ratio to the C-14:C-12 ratio in living organisms allows us to determine how long ago the organism lived (and died). (CC-BY 4.0; OpenStax).

C-14 dating does have limitations. For example, a sample can be C-14 dated if it is approximately 100 to 50,000 years old. After this range there is too little of the isotope to be detected. C-14 is obtained from the atmosphere, for this reason, aquatic samples cannot be effectively C-14 dated. Lastly, the accuracy of C-14 dating has been affected by nuclear weapons testing because fission bombs produce C-14 artificially. It is generally assumed that the levels of C-14 in the atmosphere have not changed over time, however, nuclear weapons testing above ground primarily occurred in the 1950s and 1960s, falsely increasing the atmospheric C-14 levels higher than normal during this time frame. This means that samples in existence during this time will have higher than anticipated C-14 levels. It is therefore important to incorporate additional tests and controls for these types of samples.



To calculate the age of a substance using isotopic dating, use the equation below:

$$\text{how old (time)} = n \times t_{1/2}$$

n is the number of half-lives and $t_{1/2}$ is the half-life (in time).

✓ Example 5.5.3

How long will it take for 18.0 grams of Ra-226 to decay to leave a total of 2.25 grams? Ra-226 has a half-life of 1600 years.

Solution

18.0g $\Rightarrow$ 9.0g $\Rightarrow$ 4.5g $\Rightarrow$ 2.25g, this is three half-lives

$$\text{how old (time)} = 3 \times 1600 \text{ years}$$

This decay process takes 4800 years to occur.

Radioactive Dating Using Nuclides Other than Carbon-14

Radioactive dating can also use [other radioactive nuclides](#) with longer half-lives to date older events. For example, uranium-238 (which decays in a series of steps into lead-206) can be used for establishing the age of rocks (and the approximate age of the oldest rocks on earth). Since U-238 has a half-life of 4.5 billion years, it takes that amount of time for half of the original U-238 to decay into Pb-206. In a sample of rock that does not contain appreciable amounts of Pb-208, the most abundant isotope of lead, we can assume that lead was not present when the rock was formed. Therefore, by measuring and analyzing the ratio of U-238:Pb-206, we can determine the age of the rock. This assumes that all of the lead-206 present came from the decay of uranium-238. If there is additional lead-206 present, which is indicated by the presence of other lead isotopes in the sample, it is necessary to adjust the age calculation.

Potassium-argon dating uses a similar method to uranium-lead dating. K-40 decays by positron emission and electron capture (two additional types of radioactive decay we didn't discuss) to form Ar-40 with a half-life of 1.25 billion years. If a rock sample is crushed and the amount of Ar-40 gas that escapes is measured, the determination of the Ar-40:K-40 ratio yields the age of the rock. Other methods, such as rubidium-strontium dating (Rb-87 decays into Sr-87 with a half-life of 48.8 billion years), operate on the same principle. To estimate the lower limit for the earth's age, scientists determine the age of various rocks and minerals, making the assumption that the Earth is older than the oldest rocks and minerals in its crust. As of 2014, the oldest known rocks on earth are the Jack Hills zircons from Australia, found by uranium-lead dating to be almost 4.4 billion years old.

An ingenious application of half-life studies established a new science of determining the ages of materials by half-life calculations. For geological dating, the decay of U-238 can be used. The half-life of U-238 is 4.5×10^9 years. The end product of the decay of U-238 is Pb-206. After one half-life, a 1.00 gram sample of uranium-238 will have decayed to 0.50 grams of U-238 and 0.43 grams of Pb-206. By comparing the amount of U-238 to the amount of Pb-206 in a sample of uranium mineral, the age of the mineral can be estimated. Present-day estimates for the age of the Earth's crust from this method is at 4 billion years. This radioactivity approach can be used to detect fake wine vintages too.

Isotopes with shorter half-lives are used to date more recent samples. For example, tritium ($t_{1/2} = 12.3$ years) can date samples within an age range of 1-100 years. Chemists and geologists use tritium dating to determine the age of water (ocean and freshwater). In addition, tritium dating can be useful in determining the age of wines and brandies.

Contributors and Attributions

- CK-12
- Elizabeth Gordon (Furman University)
- Emma Gibney (Furman University)
- Erin Avram (Cleveland State University)

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5.6: Applications of Nuclear Chemistry

There are many interesting applications of nuclear chemistry from energy to medicine. Choose ONE of the following sections that interests you the most to read and reflect on. In OneNote, use bullet points to outline the key ideas from your chosen reading.

[5.6.1 Using Nuclear Science to Diagnose Disease](#)

[5.6.2 Therapeutic Radiation](#)

[5.6.3 Non-medical Applications of Nuclear Science](#)

[5.6.4 Uses and Consequences of Nuclear Energy](#)

[5.6.5 Three Mile Island Nuclear Accident](#)

[5.6.6 Chernobyl Nuclear Disaster](#)

[5.6.7 Fukushima Nuclear Disaster](#)

[5.6.8 Nuclear Bombs](#)

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5.6.1: Using Nuclear Science to Diagnose Disease

Learning Objectives

- Understand the difference between diagnostic and therapeutic radiation.
- Know what testing involves nuclear chemistry.
- Calculate dosages of radiation received with a nuclear test.
- Compare nuclear testing dosages to toxic concentrations.

Diagnostic medical applications involve testing for a disease or condition. In nuclear medicine, this may involve x-rays, CT or PET scans, or isotopic studies. The radiation involved for each of these types of tools will vary in mrem or mSv amounts.

X-Ray Imaging

X-ray imaging is one of the most basic and routine forms of imaging within modern medicine. Using electromagnetic radiation, scientists and doctors have the ability to visualize internal situations. X-ray imaging takes advantage of the fact that dense structures (i.e., bone) absorb more X-rays than the softer tissues that surround it. By projecting radiation onto special film paper, the denser substance casts a shadow where less X-rays pass through the body, and a contrasting image is produced. The radiation emitted by an X-ray is of the wavelength of 10 to 0.01 nanometers (nm). Because some parts of the body may be thicker than others, X-rays are used in "soft" (10 to 0.10 nm) and "hard" (0.10 to 0.01 nm) dosages or radiation.

X-rays are produced by passing an electron beam through a tungsten anode in a vacuum. The electromagnetic waves that are released can be focused and adjusted according to what is imaged. X-rays can produce very high definition images at a low cost, but they provide poor contrast in soft tissues. Contrast agents can also be added to the body prior to imaging, which allows different tissues to absorb more radiation, creating a better contrasting image. This allows doctors and scientists to work with more specific images within their diagnosis or research (i.e., the ability to determine different aspects of soft-tissue).

Projectional radiography (Figure 5.6.1.1) can be done standing or laying down. To obtain clearer images, a patient needs to be as still as possible. A technician may ask the patient to take a big breath and hold it right before pushing the scan button. This allows the patient to keep their chest and abdomen in position and will allow a clearer image to be captured. Breathing will cause movement which will result in a blurred image.

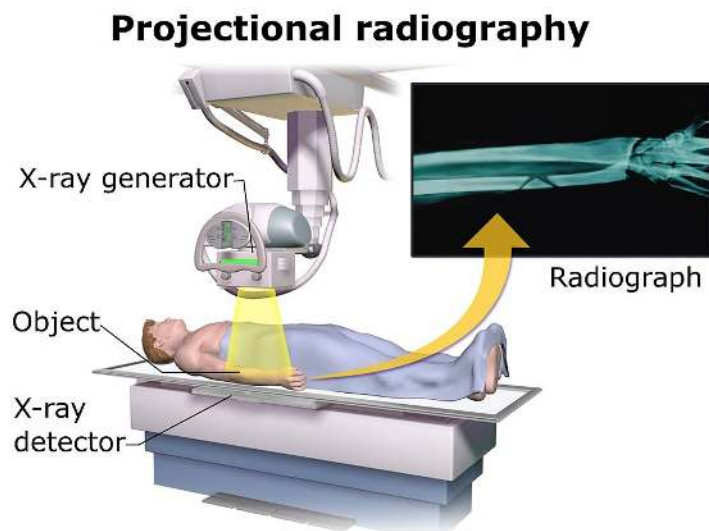


Figure 5.6.1.1: Acquisition of projectional radiography, with an X-ray generator and a detector. (CC BY-SA 4.0; [Blausen Medical](#)).

Before having an X-ray or CT scan, patients should remove all metallic clothing or jewelry. Metals can distort X-ray images easily. Sometimes, lead aprons are provided to patients. The lead aprons shield a patient's reproductive organs from ionizing radiation. X-rays and CT scans can harm a developing embryo or fetus. For this reason, female patients are required to sign a liability waiver stating they are not pregnant at the time of testing. Other options like MRI or ultrasounds (nonionizing techniques) may be an alternative option for these patients.

Diagnostic X-rays differ greatly in millisievert (mSv) dosages. Table 5.6.1.1 illustrates a few of these values. Using this table, you can calculate rem and mrem amounts and compare it to toxic values discussed in this text.

Table 5.6.1.1: Diagnostic x-rays levels

Type of X-ray Exam	Diagnostic Reference level (DRL) in mSv
Chest	0.96
Abdomen	7.10
C-spine	3.56
T-spine	8.00
L-Spine	12.0
Skull	6.00
Lumb	7.21
Hip	7.21

Source: Sonawane AU¹, Shirva VK, Pradhan AS., [Radiat Prot Dosimetry](#). 2010 Feb;138(2):129-36. doi: 10.1093/rpd/ncp235.

✓ Example 5.6.1.1: Chest X-ray

A chest x-ray yields a dosage of 6 mrem of radiation to a patient. Calculate this dose in rem and millisieverts (mSv). Compare these values to your annual background dose (around 360 mrem) and cancer dose (100 rem).

Solution

For the first calculation, recall the metric prefix milli from Chapter 2.

$$\frac{6 \text{ mrem}}{1} \times \frac{1 \text{ rem}}{1000 \text{ mrem}} = 6 \times 10^{-3} \text{ rem}$$

For the second calculation, you will need to insert the 100 rem = 1 Sv factor. Then, you will need to use your milli metric prefix again.

$$\frac{6 \times 10^{-3} \text{ rem}}{1} \times \frac{1 \text{ Sv}}{100 \text{ rem}} \times \frac{1000 \text{ mSv}}{1 \text{ Sv}} = 6 \times 10^{-2} \text{ mSv}$$

The dose of 6 mrem dose is significantly lower than the level of background radiation. In addition, the calculated $6 \times 10^{-3} \text{ rem}$ value is well below the 100 rem dosage that is associated with development of cancer.

✓ Example 5.6.1.2: Mammogram

Your grandmother/mother goes to get her yearly mammogram (x-ray image that helps detect breast cancer). The radiation level for this scan is 0.36mSv. Convert this value to mrem and rem. Compare calculated values to background and cancer-causing levels (shown in previous problem).



Figure 5.6.1.2: Sharon Grubbs, 2d Medical Group mammography technologist, demonstrates the mammography machine. (Public Domain; U.S. Air Force photo by Staff Sgt. Mike Andriacco)

$$\frac{0.36 \cancel{mSv}}{1} \times \frac{1 \cancel{Sv}}{1000 \cancel{mSv}} \times \frac{100 \cancel{rem}}{1 \cancel{Sv}} = 3.6 \times 10^{-2} \cancel{rem}$$

$$\frac{3.6 \times 10^{-2} \cancel{rem}}{1} \times \frac{1000 \cancel{mrem}}{1 \cancel{rem}} = 36 \cancel{mrem}$$

36 mrems is less than an average person's background radiation of 360 mrems. In addition, this value is 96.4 rems below the cancer level of 100 rems.

CT Scanning

Imagine a patient enters the hospital complaining of severe chest pain. After consulting the patient, the doctors decide that the source of the patient's problems is a constricted artery. However, the human body contains several long arteries. The doctors could use surgery to find the source of the clot, but this could take many hours and with no definite guarantee of success. In this case, radiation can help medical professionals reach a diagnosis much faster. Computerized Axial Tomography (CT) scans use X-ray particles along with advanced computer technology in order to produce highly detailed cross-sectional images of the body. Using these scans, medical professionals can accurately find clots and other harmful medical conditions.

CT/CAT scans have the same fundamentals of X-rays, but they produce 3-D images using tomography. Using digital geometric processing, the CT scan creates a 3-D image compiling a large number of 2-D X-rays. These 2-D images are thin cross-sections of the body taken from a single axis of rotation. A CT scanner consists of a table where the patient lies down, an X-ray device that rotates at high speed around the patient taking hundreds of cross-sections of the body, and a computer with appropriate software to compile and render the 3-D image. The table in Figure 1 moves through the tube, allowing the machine to take hundreds of pictures of the patient from head to toe.



Figure 5.6.1.3: PET/CT-System with 16-slice CT; the ceiling-mounted device is an injection pump for CT contrast agent. (Public Domain; Hg6996).

Similar to the X-ray, contrast agents can be added to the body before imaging occurs so that different tissues are more defined. Due to the fact that many more images are being taken in order to form a complete three-dimensional image, the radiation dose is much higher. While a chest x-ray is around 0.10 mSv or radiation, a chest CAT scan subjects the patient to about 8 mSv of radiation; over twice the amount that the body would absorb from its environment in a year. For this reason, it is uncommon and unsafe to have multiple CAT scans taken unless completely necessary.

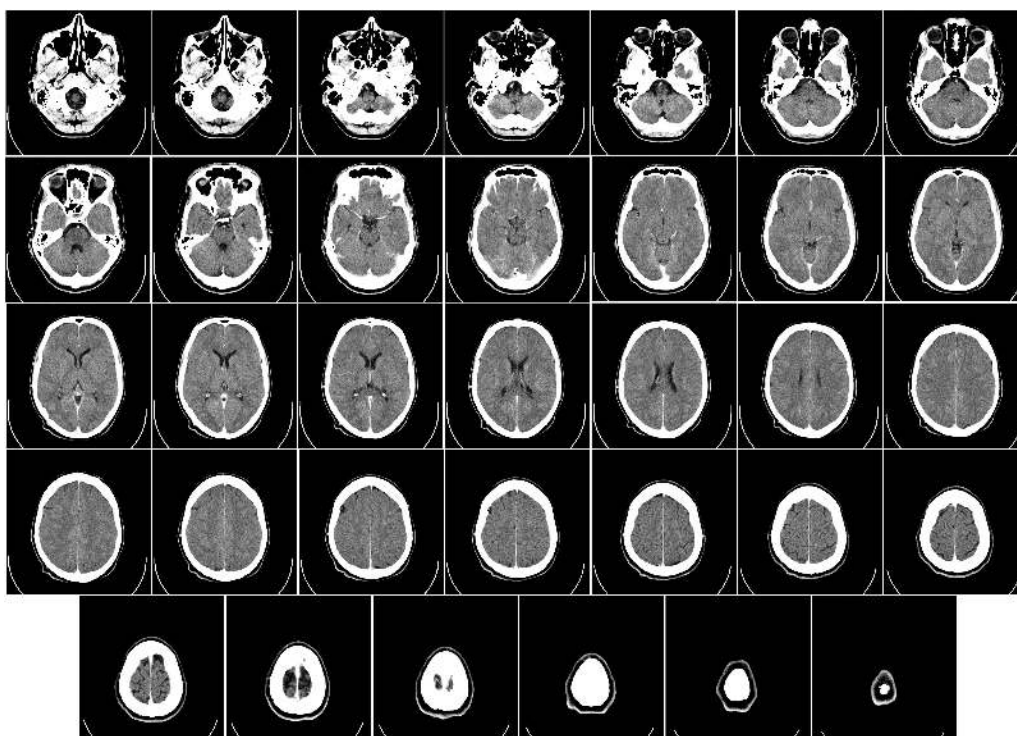


Figure 5.6.1.4: CT scan of the human brain. These images assist in diagnosing migraines, brain tumors, strokes, bleeding, and head injuries. upload.wikimedia.org/Wikipedia/commons/5/50/Computed_tomography_of_human_brain_-_large.png

✓ Example 5.6.1.3: Body CT Scan

Your doctor suspects you have appendicitis and orders a CT scan. This type of evaluation exposes you to 12 mSv of radiation. Convert the given mSv amount to rem and mrem. Lastly, compare this value to background and cancer mrem/rem values.

$$\frac{12 \cancel{\text{mSv}}}{1} \times \frac{1 \cancel{\text{Sv}}}{1000 \cancel{\text{mSv}}} \times \frac{100 \text{ rem}}{1 \cancel{\text{Sv}}} = 1.2 \text{ rem}$$

$$\frac{1.2 \cancel{\text{rem}}}{1} \times \frac{1000 \text{ mrem}}{1 \cancel{\text{rem}}} = 1200 \text{ mrem}$$

PET Scanning

Positron Emission Tomography or PET scan is a type of nuclear medicine imaging. Depending on the area of the body being imaged, a radioactive isotope is either injected into a vein, swallowed by mouth, or inhaled as a gas. When the radioisotope is collected in the appropriate area of the body, the gamma-ray emissions are detected by a PET scanner (often called a gamma camera) which works together with a computer to generate special pictures, providing details on both the structure and function of various organs. Watch this informational [video](#) on how this technique works. PET scans are used to:

- Detect cancer
- Determine the amount of cancer spread
- Assess the effectiveness of treatment plans
- Determine blood flow to the heart muscle
- Determine the effects of a heart attack
- Evaluate brain abnormalities such as tumors and memory disorders
- Map brain and heart function

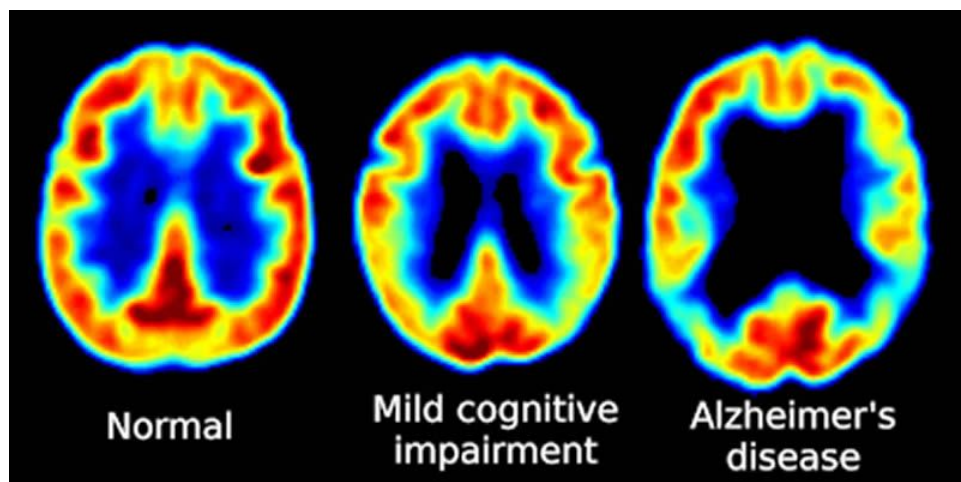


Figure 5.6.1.5: Mild cognitive impairment (MCI) is a state between normal ageing and dementia, where someone's mind is functioning less well than would be expected for their age. This image is for illustrative purposes only. (Public Domain; Center For Functional Imaging, Lawrence Berkeley National Laboratory. Alzheimer's Disease Neuroimaging Initiative (ADNI).

PET Scanning, is used to image the physiological aspects of the body rather than the anatomy. It images the function of the body rather than the form, such as where tagged molecules go and how they are used. For instance, if you were to image the brain of a deceased person, nothing would show up on a PET scan as opposed to a CAT scan, as the brain is no longer functional. PET Scanning is very useful in imaging tumors, which can be done when patients are injected with certain tracers. Often times PET scanners are used in collaboration with CAT scanners to create a composite image that shows both the function and form of the body. The animation below is a whole-body PET scan using the radioisotope of ^{18}F ($t_{1/2} = 110 \text{ min}$). Using this tracer, doctors can determine if cancer has metastasized by looking at the metabolic activity of glucose.

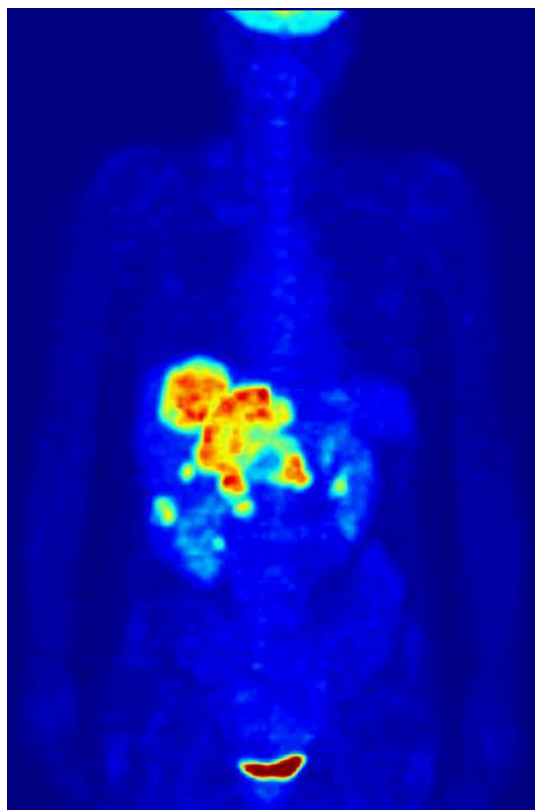


Figure 5.6.1.6: A whole body analysis by a PET scan. (Public Domain; Jens Maus, <http://jens-maus.de/>)

Other Isotopic Tests

Radioisotopes have revolutionized medical practice, where they are used extensively. More than 10 million nuclear medicine procedures and more than 100 million nuclear medicine tests are performed annually in the United States. Four typical examples of radioactive tracers used in medicine are technetium-99 ($^{99}_{43}\text{Tc}$), thallium-201 ($^{201}_{81}\text{Tl}$), iodine-131 ($^{131}_{53}\text{I}$), and sodium-24 ($^{24}_{11}\text{Na}$). Damaged tissues in the heart, liver, and lungs absorb certain compounds of technetium-99 preferentially. After it is injected, the location of the technetium compound, and hence the damaged tissue, can be determined by detecting the γ rays emitted by the Tc-99 isotope. Thallium-201 (Figure 5.6.1.1) becomes concentrated in healthy heart tissue, so the two isotopes, Tc-99 and Tl-201, are used together to study heart tissue. Iodine-131 concentrates in the thyroid gland, the liver, and some parts of the brain. It can therefore be used to monitor goiter and treat thyroid conditions, such as Grave's disease, as well as liver and brain tumors. Salt solutions containing compounds of sodium-24 are injected into the bloodstream to help locate obstructions to the flow of blood.

Small doses of I-131 (too small to kill cells) are used for purposes of imaging the thyroid. Once the iodine is concentrated in the thyroid, the patient lays down on a sheet of film and the radiation from the I-131 makes a picture of the thyroid on the film. The half-life of iodine-131 is approximately 8 days so after a few weeks, virtually all of the radioactive iodine is out of the patient's system. During that time, they are advised that they will set off radiation detectors in airports and will need to get special permission to fly on commercial flights.

Some isotopes that are used to diagnose diseases are shown in Table 5.6.1.2. All of these nuclear isotopes release one form of ionizing radiation (either α particle or γ ray). In addition, each isotopic application would involve a specific amount of mrem/mSv radiation.

Table 5.6.1.2: Selected radioisotopes used in diagnostic nuclear medicine

Symbol-mass	Half-Life ($t_{1/2}$)	Application
Xe-133	5.27 days	Lung imaging
H-3	12.26 years	Analyzing total body water
Tl-201	73 hours	Stress tests for heart problems
Fe-59	44.5 days	Detection of anemia
Gd-153	242 days	Analyzing bone density
Cr-51	27.8 days	Determining blood volume
C-11	20.4 minutes	Brain scans
Tc-99m	6.0 hours	Heart, lung, kidney, bone marrow, brain, or bone marrow imaging
Pu-238	86 years	Powering pacemakers
I-131	8.0 days	Imaging Thyroid

Keep in mind that X-rays, CT scans, PET scans, and isotopic studies involve ionizing radiation. In contrast, MRI (magnetic resonance imaging) and ultrasounds do not utilize ionizing forms of radiation.

Sources

1. <http://cerncourier.com/cws/article/cern/66177> (radioisotope information)
2. [en.Wikipedia.org/wiki/Nuclear_medicine](https://en.wikipedia.org/wiki/Nuclear_medicine) (radioisotope information)

Contributors and Attributions

- Elizabeth R. Gordon (Furman University)
- Mike Reed (UC Davis), Eugene Kwon (UC Davis)
- CK-12
- Emma Gibney (Furman University)
- Erin Avram (Cleveland State University)

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5.6.2: Therapeutic Radiation

Learning Objectives

- Distinguish between diagnostic and therapeutic radiation.
- Know the types of cancer treatments available.
- Recall the A/Z formats of a proton (particle) and gamma (ray).
- Compare/contrast external beam therapy and brachytherapy.
- Know the differences between proton and photon radiation.
- Realize the side effects of radiation therapy and medications available.

There are many techniques used to treat cancer. Surgery can be used to remove cancerous tumors inside or on the body. With chemotherapy, ingested or injected chemicals are used to kill rapidly dividing cells (cancerous and noncancerous). For more information regarding chemotherapy, please access this [link](#). Other cancer treatment methods include immunotherapy, stem cell replacement, hormone therapy, and targeted therapy.

Radiation Therapy and Chemotherapy: Two different treatment procedures

Patients diagnosed with cancer might be required to undergo chemotherapy or radiation therapy. Sometimes, both of these methods are used for a patient. In this picture, a female patient is receiving chemotherapy through an IV. She is participating in cooling therapy while receiving her treatment. By placing her hands and feet in cooling devices, this will lower her chances of losing her fingernails and toenails. Cold cap therapy is also now available for chemotherapy patients. Wearing this type of device increases a patient's chances of keeping his/her hair during chemotherapy.



Figure 5.6.2.1: (CC BY 2.0; Jenny Mealing).

Current therapeutic radiation applications involve the use of gamma, X-rays, or protons. Recently, some research facilities are investigating the use of alpha- or beta-tagged molecules to kill cancer cells. These molecules will target a cancer-related molecule on a tumor cell. Then, the alpha- or beta-tagged species will deliver its radiation to the tumor. Sr-89 (beta emitter) and Ra-223 (alpha emitter) have been used in clinical research trials of certain types of bone cancers.

Oncologists (specialists who treat patients with cancer) utilize radiation to help slow or cure the spread of cancer within individuals. Radiation is specifically applied to malignant tumors to shrink them in size. Medical professionals, mainly radiation oncologists, administer a variety of dosages to patients, contingent on factors such as the patient's current health, the concurrent use of other treatments, and the success of surgery.

External Beam Therapy (Photon and Proton Therapy)

External Beam Therapy (EBT) is a method of delivering a high-energy beam of radiation to the precise location of a patient's tumor. These beams can destroy cancer cells and with careful planning, preserve surrounding cells. Several beams of sub-lethal doses of radiation enter the body from different directions. The only place in the body where the beam is lethal is the spot where all the beams intersect. Before the EBT process, the patient is three-dimensionally mapped using CT scans and X-rays. The patient receives small tattoos to allow the therapist to line up the beams exactly. Alignment lasers are used to precisely locate the target. The video below illustrates the basic preparation and administration of external beam therapy.



Targeting Cancer - Radiation Therapy Treatment Process

Photon EBT utilizes either X-ray or gamma rays. An X-ray source would require a linear accelerator to produce high-energy electrons. In contrast, a gamma source incorporates a radioactive isotope (like Co-60). Keep in mind both of these technologies use ionizing radiation. As a result, cancer patients must be monitored throughout their life because they are at an increased risk for developing other cancers such as leukemia. EBT is used to treat many cancers such as:

- Breast cancer
- Colorectal cancer
- Head and neck cancer
- Lung cancer
- Prostate cancer

The majority of radiation patients receive photon EBT. The smaller size of this machine makes this therapy an option for all sizes of hospitals and cancer treatment centers. The equipment required for Photon EBT costs approximately three million dollars. The size and the price of this technology enables smaller facilities to keep their patients closer to home during treatment

In the past ten years, another method of radiation treatment involving protons has become more widely available. This type of therapy requires a cyclotron to generate proton beams (recall, a proton is an ionized H-1 isotope). As compared to X-rays or gamma rays (photon therapy), protons are very heavy. Today's proton facilities occupy large plots of land to house the medical cyclotron (the size of a football field that is three stories high). With a price tag of 150-200 million dollars, these facilities are sporadically located in the United States.



Figure 5.6.2.2: Current Proton Therapy Centers in the US. The world's first hospital-based proton therapy center was a low energy cyclotron centre for ocular tumors at the Clatterbridge Centre for Oncology in the UK, opened in 1989, followed in 1990 at the Loma Linda University Medical Center in Loma Linda, California. Later, The Northeast Proton Therapy Center at Massachusetts General Hospital was brought online, and the HCL treatment program was transferred to it during 2001 and 2002. By 2010 these facilities were joined by an additional seven regional hospital-based proton therapy centers in the United States alone, and many more worldwide. See [Wall Street Journal article](#) for information on the details of opening up a center.

At this time, research facilities are working on miniaturizing proton generators. An ideal technology would reduce the cost from hundreds of millions of dollars to about twenty to thirty million per device. This would make proton therapy more available and convenient for patients.



Proton Therapy versus Photon Therapy

There are many advantages to choosing proton therapy over photon therapy. Unlike photon radiation, proton beams will only penetrate to the depth of the tumor and not pass through the entire body. This reduces the overall toxicity dose. In addition, fewer treatments are required for proton therapy patients than photon therapy. Unfortunately, proton therapy is more expensive than photon therapy and less common. Once approved by a facility and medical insurance, a patient may have to move temporarily to a larger city to receive treatment. Some forms of cancer have never been clinically treated with proton therapy (namely, breast cancer). Patients desiring proton therapy might not be able to receive type of therapy due to lack of research.

Brachytherapy

The Greek term brachy means to travel a short distance. This type of radiation involves placing ionizing pellets(seeds) or rods directly at the tumor. Photons (in the form of x-ray or gamma ray) are produced inside the body and will penetrate throughout this particular area localizing the radiation. Pellets are surgically implanted while rods can be temporarily inserted to produce radiation internally. With pellets/seeds, the patient will remain radioactive as long as these devices remain inside the body. People undergoing this type of radiotherapy need to be aware of their constant emission of radiation. Radiation that is administered through rods connected to a photon device will disperse energy immediately and not leave the patient radioactive.

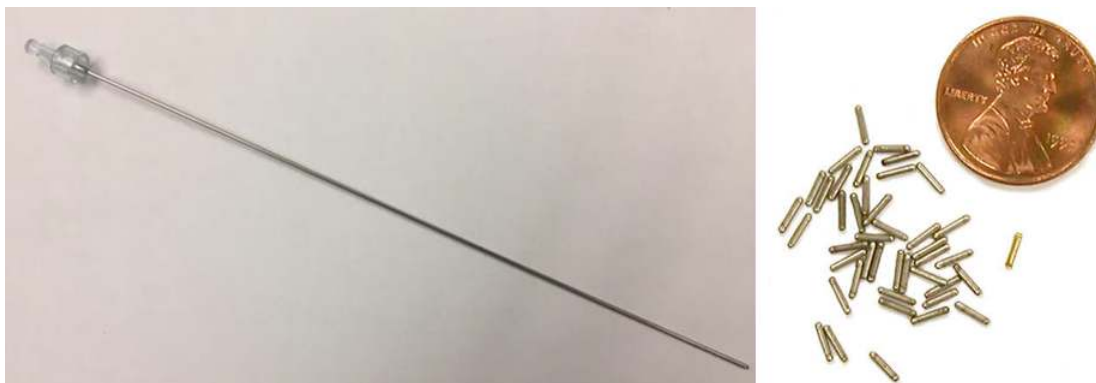


Figure 5.6.2.4: The picture on the left shows a brachytherapy rod. This device would be connected to a gamma source and connected inside a human body. The second images shows radioactive pellets or seeds that can be surgically implanted near a tumor. Note how small the pellets are when compared to an American penny. Image taken from: <https://c1.staticflickr.com/8/7254/7...74d83b7058.jpg>

Brachytherapy is widely used in the treatment of cancers involving reproductive organs. Because the radiation is isolated internally, patients are less likely to experience side effects when receiving this type of treatment. Cancers that have been treated with brachytherapy are shown below:

- Prostate
- Breast
- Esophageal
- Lung
- Uterine
- Anal/Rectal
- Sarcomas
- Head and neck

Table 5.6.2.1: Commonly used radiation sources (radionuclides) for brachytherapy.

Radionuclide	Type	Half-life
Cesium-131 (^{131}Cs)	Electron Capture, ϵ	9.7 days
Cesium-137 (^{137}Cs)	β^- - particles, γ -rays	30.17 years
Cobalt-60 (^{60}Co)	β^- - particles, γ -rays	5.26 years
Iridium-192 (^{192}Ir)	γ -rays	73.8 days
Iodine-125 (^{125}I)	Electron Capture, ϵ	59.6 days
Palladium-103 (^{103}Pd)	Electron Capture, ϵ	17.0 days
Ruthenium-106 (^{106}Ru)	β^- - particles	1.02 years
Radium-226 (^{226}Ra)	β^- - particles	1599 years

Side Effects of Radiation Therapy

Patients receiving radiation therapy can experience a variety of side effects. For example, sterility could occur if reproductive organs are irradiated. Skin that has been irradiated can appear dry and feel itchy. Some patients will lose sensation in the irradiated area. Radiation can affect the production of white and red blood cells. A reduction of white blood cells results in immunity disorders. Red blood cell loss causes anemia. Gastrointestinal issues such as diarrhea and nausea are common during radiation therapy. Some patients will lose hair as well. Lastly, dry mouth and tooth decay are prevalent during radiation treatments.

Medications are available to alleviate symptoms of radiation therapy. Narcotics can be prescribed to help alleviate intense pain. Prescription medications like Zofran and Phenergan can help with nausea. Special mouthwashes have been formulated to reduce dry mouth and cavities.

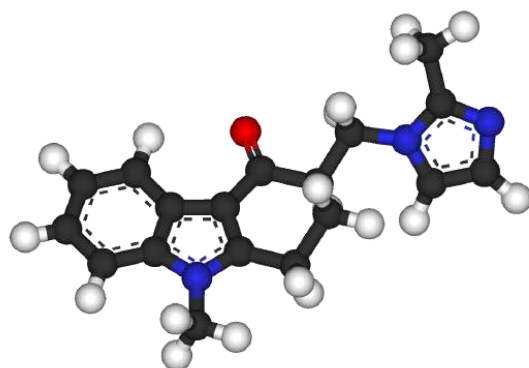


Figure 5.6.2.5: (left) Zofran is a medication used to prevent nausea and vomiting caused by cancer chemotherapy, radiation therapy, or surgery. (right) A vial of Zofran 4 mg containing ondansetron for intravenous injection. This medicine is a controlled substance and is used for nausea. Images used with permission (left: Public domain; Fvasconcellos and right: CC BY 3.0; Intropin)

Hair loss is a side effect of radiation, but only locally

Radiation therapy can cause hair loss, but hair is only lost in the area being treated. For instance, radiation to your head may cause you to lose some or all the hair on your head (even eyebrows and lashes), but if you get treatment to your hip, you won't lose the hair on your head.

Access to Radiotherapy

For those countries who do not have access to nuclear technologies, the International Atomic Energy Agency (IAEA) provides support and technical assistance to its member countries. When nations chose to join the IAEA, then they will gain nuclear materials and technologies for energy, medical, industrial, and agriculture purposes. Once accepted into the IAEA, a country must allow weapons inspectors to monitor any equipment that could be used to make nuclear weapons. The IAEA does not allow countries that do not have nuclear weapons to obtain them.



IAEA video: *Fighting Cervical Cancer: A Fair Chance for Every Woman Around the World*

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1. <https://www.cancer.gov/about-cancer/treatment/types>
2. <https://weillcornellgucancer.org/201...-cancer-cells/>
3. <http://www.proton-therapy.org/zapper.htm>
4. www.brachytherapy.com/
5. <https://www.iaea.org/>

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- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a [Creative Commons Attribution License 4.0](https://creativecommons.org/licenses/by-nc-sa/4.0/) license. Download for free at <http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110>.
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- Emma Gibney (Furman University)

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5.6.3: Non-medical Applications of Nuclear Science

Learning Objectives

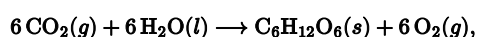
- Appreciate how radiation is used in agriculture and industry.
- Realize various isotopes and wave radiation are used as tracers and sanitizers.
- Distinguish between millimeter and back scatter technologies uses by the Transportation Security Administration (TSA).
- Know examples of foods that are irradiated in the United States.

Tracer Applications

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A **tracer** is a substance that can be used to follow the pathway of that substance through some structure. For instance, leaks in underground water pipes can be discovered by running some tritium-containing water through the pipes and then using a Geiger counter to locate any radioactive tritium subsequently present in the ground around the pipes. (Recall that tritium is a radioactive isotope of hydrogen.)

Tracers are used in diverse ways to study the mechanisms of chemical reactions in plants and animals. These include labeling fertilizers in studies of nutrient uptake by plants and crop growth, investigations of digestive and milk-producing processes in cows, and studies on the growth and metabolism of animals and plants.

For example, the radioisotope C-14 was used to elucidate the details of how photosynthesis occurs. The overall reaction is:



The process is much more complex, proceeding through a series of steps in which various organic compounds are produced. In studies of the pathway of this reaction, plants were exposed to CO_2 containing a high concentration of ^{14}C . At regular intervals, the plants were analyzed to determine which organic compounds contained carbon-14 and how much of each compound was present. From the time sequence in which the compounds appeared and the amount of each present at given time intervals, scientists learned more about the pathway of the reaction.

Airport Applications

Commercial airports use X-rays and CT imaging to identify items contained in luggage. An X-ray machine can produce images of hard and soft materials. Based upon relative densities, an X-ray image will display a different color for different types of material. For example, foods and organic materials will appear **orange** on an x-ray scan. Harder materials like thicker plastics will appear **green**. Lastly, metals and extremely thick materials will produce a **bluish color** when scanned. Click on this [video](#) for how technology for scanning bags has changed in the last five years.



Figure 5.6.3.1: Photograph of a baggage screening x-ray display. (CC BY-SA; IDuke via Wikimedia).

American airports once used X-rays to scan human bodies for weapons. Up through 2013, back-scatter technology was utilized to give very detailed body images. Due to the "naked body" pictures, this type of technology was banned in American airports and has now been replaced with millimeter wave generators. These machines produce cartoon-like images and expose the body to high-energy radio waves.

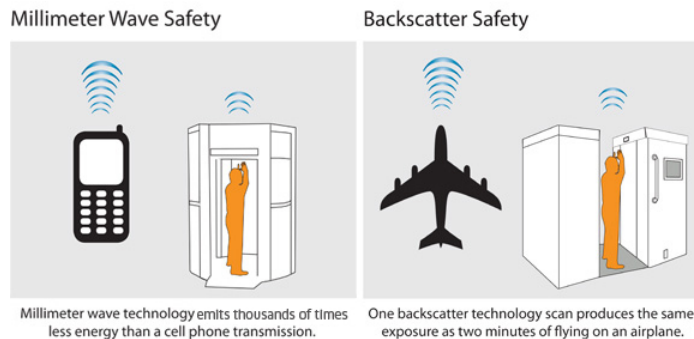


Figure 5.6.3.2: Different whole body scanning technologies have been used in airport security. (Public Domain; Transportation Security Administration via Wikipedia)

Irradiation of Food, Mail, and Water

Co-60 (gamma source) and X-rays are used to irradiate many foods in the United States. Ionizing radiation can be used to kill food-borne illnesses like *Salmonella* and *E. coli*. Irradiating food can also prolong its shelf-life, delay ripening, and destroy insects. In addition, gamma/X-ray can sterilize foods making refrigeration unnecessary. Often, sterilized food is served to hospital patients who have impaired immune systems.



In the United States, the irradiation of food is regulated by the FDA (Food and Drug Administration). Foods that have been irradiated must display the symbol called the Radura (Figure 5.6.3.3) on their packaging. Some foods that could display this symbol are: meats (beef, pork, chicken), shrimp, lobster, fruits, vegetables, shellfish, and spices. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.



Figure 5.6.3.3 Radura- symbol that irradiation has been used to sanitize food. Image taken from: commons.wikimedia.org/wiki/File:Radura-Symbol.svg

Ultraviolet (UVC) radiation has been used in the purification of drinking water. The peak wavelength for this process is around 222 nm. By using UVC, many viruses, bacteria, and other pathogens can be annihilated in air, water, and nonporous materials. UVC radiation is even powerful enough to destroy the protein coating of the SARS-CoV-2 virus. For more information on UVC and sanitation in water, click [here](#).

In the fall of 2001, anthrax-laced letters were sent to various news agencies and two United States senators. **Anthrax** is an infectious disease caused by bacteria. It exists naturally in some soils and can be isolated in a laboratory. After September 11, 2001, this substance was used as a biochemical weapon. This particular incident killed five people and injured twenty-two other individuals. Anthrax produces ulcers on the skin and causes severe breathing difficulties. A vaccine to protect against anthrax is available, but only members of the United States military are able to receive it.

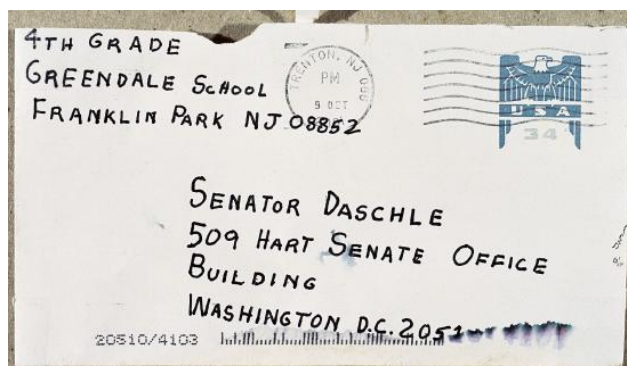


Figure 5.6.3.4: Anthrax-laced letter that was sent to Senator Tom Daschle. (Public Domain; FBI via Wikipedia)

To combat bioterrorism with anthrax, the United States Postal Service (USPS) and the Federal Bureau of Investigation (FBI) installed X-ray generators to irradiate suspicious-looking mail sent to some governmental facilities. X-rays will kill the majority of these bacteria and some viruses as well. Once irradiated, mail (packages and letters) might change in color, texture, and smell. This ionizing radiation chemically alters the composition of the mail's paper component. The X-rays do not leave the mail radioactive.

Smoke Detectors

Americium-241, an α emitter with a half-life of 458 years, is used in tiny amounts in ionization-type smoke detectors (Figure 5.6.3.5). The α emissions from Am-241 ionize the air between two electrode plates in the ionizing chamber. A battery supplies a potential that causes movement of the ions, thus creating a small electric current. When smoke enters the chamber, the movement of the ions is impeded, reducing the conductivity of the air. This causes a marked drop in the current, triggering an alarm.

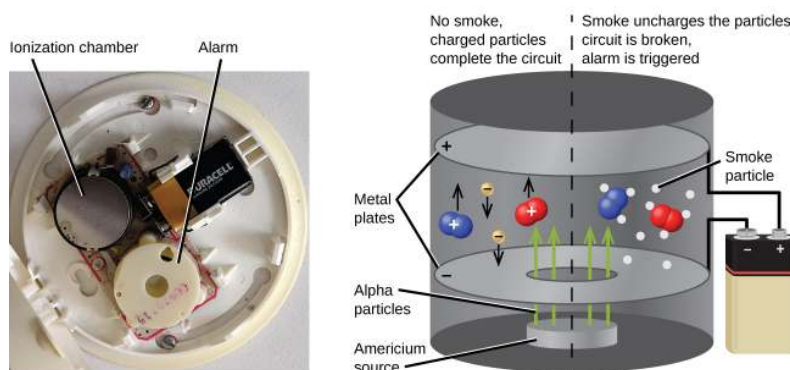


Figure 5.6.3.5: Inside a smoke detector, Am-241 emits α particles that ionize the air, creating a small electric current. During a fire, smoke particles impede the flow of ions, reducing the current and triggering an alarm. (credit a: modification of work by "Muffet"/Wikimedia Commons)

References

1. <https://medlineplus.gov/anthrax.html>
2. <https://www3.epa.gov/radtown/mail-irradiation.html>

Contributors

- Elizabeth R. Gordon (Furman University)
- Emma Gibney (Furman University)

- Erin Avram (Cleveland State University)

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5.6.4: Uses and Consequences of Nuclear Energy

Learning Objective

- Know the benefits and consequences of nuclear energy.

Nuclear Power Generation

The generation of electricity is critical for the operation of businesses, health care delivery, schools, homes, and other areas that require electrical power. According to 2011 statistics, coal is used for **42%** of the total power generated, with natural gas being employed for another **25%**. Nuclear power plants are employed in about **19%** of the cases, with renewable energy sources supplying the last **13%**. All of these fuels are used to heat water to generate steam. The steam then turns a turbine to generate electricity.

Nuclear power plants use the energy from nuclear fission to produce electricity. **U-235** is the preferred nuclear fuel because when its atoms are split (fissioned), they not only emit heat and high-energy radiation but also enough neutrons to maintain a chain reaction and provide energy to power a nuclear power plant. Uranium is found in rocks all over the world but is relatively rare and the supply is finite making it a nonrenewable energy source.

Uranium usually occurs in combination with small amounts of other elements and once it is mined, the U-235 must be extracted and processed before it can be used as a fuel in a nuclear power plant to generate electricity. The processed uranium is formed into fuel rods and then bundled into fuel assemblies. Fuel assemblies are stored onsite until the reactor operators need them. At this stage, the uranium is only mildly radioactive, and essentially all radiation is contained within the metal tubes. When needed, the fuel is loaded into a reactor core (Figure 5.6.4.1). Typically, about one third of the reactor core (40 to 90 fuel assemblies) is changed out every 12 to 24 months.

The most common type of reactors are the **pressurized water reactors (PWR)** (Figure 5.6.4.1) in which water is pumped through the reactor core and heated by the fission process. The water is kept under high pressure inside the reactor so it does not boil. The heated water from the reactor passes through tubes inside the steam generator where the heat is transferred to water flowing around the tubes in the steam generator. The water in the steam generator boils and turns to steam. The steam is piped to the turbines. The force of the expanding steam drives the turbines, which spin a magnet in a coil of wire – the generator – to produce electricity.

After passing through the turbines, the steam is converted back to water by circulating it around tubes carrying cooling water in the condenser. The condensed steam – now water – is returned to the steam generators to repeat the cycle.

The three water systems (condenser, steam generator, and reactor) are separate from each other and are not permitted to mix. Water in the reactor is radioactive and is stays within the containment structure whereas water in the steam generator and condenser is nonradioactive.

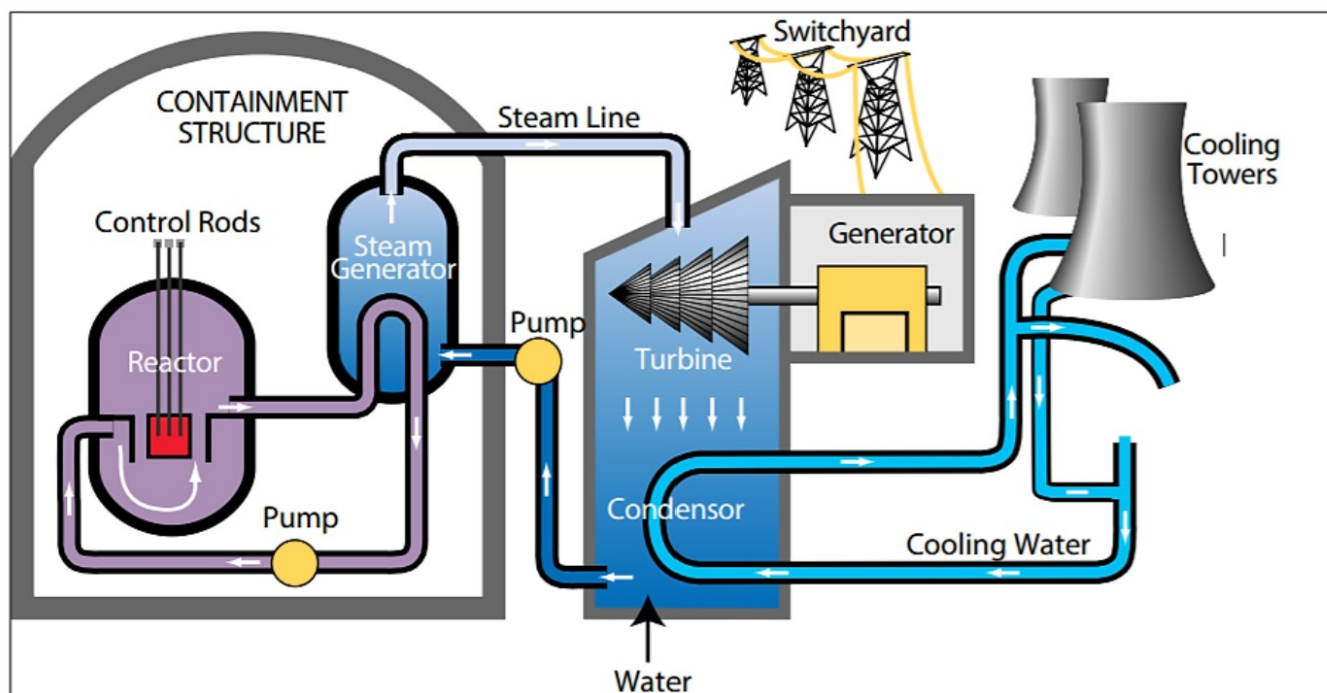


Figure 5.6.4.1: A schematic diagram of a pressurized water reactor (PWR), the most common type of nuclear reactor. Diagram from Tennessee Valley Authority (public domain). www.tva.com

Benefits of Nuclear Energy

By using fission, nuclear power plants generate electricity without emitting air pollutants like those emitted by fossil fuel-fired power plants. This means that financial costs related to chronic health problems caused by air pollutants such as particulate material, carbon monoxide, nitrogen oxides and ozone, among others, are significantly reduced. In addition, nuclear reactors do not produce carbon dioxide while operating which means that nuclear energy does not contribute to the global warming problem.

Another benefit of nuclear energy over fossil fuels especially coal is that uranium generates far more power per unit weight or volume. This means that less of it needs to be mined and consequently the damage to the landscapes is less especially when compared to the damage that results from coal mining such as mountaintop removal.

Nuclear power can also be used to propel ships. The turbine can be connected to a propeller system. The rotating turbine shaft will turn the propeller to move the ship.



Figure 5.6.4.2 Nuclear submarine.

The Drawbacks of Nuclear Energy

The main environmental concern related to nuclear power is the creation of radioactive wastes such as uranium mill tailings, spent (used) reactor fuel, and other radioactive wastes. These materials can remain radioactive and dangerous to human health for thousands of years. Radioactive wastes are classified as low-level and high-level. By volume, most of the waste related to the nuclear power industry has a relatively low-level of radioactivity. **Uranium mill tailings** contain the radioactive element radium, which decays to produce **radon**, a radioactive gas. Most uranium mill tailings are

placed near the processing facility or mill where they come from. Uranium mill tailings are covered with a barrier of material such as clay to prevent radon from escaping into the atmosphere, and they are then covered by a layer of soil, rocks, or other materials to prevent erosion of the sealing barrier.

Spent fuel rods contain a variety of products, consisting of unstable nuclei ranging in atomic number from 25 to 60, some transuranium elements, including plutonium, americium, and unreacted uranium isotopes. The unstable nuclei and the transuranium isotopes give the spent fuel a dangerously high level of radioactivity. The long-lived isotopes require thousands of years to decay to a safe level. The ultimate fate of the nuclear reactor as a significant source of energy in the United States probably rests on whether or not a politically and scientifically satisfactory technique for processing and storing the components of spent fuel rods can be developed.

The other types of **low-level radioactive waste** are the tools, protective clothing, wiping cloths, and other disposable items that get contaminated with small amounts of radioactive dust or particles at nuclear fuel processing facilities and power plants. These materials are subject to special regulations that govern their handling, storage, and disposal so they will not come in contact with the outside environment.

High-level radioactive waste consists of spent nuclear reactor fuel (i.e., fuel that is no longer useful for producing electricity). The spent reactor fuel is solid and consists of small pellets in long metal tubes called rods. Spent reactor fuel assemblies are initially stored in specially designed pools of water, where the water cools the fuel and acts as a radiation shield. Spent reactor fuel assemblies can also be stored in specially designed dry storage containers. An increasing number of reactor operators now store their older spent fuel in dry storage facilities using special outdoor concrete or steel containers with air cooling. Currently, there is no permanent disposal facility in the United States for high-level nuclear waste.

When a nuclear reactor stops operating, it must be **decommissioned**. This involves safely removing the reactor and all equipment that has become radioactive from service and reducing radioactivity to a level that permits other uses of the property. The U.S. Nuclear Regulatory Commission has strict rules governing nuclear power plant decommissioning that involve the cleanup of radioactively contaminated plant systems and structures, and the removal of the radioactive fuel.

A **nuclear meltdown**, or uncontrolled nuclear reaction in a nuclear reactor, can potentially result in widespread contamination of air and water. Some serious nuclear and radiation accidents have occurred worldwide. The most severe accident was the Chernobyl accident of 1986 in the then Soviet Union (now Ukraine) which killed 31 people directly and sickened or caused cancer in thousands more. The Fukushima Daiichi nuclear disaster (2011) in Japan (Figure 5.6.4.3) was caused by a 9.0 magnitude earthquake that shut down power supply and a tsunami that flooded the plant's emergency power supply. This resulted in the release of radioactivity although it did not directly result in any deaths at the time of the disaster. Another nuclear accident was the Three Mile Island accident (1979) in Pennsylvania, USA. This accident resulted in a near-disastrous core meltdown that was due to a combination of human error and mechanical failure, but did not result in any deaths. No cancers or other health impacts have been found in follow-up studies of this accident. While there are potentially devastating consequences to a nuclear meltdown, the likelihood of one occurring is extremely small. After every meltdown, including the 2011 Fukushima Daiichi disaster, new international regulations were put in place to prevent such events from occurring again.



Figure 5.6.4.3 (a) After the accident, contaminated waste had to be removed, and (b) an evacuation zone was set up around the plant in areas that received heavy doses of radioactive fallout. (credit a: modification of work by "Live Action Hero"/Flickr)

The processes for mining and refining uranium ore and making reactor fuel require large amounts of energy. Nuclear power plants have large amounts of metal and concrete, which also require large amounts of energy to manufacture. If fossil fuels are used for mining and refining uranium ore or in constructing the nuclear plant, then the emissions from burning those fuels could be associated with the electricity that nuclear power plants generate.



Video 5.6.4.1 *The design and safe operation of a nuclear reactor.*

Summary

- The importance of nuclear power in generating electricity is described.
- The operation of a nuclear power plant is described.
- The drawbacks of nuclear energy is described.

Contributors and Attributions

- CK-12
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- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a [Creative Commons Attribution License 4.0](https://creativecommons.org/licenses/by-nc-sa/4.0/) license. Download for free at <http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110>.
- Marisa Alviar-Agnew (Sacramento City College)

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5.6.5: Three Mile Island Nuclear Accident

The importance of regulating cooling and containment systems at nuclear power generating stations are amply illustrated by three major accidents that occurred with the nuclear reactors in the United States (Three Mile Island), the former Soviet Union (Chernobyl), and Japan (Fukushima). Other nuclear incidents have occurred around the world that have involved the storage or transportation of nuclear materials.

Three Mile Island (Harrisburg, Pennsylvania, March 28, 1979)

The two reactors housed at this facility are located in Middletown, Pennsylvania. This site is approximately 10 miles southeast of Harrisburg, Pennsylvania. Both units are PWRs (pressurized water reactors) and were commissioned in the mid to late 1970s. Before the accident, each of the Three Mile Island (TMI) reactors generated around 800 MW of electrical power. The smaller cylindrical buildings in the photograph below are reactors named Units 1 and 2. The four larger structures in this picture are the cooling towers that allow excess heat to be transferred to the atmosphere.



Figure 5.6.5.1: Color photograph of the Three Mile Island nuclear- generating station, which suffered a partial meltdown in 1979. The reactors are in the smaller domes with rounded tops (the large smokestacks are just cooling towers). (United States Department of Energy)

On March 28, 1979, Three Mile Island's Unit 2 (TMI-2) reactor lost cooling water in the secondary loop (note the blue loop in the figure below). This event occurred around 4:00 a.m. and alarms sounded in the control room immediately. Within one minute, the reactor automatically shut down. Meanwhile, water in the primary loop (yellow tube) overheated and the pressure of the reactor increased. To relieve the excess pressure inside the reactor, the pressure relief valve (top yellow/red tube) opened up. Backup pumps restarted the water in the secondary loop to assist the cooling efforts.

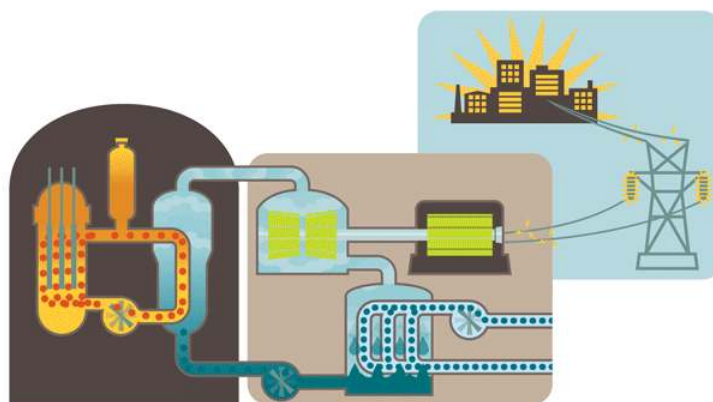


Figure 5.6.5.2: The following animated diagram graphically depicts the sequence of events associated with the accident at TMI-2. (United States Nuclear Regulatory Commission; Public Domain).

In the control room, operators noted the decrease in pressure of the reactor and a reduction of cooling water in the primary loop. Within fifteen minutes of the original malfunction, approximately 3000 gallons of cooling water had evaporated and exited the cooling towers. The reactor components (fuel and control rods) became exposed due to the lack of water. This caused excess heat

and corresponding melting of materials to various degrees. Back in the control room, the instrumentation indicated incorrectly that the pressure valve was closed, even though it was stuck in the open position. Operators attempted to cool down the control and fuel rods by pumping more water into the primary loop of the reactor. Unfortunately, since the pressure valve was stuck open, most of the water converted to steam. This made temperatures increase greatly in the reactor core.



Figure 5.6.5.3: President Jimmy Carter leaving Three Mile Island for Middletown, Pennsylvania., 04/01/1979. (<https://catalog.archives.gov/id/540021>)

By 6:22 a.m., the operators closed a manual back-up valve before the pressure release valve. This allowed water to remain in the primary loop. Unfortunately, most of the water in the core was in the form of steam and did not circulate efficiently throughout the core. In addition, the steam reacted with metals in the control and fuel rods. These reactions produced combustible hydrogen gas which rose to the top of the reactor core. The operators turned on the pumps to inject more water into the reactor. By 7:50 a.m., all cooling water was restored to the core successively, but a large hydrogen bubble hovered over the top of the reactor. At this time, operators were fearful the bubble might ignite and then explode the core to the outside environment. To prevent this explosion from happening, operators and engineers vented the hydrogen gas outside the nuclear reactor into the atmosphere. This venting occurred from March 30 until April 1, 1978.

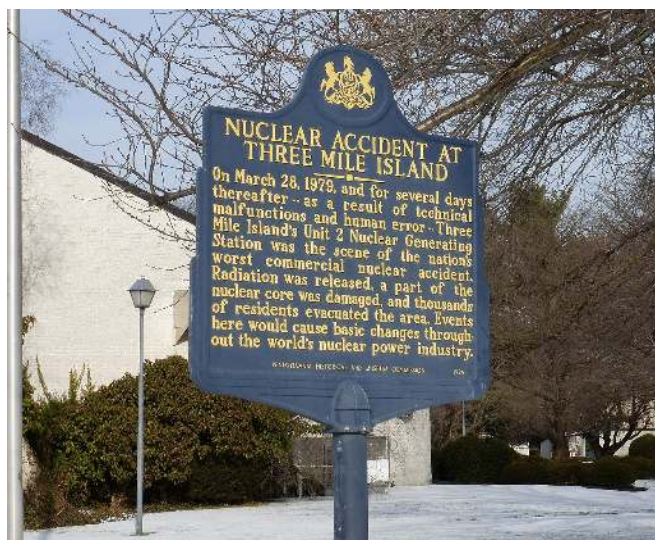


Figure 5.6.5.4: A sign in Middletown, Pennsylvania dedicated in 1999 describing the Three Mile Island accident, the evacuation of the residents in the area, and the impacts to the nuclear power industry. (Wikipedia Commons)

Residents in nearby communities were informed of a nuclear issue around 7:24 a.m. on March 28, 1979. Radio and television broadcasts started within the next hour. By 12:45 p.m., representatives of the Department of Energy (DOE) arrived to evaluate the current status of the reactor. DOE agents put together a plan of action to keep the reactor under control. Numerous state governmental officials met with the DOE and devised safety protocols (remain indoors) and emergency evacuation plans if the hydrogen bubble were to burst. To see a comprehensive timeline of these events, please click [here](#).

In the venting process, a minimal amount of radioisotopes were released to the areas surrounding the TMI nuclear power station. Several independent research facilities have quantified the radiation that was released during the accident. These studies state that residents within 10 miles of TMI were exposed to radiation within the range of 0.08 mSv to 1 mSv. This level of radiation is comparable to having a chest X-ray exam (Table 5.6.5.2).

Table 5.6.5.1: Fission products vented to the atmosphere from TMI 2

Isotope	Emits	Half-life ($t_{1/2}$)
Kr-85	beta particles, gamma rays	10.8 years
Cs-137	beta particles, gamma rays	30.2 years
I-131	beta particles, gamma rays	8.0 days
Xe-133	beta particles, gamma rays	5.3 days

In 1984, nuclear scientists were able to survey the destruction of the reactor core. They concluded that approximately 1/3 of the fuel rods melted inside the reactor. This accident was labeled a partial meltdown (full meltdown would indicate the majority of fuel rods were melted). The contaminant structure remained intact and protected the citizens of the area.

The clean-up process for TMI unit 2 started in August of 1979. It was a twelve-year project that employed over 1000 people. All radioactive fuel and water were shipped to existing nuclear waste storage facilities in the United States (Washington State and Idaho). The containment building for this reactor still remains and houses waste for the TMI unit 1. The approximate clean-up price for this accident was \$973 million US dollars.

The accident at Three Mile Island halted all construction of new nuclear power plants in the United States. Many citizens became fearful of nuclear power in general and became supportive of other types of energy sources. Films like [The China Syndrome](#) catalyzed this feeling for decades after this catastrophe.



Video 5.6.5.1: Description of TMI Unit 2 partial meltdown.

At the time of the partial meltdown of TMI Unit 2, Unit 1 was shut down for refueling. Despite much opposition, Unit 1 was restarted in the Fall of 1985 and has been providing power to the residents of this area until September 2019. Funding could not be secured to keep this unit running due to the relatively low cost of fossil fuels.

✓ Example 5.6.5.1

The average radiation exposure due to the TMI Unit 2 accident was between .08mSv and 1 mSv. Calculate this dosage in rem and mrem. Compare these values to average background radiation (360 mrem) and Acute Radiation Syndrome (ARS, 25 rem). For a related example, see [5.8.1](#)

Solution

For the first calculation, recall the metric prefix milli from earlier this semester.

$$\frac{.08 \cancel{mSv}}{1} \times \frac{1 Sv}{1000 \cancel{mSv}} = 8 \times 10^{-5} Sv$$

For the second calculation, you will need to insert the 100 rem = 1 Sv factor. Then, you will need to use the milli metric prefix again.

$$\frac{8 \times 10^{-5} \cancel{Sv}}{1} \times \frac{100 \cancel{rem}}{1 \cancel{Sv}} \times \frac{1000 mrem}{1 \cancel{rem}} = 8 mrem$$

Comparing the calculated mrem level to background radiation, one can see that this is a small concentration of radiation. As for the ARS level, the first calculated rem value is significantly lower than 25 rems.

Sources

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2. Larabee, Ann. Decade of Disaster. Chicago: Board of Trustees of the University of Chicago, 2000.
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Contributors and Attributions

- Abheetinder Brar (UCD)

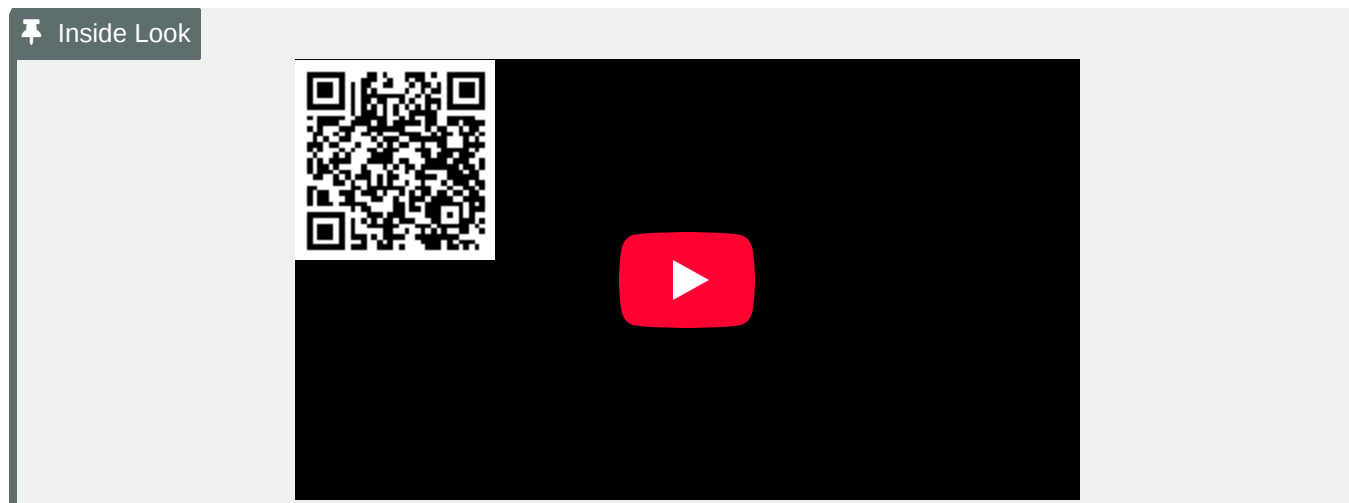
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5.6.6: Chernobyl Nuclear Disaster

Chernobyl (Pripyat, Ukraine, April 26, 1986)

Another major nuclear accident involving a reactor occurred in April 1986, at the Chernobyl Nuclear Power Plant in Ukraine, which was still part of the former Soviet Union. While operating at low power during an unauthorized experiment with some of its safety devices shut off, one of the reactors at the plant became unstable. Its chain reaction became uncontrollable and increased to a level far beyond what the reactor was designed for. The steam pressure in the reactor rose to between 100 and 500 times the full power pressure and ruptured the reactor. Because the reactor was not enclosed in a containment building, a large amount of radioactive material spewed out, and additional fission products were released as the graphite (carbon) moderator of the core ignited and burned.



Video 5.6.6.1: High detail 3D animations and explanations of the inner workings of Chernobyl nuclear power station. Showing why it was so vulnerable to blowing itself up and how it was different from western reactors.

Watch the video above (8:42 in length) and answer the questions below:

1. How many working reactors were at the Chernobyl Power Plant site? How many more were under construction?
2. Which unit experienced a full meltdown?
3. What type of fuel and moderator did Chernobyl use?
4. Describe the positive void coefficient. Are Western reactors designed to experience positive void coefficients?
5. Compare the basic core size of Fukushima to Chernobyl.
6. Compare the basic containment layers of Fukushima to Chernobyl.
7. What is the purpose of leaving the top of an RBMK reactor open? Do Western reactors have this design?
8. What events led to the explosion of Chernobyl? Was this a chemical or a nuclear explosion?
9. Name the two countries that were most affected by the Chernobyl meltdown.
10. Did the people who cleaned up Chernobyl have special training in radiation? What is the nickname of these individuals?
11. How fast did the USSR construct the first sarcophagus? Why did it need to be replaced?
12. Look below to the section of Today's Chernobyl. The EU constructed this massive structure. How much did this cost and how long did it take to build?
13. Does Russia still use RBMKs?
14. Describe the ghost town of Pripyat today. Just before the accident, an amusement park was constructed to celebrate May Day. Did the people of Pripyat get to enjoy the Ferris Wheel? Would you have gotten on that Ferris Wheel during that time (assuming you were on this planet)?

Basics of the Disaster

On April 26, 1986, a test was scheduled at the Chernobyl Nuclear Power Plant to evaluate residual cooling capacity in the event of a power grid failure. At the time of the accident, the Chernobyl facility used four RBMK (Reaktor Bolshoy Moshchnosty Kanalny or high power channel) reactors to produce a total of 4000 MW of energy. Two more reactors were being constructed to produce additional wattage. RBMKs not only produce a large amount of energy but also produce weapons-grade Pu-239. The excessive power, design, and materials of these reactors have discouraged many countries (including the United States) from building them.



Video 5.6.6.2: Video from the 2019 HBO miniseries "Chernobyl."

Components of the Chernobyl RBMK reactors.

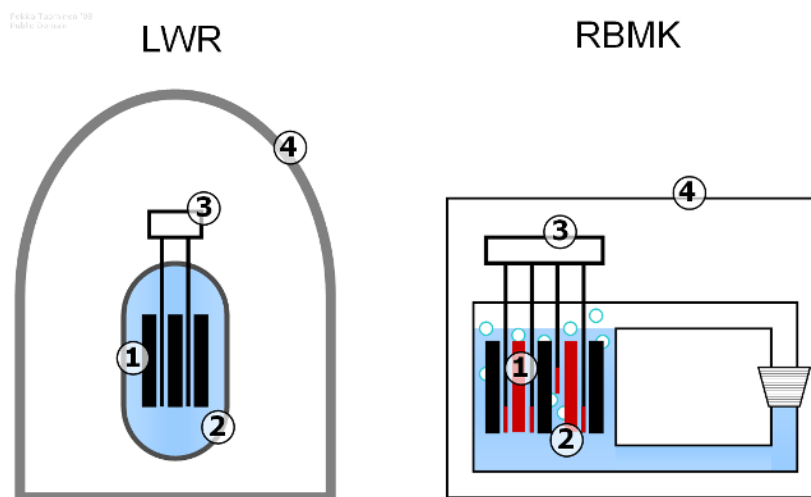


Figure 5.6.6.1: The critical differences in the RBMK reactor compared to a LWR (Light Water Reactor) that directly contributed to the Chernobyl disaster. 1. The flammable graphite moderator in the reactor core that burned in the fire, 2. The positive void coefficient in the water made possible the power peak that blew the reactor vessel, 3. The control rods were very slow, they took 18-20 seconds to be inserted into the reactor. Moreover they had graphite tips that actually intensified the fission chain reaction in the beginning of the insertion. 4. No containment building at all, data from IAEA: [INSAG-7 The Chernobyl Accident: Updating of INSAG-1](#). Vienna, 1992; NEA: [Chernobyl - Assessment of Radiological and Health Impacts](#). Paris, 2002; Nakao, M.: [Chernobyl Accident](#). University of Tokyo, 2006.

Unlike light water reactors, RBMK reactors use graphite moderators to slow the neutrons of the fission process. This type of material can combust easily. The less enriched U-235 fuel of this reactor produces Pu-239 to be produced. Unfortunately, the Chernobyl reactor technicians felt that removing the new Pu-239 fuel would be too difficult with a full containment covering of unit 4. The new Pu-239 material would be extracted while the reactor continued to produce energy. The high power wattage, combustible moderator, and open containment convinced many other countries to steer clear of building RBMKs. When comparing the differences of RBMK reactors to those typically used in most other countries, view the table below:

Table 5.6.6.1: Types of Nuclear Reactor

Type of Reactor	Containment	Fuel	Coolant	Moderator
PWR	repetitive layers of lead/concrete and the top is closed inside the reactor building	slightly enriched U-235 (2-4%)	LW = light water	LW = light water
BWR	repetitive layers of lead/concrete and the top is closed inside the reactor core building	slightly enriched U-235 (2-4%)	LW = light water	LW = light water
RBMK	layers of lead/concrete and the top is open inside the reactor core building to allow extraction of rods while the reactor is producing energy	low enriched U-235 (<2%)	LW = light water	Graphite (carbon)

Before the accident

Nuclear reactors require active cooling in order to remove the heat generated by radioactive decay. Even when not generating power, reactors still generate some heat, which must be removed in order to prevent damage to the reactor core. Cooling is usually accomplished through the flow of fluids. In the case of Chernobyl, this fluid was water.

The problem at the Chernobyl plant was that following an emergency shutdown of all power, diesel generators were needed to run the cooling pumps. These generators took about a minute to attain full speed, which was deemed an unacceptably long time for the reactor to be without cooling. It was suggested that the rotational momentum of the winding down steam turbine be used to power the pumps in the time between shutdown and the generators being ready. A test was devised to test this method in 1982, but the turbine was not able to provide the required voltage as it spooled down. Two more tests were conducted in the following years, but were also unsuccessful. The fourth test was scheduled to be run on April 25, 1986.



Figure 5.6.6.2: Interior view of the control room of Chernobyl Nuclear Power Plant unit 3. Over 3,000 people continue to work at Chernobyl to monitor nuclear fuel and carry out the decommissioning of the facility. (www.iaea.org/NewsCenter/Multi...bank/index.js)

The experiment was devised in such a way that if it had gone as planned, the disruption and danger to the plant would be very minimal. First, the reactors would be brought down to low power, between 700 and 800 megawatts. Then the steam turbine would be run up to full speed and then turned off. The power generated by the winding down generators would then be measured to determine if it was sufficient to power the cooling pumps in the time before the diesel generators got up to full speed.

By 1986, the plant had been running for two years without the implementation of a method to keep the cooling pumps running continuously following an emergency shutdown. This was an important safety measure that the plant was lacking, which presumably gave the plant managers a considerable amount of urgency in completing another test.

Preparing for the test

The experiment was scheduled to run during the day shift of 1985, while the night shift would only have to maintain cooling of the radioactive decay in the shut-down plant. However, another power generator nearby unexpectedly shut down, necessitating the need for the Chernobyl plant to delay the test and continue producing power. The experiment would be resumed at 11:04 PM, by which time the day shift had departed and the evening shift was about to leave. This meant that the experiment would be conducted in the middle of two shifts, leaving very little time for the night shift employees to be briefed about the experiment and told what to do.

The power reduction of reactor 4 to 700 MW was accomplished at 00:05 AM on the 26th of April. However, the natural production of a neutrino absorber, Xenon-135, led to a further decrease in power. When the power dropped to about 500 MW, the night shift operator committed an error and inserted the reactor control rods too far. This caused the reactor to go into a near-shutdown state, dropping power output to around 30 MW.

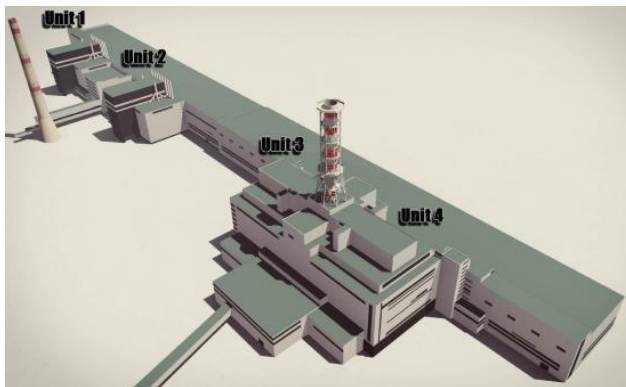


Figure 5.6.6.3: Layout of four working RBMK reactors at the Chernobyl Nuclear Power plant. (Copyright; By Tadpolefarm - Own work, CC BY-SA 4.0, <https://commons.wikimedia.org/w/index.php?curid=81169789>)

Since this was too low for the test, it was decided to restore power by extracting the control rods. Power would eventually rise and stabilize at around 200 MW.

The operation of the reactor at such a low power level would lead to unstable temperature and flow. Numerous alarms and warnings were recorded regarding the emergency measures required to keep the reactor stable. In the time between 12:35 and 12:45 AM, alarm signals for the thermal-hydraulic parameters were ignored in order to preserve the reactor's power level.

The test continued, and at 1:05 AM extra water pumps were activated in order to increase the water flow. The increased coolant flow rate led to an increase in the coolant temperature in the core, reducing the safety margin. The extra water flow also led to a decrease in the core's temperature and increased the neutron absorption rate, decreasing the reactor's power output. Operators removed the manual control rods in order to maintain power.

All these actions led to the reactor being in an unstable state that was clearly outside safe operation protocol. Almost all the control rods had been removed, which reduced the effectiveness of inserting safety rods in an emergency shutdown. The water was very close to boiling, which meant that any power increase would cause it to boil. If it started boiling, it would be less effective at absorbing neutrons, further increasing the reactor's power output.

Conducting the test

The experiment was started at 1:23:04 AM. The steam to the turbines was shut off, causing the turbines to start spooling down. Four of the eight cooling pumps were also shut down. The diesel generator was started and began powering the cooling pumps after at 1:23:43. Between this time, the four pumps were powered by the slowing steam turbines. As the turbines slowed down, their power output decreased, slowing the cooling pumps. This led to the increased formation of steam voids in the core, reducing the ability of the cooling water to absorb neutrons. This increased the power output of the reactor, which caused more water to boil into steam, further increasing the reactor's power. However, during this time the automatic control system was successful in limiting power increase through the insertion of control rods.

Video 5.6.6.3: [Chernobyl nuclear accident why happened? Animated Demonstration](#)

At 1:23:40, a button was pressed that initiated the emergency shutdown of the reactor and the insertion of all control rods. It is believed that this was done as a routine method to shut down the reactor to conclude the experiment and not as an emergency measure.

The process of inserting the control rods was initiated, but it took about 20 seconds for the rods to be completely inserted. A flawed design in the graphite-tip control rod meant that coolant was displaced before the neutron-absorbing material could be fully inserted and slow down the reaction. This meant that the process of inserting the control rods actually increased the reaction rate in the lower half of the core.



Video 5.6.6.2: Trailer video from the 2019 HBO miniseries "Chernobyl."

A massive power spike occurred, causing the core to overheat. Some of the fuel rods fractured, causing the control rods to become stuck before they were fully inserted. Within three seconds the core's power output rose above 500 MW. According to the simulation, it is estimated that power output then rose to 30 GW, ten times the normal power output. This was caused by the rising power output causing massive steam buildup, which destroyed fuel elements and ruptured their channels.

It is not possible to know precisely what sequence of events led to the destruction of the reactor. It is believed that the steam buildup entered the reactor's inner structure and lifted the 2000 ton upper plate. This steam explosion further ruptured fuel channels, resulting in more coolant turning into steam and leaving the reactor core. This loss of coolant further increased the reactor's power. A nuclear excursion (an increasing nuclear chain reaction) caused a second, even more, powerful explosion.

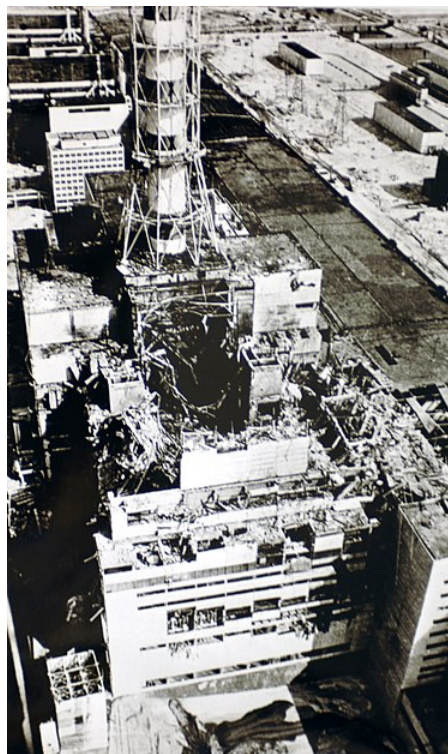


Figure 5.6.6.4: A new reckoning of time began on 26 April 1986 At 1:23:45 AM. This photo was taken from a helicopter on the day following the explosion of the Chernobyl #4 reactor. [IAEA Imagebank - 02790015](#)

The explosion destroyed the core and scattered its contents in the surrounding area, igniting the red-hot graphite blocks. Against safety regulations, a flammable material, bitumen, had been used in the roof of the reactor. When this was ignited and scattered into the surrounding area, it started several fires on reactor 3. Those working there were not aware of the damage that had been done and continued running the reactor until it was shut down at 5:00 AM.

Radiation Levels and Fallout

In the worst-hit parts of the reactor building, radiation levels were high enough to cause fatal doses in a matter of minutes. However, all dosimeters available to the workers did not have the ability to read radiation levels so high and thus read "off-scale." Thus, the crew did not know exactly how much radiation they were being exposed to. It was assumed that radiation levels were much lower than they actually were, leading the reactor crew chief to believe that the reactor was still intact. He and his crew tried to pump water into the reactor for several hours, causing most of them to receive fatal doses of radiation.

Since containment failed at the Unit 4 Reactor, particle and wave radiation emerged from the burning power plant. Winds carried the fallout throughout Europe. For more information on radioactive fallout, refer to this text: [Nuclear Fallout and Weapons Testing](#).

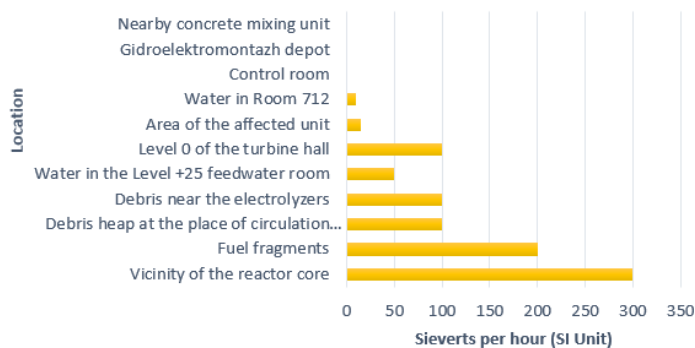


Figure 5.6.6.5: Chernobyl radiation map from CIA handbook. (http://www.lib.utexas.edu/maps/commonwealth/chornobyl_radiation96.jpg and File:Tchernobyl_radiation_1996.svg for the vector version)

Concentration of Radiation around Unit 4

Radiation is measured with a variety of units. In order to understand biological damage, click on this previous [OER page](#) and read about Sieverts (Sv) and Rem (rem).

Approximate radiation levels in and around Unit 4 shortly after the explosion



Source B. Medvedev (June 1989). "JPRS Report: Soviet Union Economic Affairs Chernobyl Notebook"

Figure 5.6.6.6: Čestina: Přibližné stupně radiace krátce po explozi 4. reaktoru v černobylské jaderné elektrárně. https://commons.wikimedia.org/w/index.php?title=User:Sandra_K%C5%99%C3%AD%C5%BEov%C3%A1&action=edit&redlink=1

Evacuation of Pripyat

Pripyat, a city nearby the power plant, was not immediately evacuated. At first, the government denied that the reactor had exploded and insisted that it was only a small accident. By April 27, though, investigators were forced to acknowledge that the reactor had exploded and ordered Pripyat to be immediately evacuated. The process of moving over 40,000 people from this area was started around 2 p.m. following the nuclear explosion. Over 1200 buses from Kyiv, transported this city's inhabitants beyond the exclusion zone.



Video 5.6.6.3: Message of evacuation the city of Pripyat (1986)

On April 28, Sweden's Forsmark Nuclear Power plant detected the presence of I-131 on their radiation detectors. Late that night, Moscow news reported that an accident had occurred at the Chernobyl Nuclear Power Plant. Within an hour of this announcement, Denmark had also detected isotopes that are only created during nuclear fission.

Fatalities and Health Effects

At the time of the accident, around 600 people were employed at the Chernobyl Nuclear Power Plant. When the explosion occurred, two workers immediately died. The World Health Organization (WHO) documented the radiological health impact for two-decades following this incident. They determined that approximately 134 Chernobyl employees were exposed to high-dose radiation (0.8 - 1.6 Gray). These workers were diagnosed with Acute Radiation Sickness (ARS). Of the 134 ARS victims, 28 of these individuals passed away within three months of the accident.



Figure 5.6.6.7: Erected near the Chernobyl site is a monument to the "liquidators", an estimated quarter of a million soldiers and citizens enlisted as the 1986 accident's "first responders". (Chernobyl, Ukraine, 1991) (Petr Pavlicek/IAEA)

In the 2011 report entitled, "Health effects due to radiation from the Chernobyl accident," the United Nations Scientific Committee on the Effect of Atomic Radiation (UNSCEAR), stated that 530,000 of the reactor's recovery workers were exposed to ionizing radiation dosages between 20-500 mSv. This test group was initially studied from April 1986 through 1990. Currently, health monitoring of these individuals continues. Studies have determined that workers who were exposed to over 200mSv were more likely to develop leukemia. Radiation-induced cataracts were noted in workers exposed to levels lower than 200 mSv.

Over 115,000 people who lived in Ukraine and Belarus evacuated after the meltdown of Unit 4. Radiation levels reported for this particular group were around 30 mSv. As for other parts of Europe, radiation exposure from the Chernobyl meltdown was considered to be less than yearly background radiation (1 mSv).

UNSCEAR reported a significant increase in the diagnosis of thyroid cancer, often attributed to exposure to I-131. Over 6,000 people (with less than 1% mortality rate) in Belarus and the Ukraine have developed this cancer during the study. Exposure to this

radioactive isotope could have been due to direct fallout exposure or indirect means (drinking milk from cows who ingested I-131 or eating foods grown in radioactive soil).

There seems to be little or no effect in regards to fertility impairment with any of the evacuees. However, residents of Belarus and the Ukraine have reported more birth defects since the accident occurred. UNSCEAR stated that this increase could be because these individuals are more likely to report all birth defects after the meltdown than they were before the accident occurred. Due to the nature of the accident and evacuation effort, many citizens of Belarus and Ukraine have reported an increase in mental health conditions.

Temporary Containment

After the accident, the reactor was encapsulated in steel and concrete, a now-decaying structure known as the sarcophagus. The construction of this covering occurred in May of 1986 and was completed by November of that same year. Not long after construction, the protective covering cracked, causing the release of additional radiation into the environment.



Figure 5.6.6.8: The Chernobyl reactor #4 building as of 2006, including the later-built sarcophagus and elements of the maximum-security perimeter. (<https://www.flickr.com/photos/83713082@N00/535916329/>; [Carl Montgomery](#))

Today's Pripyat

Today, radiation levels are still higher than normal in the areas surrounding the plant but have dropped considerably from the levels twenty years ago. It is now considered safe to visit the areas immediately surrounding the plant for short periods of time. However, it is estimated that it will take 20,000 years for reactor 4's core to be completely safe.



Figure 5.6.6.9: Montage of Prip'yat(Copyright; Wikiwind (montage); (WT-en) RealLeo at English Wikivoyage (Prip'yat CentralSquare.jpg); Jorge Franganillo from Barcelona, Spain (Prip'yat (38307778522).jpg, Prip'yat welcome sign (38071252145).jpg, Prip'yat (38338630751).jpg); Alexander Blecher, blecher.info (Chernobyl Exclusion Zone (2015) 68.JPG); Shanomag (The Abandoned Sports Hall in Prip'yat - Chernobyl.jpg).)

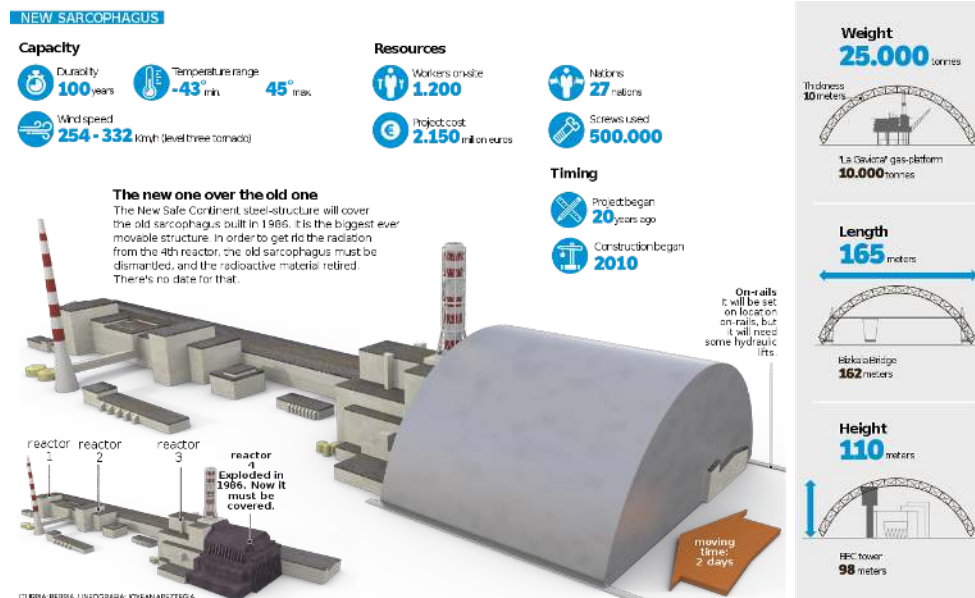


Figure 5.6.6.10: This image covers the construction of the Chernobyl New Sarcophagus. ([File:1121-Txernobylgo zentral nuklearrerako arkofago berria.svg](#);))

Nuclear Tourism



For tour info: <https://www.chernobyl-tour.com/english/>

Resources

- <http://www.nucleartourist.com/events/chernobl.htm>
- [en.Wikipedia.org/wiki/Chernobyl_disaster](https://en.wikipedia.org/wiki/Chernobyl_disaster)
- Microsoft Word - WHO Report on Chernobyl Health Effects corrected July 5 06.doc
- Mitigating health consequences of Chernobyl (who.int)

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5.6.7: Fukushima Nuclear Disaster

Learning Objectives

- Explore the components of a Fukushima nuclear reactor.
- Compare and contrast Fukushima reactors to United States reactors.
- Determine how a partial meltdown occurred at Units 1 and 2 at the Fukushima Nuclear Power Plant.
- Explain whether a partial meltdown at Fukushima could have been avoided.
- Note the timeliness of evacuation efforts and compare them to the Chernobyl Nuclear accident.
- Compare radiation dosages and fatalities at Fukushima to those at Chernobyl and Three-Mile Island
- Review the safety of building nuclear reactors at geographical spots around the globe.

Nuclear Power Plants in Japan

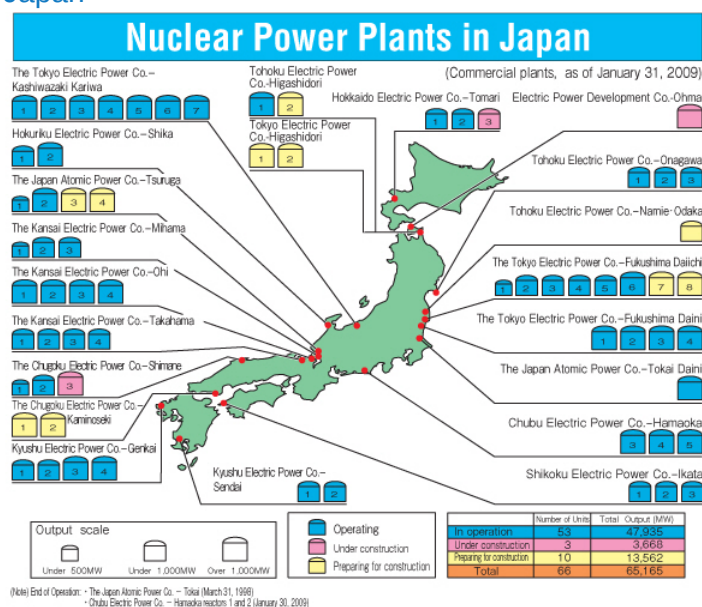


Figure 5.6.7.1: Copy and Paste Caption here. (Copyright;))

Background of Fukushima Dai-ichi nuclear facility

Fukushima Dai-ichi nuclear station was commissioned in March 1971. Six boiling water reactors (BWR) were designed to produce a net energy output of 2719MW. General Electric designed Units 1 (460 MW), 2(784 MW), and 6(1100 MW). Toshiba constructed Units 3 (784 MW) and 5 (784 MW). Lastly, Hitachi devised Unit 4 (784 MW). All six reactors were managed and maintained by the Tokyo Electric Power Electric Company (TEPCO).

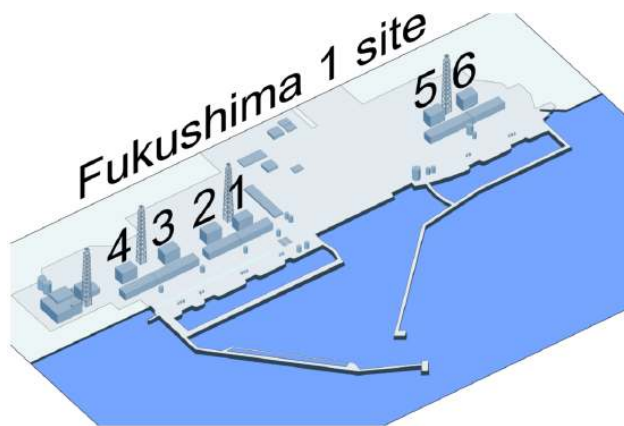


Figure 5.6.7.2: Fukushima I Nuclear Powerplant site close-up without text. (Copyright; <https://commons.wikimedia.org/w/index.php?title=File:Fukushima Daiichi Nuclear Power Plant 3D model.jpg&redlink=1>)

These six units all employed light water as a coolant and a moderator. All but Unit 3 utilized low-enriched uranium (LEU) for fissionable fuel. Unit 3 produced energy from mixed oxide fuel (M.O.X), composed of LEU oxides and fissionable Pu-239 oxides. Units 1-5 had a Mark 1 containment system while Unit 6 had upgraded to a Mark 2 containment model.

The Tsunami

On March 11, 2011, the Fukushima Daiichi Nuclear Power Plant in Japan was badly damaged by a 9.0-magnitude earthquake and resulting tsunami. At the time, Units 4, 5, and 6 were shut down for refueling and maintenance. Once the earthquake arrived at 2:46 p.m., the control rods of Units 1, 2, and 4 were immediately and completely submerged in their corresponding reactor vessels. Automatically, emergency generators came online to power electronics and coolant systems. However, the tsunami quickly flooded the emergency generators and cut power to the pumps that circulated coolant water through the reactors. High-temperature steam in the reactors reacted with zirconium alloy to produce hydrogen gas. The gas escaped into the containment building, and the mixture of hydrogen and air exploded. Radioactive material was released from the containment vessels as a result of the deliberate venting to reduce the hydrogen pressure, deliberate discharge of coolant water into the sea, and accidental or uncontrolled events.



Watch the video above (13:01 minutes) and answer the questions below. **This video is quite technical. Please focus on the questions and don't become overwhelmed with the comprehensive explanation of the reactor design.**

- 1) What two natural disasters occurred prior to the Fukushima nuclear disaster? How many people lost their lives due to these natural disasters?

- 2) How many reactors were producing energy before the explosion? How many reactors were unloading fuel? Lastly, how many reactors were shutdown?
- 3) What type of reactors were used at this site (RBMK, PWR, BWR, or AP1000)? Does the United States use this type of reactor?
- 4) Where was the spent fuel stored in the reactor? Note: The United States does not store spent fuel in this part of the reactor.
- 5) When the earthquake hit, did the control rods work?
- 6) If a reactor site loses power, a backup power supply is needed to keep the reactor core cool. Did Fukushima eventually lose back-up power? What did this do to reactors 1-3?
- 7) Several times, the NRC and IAEA told Fukushima that their sea walls were too low. How did this contribute to the loss of backup power?
- 8) Not having enough coolant, led to the production of hydrogen gas. What is the danger of having this gas inside the reactor core?
- 9) To reduce the pressure of the core, workers at Fukushima vented the radioactive gas into the atmosphere (similar to the Three-Mile Island situation). How else did this radiation affect the planet?
- 10) Workers at Fukushima used seawater to cool the core. What is the problem with using this substance?
- 11) Did the containment of the reactors hold?
- 12) What is a cold shutdown? Has this been achieved yet?
- 13) Did Japan force its citizens to man and then clean up Fukushima? Do you think the people involved understood the hazards of radiation (compare this to Chernobyl)?
- 14) At Fukushima, there were two deaths at the reactor site due to the natural disaster. There have been no reported deaths due to the meltdowns. How does this differ from Chernobyl?

Diagrams of the facility

Sea wall

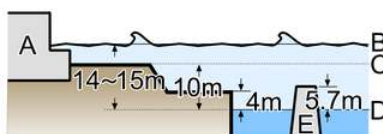


Figure 5.6.7.3: The height of the tsunami that struck the station approximately 50 minutes after the earthquake. A: Power station buildings B: Peak height of tsunami C: Ground level of site D: Average sea level E: Seawall to block waves. (Copyright, https://commons.wikimedia.org/w/index.php?title=File:Fukushima_tsunami.jpg&redlink=1)

Storage of spent fuel rods

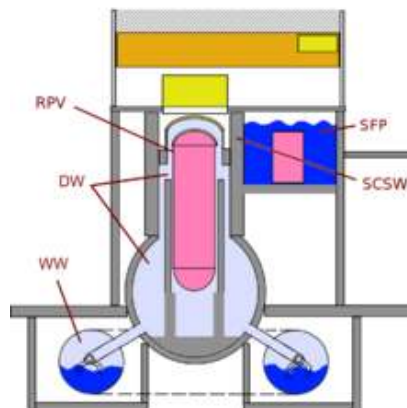


Figure 5.6.7.4: Cross-section of a typical BWR Mark I containment as used in units 1 to 5. **RPV**: reactor pressure vessel **DW**: dry well enclosing reactor pressure vessel. **WW**: wet well – torus-shaped all around the base enclosing steam suppression pool. Excess steam from the drywell enters the wet well water pool via downcomer pipes. **SFP**: spent fuel pool area **SCSW**: secondary concrete shield wall (Copyright; Sketched in Paint.net using 826 x 852 canvas inspired by Page 15 (pdf: 40), Containment Integrity Research at Sandia National Laboratories - An Overview, Manuscript completed: March 2006, Published: July 2006, id: NUREG/CR-6906 and SAND2006-2274P [1] - Prepared by Sandia National Laboratories for the U.S. Nuclear Regulatory Commission, Job Code Y6757. The two diagrams on the same page are labeled: *Figure 5 Typical BWR Mark I Concrete Containment with Steel Torus*.

Backup power sources

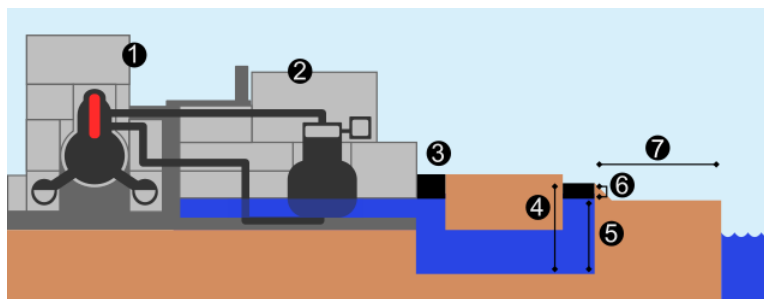


Figure 5.6.7.4: The diagram is labeled simply with numbers to be language neutral. Hope this helps! Key: 1 Reactor building; 2 Turbine generator and associated condenser; 3,4,5,6 and 7: various trenches and pipe tunnels (Copyright;<https://commons.wikimedia.org/wiki/User:Nesnad>)

Explosions and Meltdown

Immediately after the earthquake, the six external power supplies shut down. Loss of external power could result in overheating in the reactor cores. Of the six diesel backup generators, only one remained functional to provide cooling water. This generator cooled Units 5 and 6 until it lost power. The five other diesel generators were flooded when the tsunami hit. As a last resort, Fukushima accessed their 125-volt DC backup batteries. One of these energy sources was able to keep Unit 3 cool for 30 hours until it failed to produce any more power.

Unfortunately, hydrogen gas was produced in Units 1-4. Explosions of this gas occurred in Units 1,3 and 4. On March 14, hydrogen gas explosions blew the rooftops off of Units 1 and 3. There was no damage to the 20 cm thick steel containment structure.



Evacuation of the area surrounding Fukushima Dai-ichi

Around 3:40 p.m., backup sources for the reactors lost power. Evacuation of a 2 km radial zone began at 7:03 p.m. Two hours later, the radius of evacuation was increased to 3 km. The Japanese government announced the following morning that the evacuation zone would increase to 10 km. Later that day, the final evacuation zone of 20 km was publicized. An estimated 200,000 people were evacuated from this area.

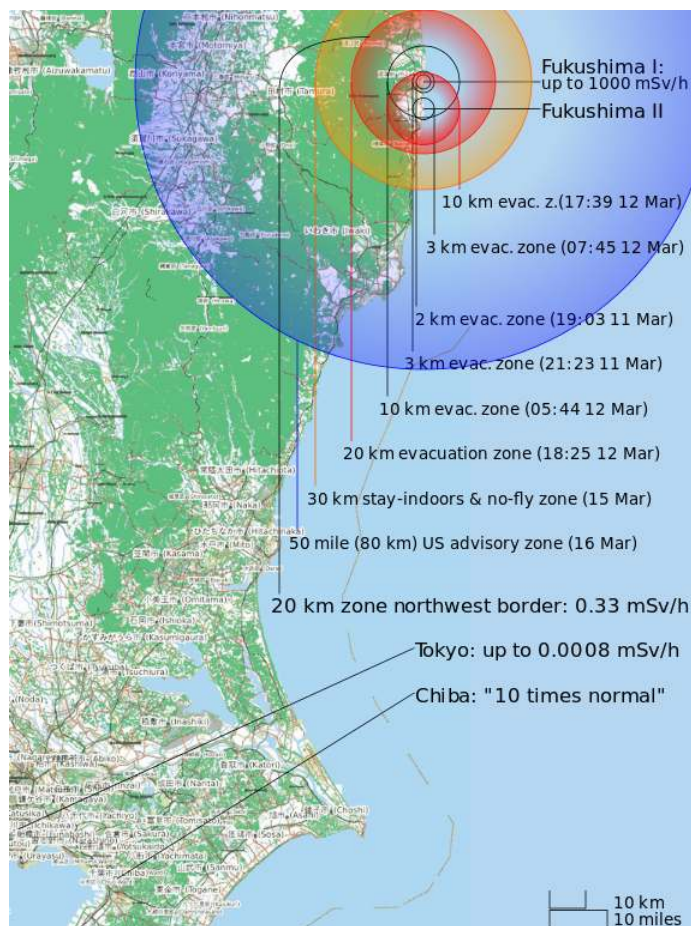


Figure 5.6.7.6: Fukushima I and II Nuclear Accidents Overview Map showing evacuation and other zone progression and selected radiation levels as of March 15. (Copyright;CMG Lee and OpenStreetMap contributors)

Radiation Levels

On March 16, radiation dosages inside the nuclear power station were 400 mSv per hour. Outside the facility, dosimeters measure 1.9 mSv per hour. The average TEPCO worker was exposed to 50-100 mSv per hour while on-site at the nuclear power plant.

As of 2021, there have been no reported deaths due to the partial meltdowns, explosions, or fallout from this nuclear disaster. Unlike the Chernobyl disaster, TEPCO workers (mostly composed of older males) volunteered in shifts to monitor the reactors during the accident. They were provided protective equipment and potassium iodide pills.

? Exercise 5.6.7.1

Compare the radiation dosages of a TEPCO employee (400 mSv/hour) and a person outside the Fukushima Dai-ichi facility (1.9 mSv/hour) to that of yearly background radiation (360 mrem) and Acute Radiation Syndrome (25 rem).

Part A. How much time could a TEPCO employee (400 mSv/hour) spend inside the facility until they exceed the yearly threshold for Acute Radiation Syndrome (25 rem)?

Part B. How much time could a person spend outside the Fukushima Dai-ichi facility (1.9 mSv/hour) until they exceed the yearly background radiation (360 mrem)?

Part C. How much time could a person spend outside the Fukushima Dai-ichi facility (1.9 mSv/hour) until they exceed the yearly threshold for Acute Radiation Syndrome (25 rem)?

Answer

For these calculations, recall the metric prefix milli from earlier this semester (1000 mSv = 1 Sv and 1000 mrem = 1 rem). You will also use the conversion 100 rem = 1 Sv.

Step 1. Convert mSv/hour to rem/hour

Step 2. Compare the annual exposure limit to the hourly limit to determine the number of hours

Part A.

$$\frac{400 \cancel{\text{mSv}}}{\text{hour}} \times \frac{1 \cancel{\text{Sv}}}{1000 \cancel{\text{mSv}}} \times \frac{100 \text{ rem}}{1 \cancel{\text{Sv}}} = \frac{40 \text{ rem}}{\text{hour}}$$

This tells us that TEPCO workers are exposed to 40 rem each hour they are in the facility. If the maximum annual dose for acute radiation syndrome is 25 rem per year, then these employees can spend no longer than 0.625 hours (or 37.5 minutes) in the facility before they are at risk for ARS.

$$\frac{25 \cancel{\text{rem}}}{40 \cancel{\text{rem}}} \times \frac{1 \cancel{\text{hour}}}{1 \cancel{\text{hour}}} \times \frac{60 \text{ min}}{1 \cancel{\text{hour}}} = 37.5 \text{ min}$$

Part B.

$$\frac{1.9 \cancel{\text{mSv}}}{\text{hour}} \times \frac{1 \cancel{\text{Sv}}}{1000 \cancel{\text{mSv}}} \times \frac{100 \cancel{\text{rem}}}{1 \cancel{\text{Sv}}} \times \frac{1000 \text{ mrem}}{1 \cancel{\text{rem}}} = \frac{190 \text{ mrem}}{\text{hour}}$$

This tells us that TEPCO workers are exposed to 190 mrem each hour they are on the outside of the facility. If the typical background radiation annual dose is 360 mrem per year, then these employees can spend no longer than 1.89 hours outside the facility before they exceed the background level.

$$\frac{360 \cancel{\text{mrem}}}{190 \cancel{\text{mrem}}} \times \frac{1 \text{ hour}}{1 \cancel{\text{hour}}} = 1.89 \text{ hours}$$

Part C.

$$\frac{1.9 \cancel{mSv}}{hour} \times \frac{1 \cancel{Sv}}{1000 \cancel{mSv}} \times \frac{100 \cancel{rem}}{1 \cancel{Sv}} = \frac{0.19 \cancel{rem}}{hour}$$

This tells us that a person outside the facility is exposed to 0.19 rem each hour. If the maximum annual dose for acute radiation syndrome is 25 rem per year, then a person should spend no longer than 131.6 hours (or around 5.5 days) at this location before they are at risk for ARS.

$$\frac{25 \cancel{rem}}{1 \cancel{rem}} \times \frac{1 \cancel{hour}}{0.19 \cancel{rem}} = 131.6 \cancel{hours}$$

Resources

- 1) <https://www.scientificamerican.com/a...r-power-plant/>
- 2) <https://world-nuclear.org/informatio...-accident.aspx>
- 3) <https://www.britannica.com/event/Fukushima-accident>

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5.6.8: Nuclear Bombs

Learning Objectives

- Describe the key features of an atomic bomb.
- Define nuclear fallout.

World War II brought about many advances in weaponry. The Nazis used the immense scientific ability of their newly conquered countries to attempt to create super weapons to keep the new territory. The weapons included new tanks like the feared Tiger Tank, or the highly advanced Luftwaffe Air Force. However, the one weapon the Nazi's really wanted to hold was the atomic bomb. The problem was that many of the great minds they once had left the country and came to the United States. In the end, Allied troops got to Hitler and his scientists before they could complete any atomic bombs. The U.S. also wanted to create an atomic bomb, and started The Manhattan to accomplish this goal.

Unlike the Nazi's efforts, The Manhattan Project in the United States was successful in part due to the vast resources that could be pulled together. The great minds worked under the University of Chicago football field to try and create fission (a reaction where a neutron collides with the nucleus of an atom causing it to split into smaller particles). A plant in Tennessee had the sole purpose of producing enough uranium for atomic bomb testing. Another testing center was established in New Mexico, becoming the epicenter of the operation. This is where the bombs were constructed and tested. The resources needed were immense. There had to be enough fuel for the bomb, and the materials required to make it. There had to be housing and food for all the scientists. Maybe most importantly, there had to be enough money to support the project. After years of hard work, an atomic bomb test took place on July 16, 1945. The bomb was held together in some parts by just tape, yet it was a huge success.

At the time of the atomic bomb test, the war in Europe had already ended, however, the war in the Pacific continued. This conflict was a lot different than the war in Europe. In Europe, Allied forces landed once and fought their way through to Berlin to defeat the Nazi's. However, in the Pacific, Japan controlled hundreds of Islands between the United States and Japan, the number of invasions required would have cost large amounts of American lives and resources. President Truman, first proposed the Potsdam Declaration, warning of "prompt and utter destruction" if Japan did not agree. After 11 days, when there was no reply from Japan, President Truman authorized the use of the atomic bomb on Japan.

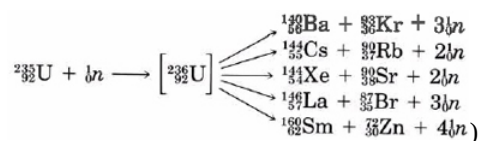
The first nuclear weapon used in combat, nicknamed "Little Boy," was dropped on Hiroshima on August 6, 1945, by the crew of the plane called the Enola Gay. The bomb destroyed the city, killing almost 100,000 people. The city of Hiroshima was essentially leveled. Following this attack, on August 9, 1945, another atomic bomb, nicknamed the "Fat Man," was dropped on Nagasaki, Japan. The same effects were felt in this city. The United States was ready to drop another atomic bomb in the third week of August, 3 more in September, and 3 more in October if necessary. However, on August 15, 1945, Japan surrendered to the Allies and World War II was officially over.



Figure 5.6.8.1 Hiroshima before the atomic bomb (top), Hiroshima after the atomic bomb (bottom).

The Workings of the Atomic Bomb

The Little Boy nuclear weapon contained enriched uranium-235. A crucial feature of this weapon was the fission of uranium. This type of fission reaction *produces more neutrons than it consumes*. As can be seen from Eqs. (1), for every neutron captured by a $^{235}_{92}\text{U}$ nucleus, between two and four neutrons are produced. Suppose now that we have a very large sample of the pure $^{235}_{92}\text{U}$ isotope and a stray neutron enters this sample. As soon as it hits a $^{235}_{92}\text{U}$ nucleus, fission will take place and about three neutrons will be produced. These in turn will cause the fission of three more $^{235}_{92}\text{U}$ nuclei, producing a total of nine neutrons. A third fission cycle will produce 27 neutrons; a fourth, 81; and so on. This process (which is called a *chain reaction*) escalates very rapidly. Within a few microseconds, a very large number of nuclei fission reactions occur releasing a tremendous amount of energy, resulting in an atomic explosion.



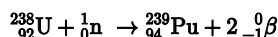
There are two reasons why a normal sample of uranium metal does not spontaneously explode in this way. First, natural uranium consists mainly of the isotope $^{238}_{92}\text{U}$, while the fissionable isotope $^{235}_{92}\text{U}$ comprises only 0.7 percent of the total. Most of the neutrons produced in a given fission process are captured by $^{238}_{92}\text{U}$ nuclei without any further production of neutrons. The escalation of the fission process thus becomes impossible. However, even a sample of pure $^{235}_{92}\text{U}$ will not always explode spontaneously. If it is sufficiently small, many of the neutrons will escape into the surroundings without causing further fission. The sample must exceed a **critical mass** before an explosion results. In an atomic bomb, several pieces of fissionable material, all of which are below the critical mass, are held sufficiently far apart to prevent the chain reaction from occurring. When these are suddenly brought together, an atomic explosion results immediately.

Isotopic Enrichment

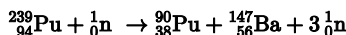
A great deal of the five years of the Manhattan Project was spent separating the 0.7% of $^{235}_{92}\text{U}$ from the more abundant $^{238}_{92}\text{U}$. Most of this work was performed at the at Kingston Demolition Range and Clinton Engineer Works in Oak Ridge, Tennessee. These two isotopes are nearly identical in physical properties, making it very challenging to isolate the uranium-235 isotope. Several methods for enriching uranium were developed and used in Oak Ridge. One method capitalized on the slight mass difference between the two isotopes. Gaseous UF_6 was heated and allowed to effuse through a porous screen. Each effusion resulted in a gas that was slightly richer in the lighter isotope. Repeating this process eventually produced a compound rich enough in $^{235}_{92}\text{U}$ for the purposes of bomb manufacturing.

Synthesis of Plutonium

In parallel with the work on uranium was an effort to produce plutonium-239, which was discovered at the Berkeley Radiation Laboratory in 1940. Only the first bomb dropped on Japan used uranium. The second bomb used the artificial element *plutonium*, produced by the neutron bombardment of ^{238}U :



Fission of Pu-239 occurs in much the same way as for U-235, giving a variety of products; for example,



Again this is a highly exothermic reaction yielding about the same energy per mole (20 000 GJ mol⁻¹) as ^{235}U .

This immense amount of energy released is the reason why plutonium and uranium were used. By comparison, 1 ton of TNT yields about 4 GJ mol⁻¹ of energy. This means that the atomic bombs had about 5,000 times more energy than 1 ton of TNT, explaining why the bomb was so effective and destructive.

Bomb Construction

At Los Alamos, New Mexico, J. Robert Oppenheimer, and his research group were given the task of assembling and testing the fission bomb. Both types of fuel that were manufactured (U-235 and Pu-239) would be placed separately into two different bomb designs. The smaller of the two bombs would contain U-235 and was code-named "Little Boy." The larger bomb, which housed Pu-239, was named "Fat Man." Of the two, only the Pu-239 bomb would be tested before it was dropped in wartime.

The basic design of the U-235 bomb is shown in the figure below. To prevent the spontaneous detonation of an atomic bomb, the fissile U-235 is kept in a subcritical configuration. It is then rapidly assembled into a supercritical mass using conventional explosives. Once the bomb has achieved this mass, any neutron introduced into it will be likely to initiate a chain reaction. The mechanism for "Little Boy" was a gun that fired one subcritical piece of U-235 into another to form a supercritical mass (Figure 5.6.8.2). The pieces had to be assembled faster than the spontaneous neutrons from either U-235 or cosmic radiation. A conventional explosive in an artillery barrel could fire the U-235 mass at speeds of a few millimeters per second, fast enough to prevent a fizzle caused by a spontaneous neutron setting off a premature chain reaction.

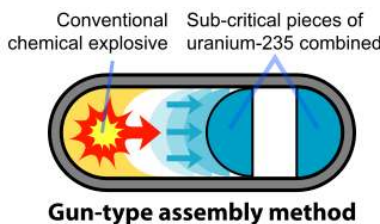


Figure 5.6.8.2: Little Boy: Gun-Type Device. (Public Domain).

Originally, the gun-type mechanism was planned for both the U-235 and Pu-239 weapons. However, a problem arose with the Pu-239 bomb that required a different assembly mechanism because of the small amount of Pu-240 that is produced with the Pu-239 in the reactor. Pu-240 emits large numbers of neutrons spontaneously: 1,030 neutrons per gram per second compared with 0.0004 neutrons per gram per second for U-235. Even at a concentration of 1% Pu-240 in the fissile Pu-239, the required mass of Pu emits 52,000 neutrons per second or one neutron every 20 microseconds. Thus, it is very probable that a neutron from Pu-240 will initiate a premature chain reaction during the critical last 100 microseconds of the critical mass in a gun-type assembly. This problem was discovered in mid-1944, well after the start of the construction of the massive Hanford plutonium production facilities.

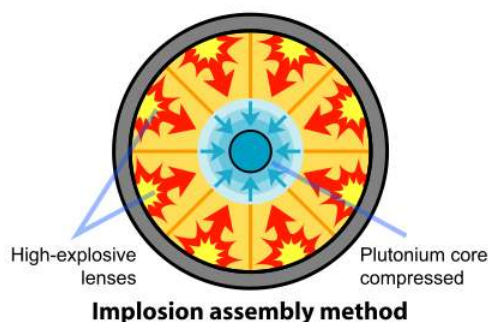


Figure 5.6.8.3: Fat Man: Implosion-type bomb. (Public Domain).

The removal of the Pu-240 was impractical. So the scientists and engineers looked for a faster method of assembling the plutonium. A mechanism based on implosion provided the solution to this problem. In this design, the fissile material is shaped into a single sphere with a mass slightly less than critical (Figure 5.6.8.3). Layers of carefully shaped high explosives surround the sphere. When the explosives are detonated, the force of the shock wave compresses the fissile material into a smaller volume, forming a supercritical mass. This method of assembly is much faster than the gun-type mechanism and thus eliminates the problems resulting from spontaneous neutron emission of Pu-240. The spherical mass resulted in a pumpkin-shaped weapon called "Fat Man".

✓ Example 5.6.8.1

Watch the from the 23-minute mark through the 32-minute mark of the third installment of the History Channel's Manhattan Project and answer the questions below.



1. What were the problems with the two different types of fuel that were to be used in the fission weapons?
2. What was the configuration or set-up of the Pu-239 bomb?
3. The combination of electromagnetic separation and gaseous diffusion enriched enough U-235 to produce how many bombs?
4. Why was the U-235 bomb never tested?
5. When was the construction of the U-235 completed?
6. How much did Truman know about the nuclear bombs when FDR was alive?
7. Who was selected to pilot the plane for the nuclear weapons?
8. What island would serve as a base to house the nuclear bombs before delivering them to their targets?

In July 1945, the United States had enough fissile material for one uranium and two plutonium weapons. The scientists and engineers felt confident that the gun-type assembly mechanism for Little Boy would function properly. Besides, they did not have the material for a test device. They were less confident about the implosion mechanism on the plutonium weapon and felt that a test was necessary. On July 16, 1945, the first nuclear device, known as "The Gadget", was placed on a 100-foot tower and successfully detonated in the Alamogordo Desert, 200 miles south of Los Alamos (Figure 5.6.8.4).

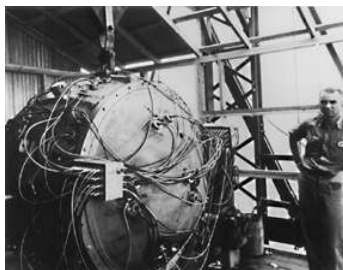


Figure 5.6.8.4: The Gadget at the Trinity Test Site, 1945. Courtesy of the U.S. Department of Defense

As the war with Japan continued and a costly allied invasion loomed as a real possibility, President Truman approved the use of nuclear weapons against selected Japanese targets. The U.S. Army Air Force received orders to use these weapons anytime after August 3, 1945. On August 6, "Little Boy" was dropped on Hiroshima (Figure 5.6.8.5).



Figure 5.6.8.5: Little Boy - The only U-235 weapons constructed during WWII. Courtesy of the U.S. Department of Defense

Little Boy was a uranium weapon containing 141.4 pounds of fissionable material containing 82.7% U-235. Only about two pounds fissioned, releasing an energy equivalent to 15-16,000 tons of TNT. The immediate effects of the blast killed an estimated 70,000 people, and by the end of 1945 an additional 20,000 to 70,000 deaths occurred, many due to lack of adequate medical resources.

Three days later, "Fat Man" destroyed a large part of Nagasaki (Figure 5.6.8.6). Fat Man contained 13.6 pounds of Pu-239, of which only 2 pounds underwent fission. The explosive yield was equivalent to about 22,000 tons of TNT. Fat man resulted in 35,000 immediate deaths. By the end of 1945, at least 70,000 people perished in this event.

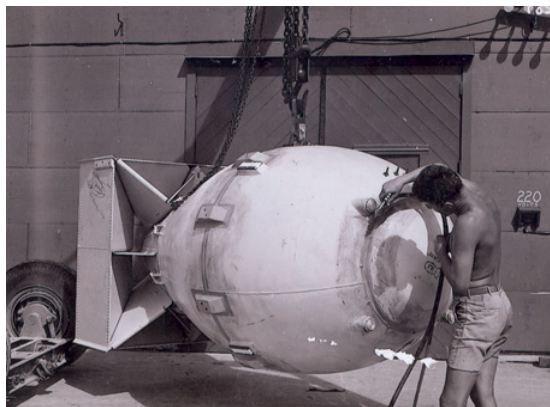


Figure 5.6.8.6: Fat Man – One of many Pu-239 nuclear weapons constructed during WWII. Courtesy of the U.S. Department of Defense

✓ Example 5.6.8.2

Start the video below at the 23:20 mark and stop it at the 29-minute setting. Then, answer the questions below:



1. When and where was the first nuclear bomb tested?
2. Was the bomb placed on the ground?
3. What were scientists afraid the first nuclear bomb would do?
4. What did Edward Teller (father of the fusion bomb) apply to his skin before the bomb was ignited? Why did he do this?
5. What did the bomb do to the tower and the surrounding sand?
6. In the photograph, where did General Groves focus his attention?

Nuclear Fallout

Nuclear fallout refers to the radioactive material that "falls out" of the atmosphere after a nuclear explosion (reactor or bomb). It consists of dust and radioactive particles that can contaminate an area with radioactivity and pose a huge health hazard to biological organisms. It can contaminate the animal food chain which can have drastic effects on the affected region. Weather has a huge impact on fallout, wind currents can spread radioactive fallout either over a large area, such as in the case of *Castle Bravo*, or not.

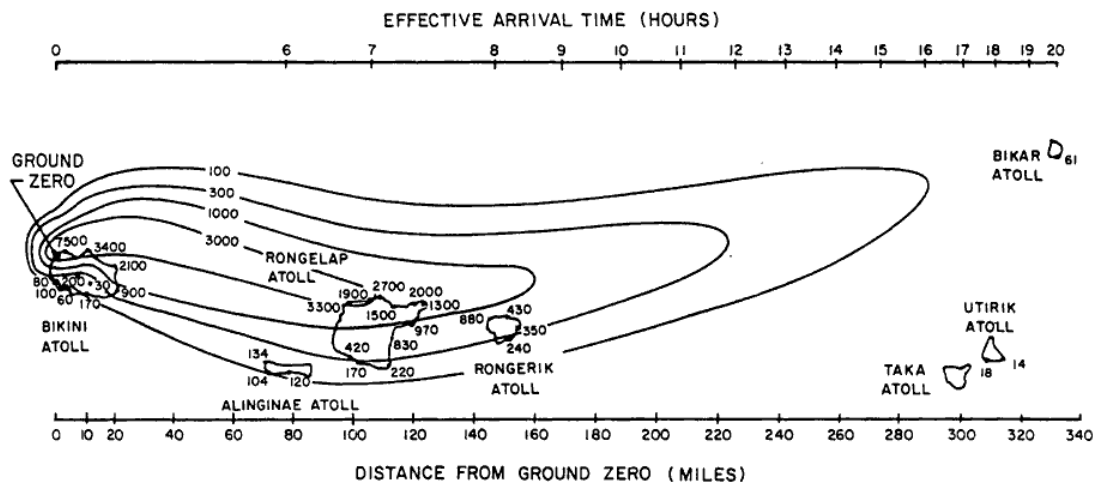


Figure 5.6.8.2: The 450 km (280 mi) fallout plume from 15 Mt shot Castle Bravo, 1954. (Public Domain; United States Department of Energy via Wikipedia)

In regards to nuclear reactions, fission produces more fallout products than fusion. Intense heat of the latter type of reactions reduces the amount and type of radioisotopes produced. Although more powerful, fusion is a "cleaner" type of nuclear reaction. By increasing the fission component of a fusion weapon, bombs can be designed to generate more fallout products. Of the different types of isotopes that fission produces, the most worrisome are **Sr-90**, **Cs-137**, and **I-131**. The first of these isotopes has a half-life of 28.8 years and decays by gamma and beta emission. Once ingested or inhaled, Sr-90 deposits in teeth, bone, and bone marrow. Sr-90 exposure can lead to bone cancer and/or leukemia. Cs-137 ($t_{1/2} = 30.17$ years) emits beta and gamma radiation. This ionizing

radiation deposits in soft tissues of the body. In addition, this particular radioisotope is water soluble and can easily contaminate food and water supplies. I-131 had a half-life over 8.1 days and undergoes gamma/beta decay. Once this fallout product enters the body, it attacks the thyroid gland. Extensive damage could affect heart rate, blood pressure, body temperature, and childhood growth. In addition, I-131 exposure could result in thyroid cancer.

✓ Nuclear Fallout

Watch the fallout video and then answer the questions below:



Summary

- In an atomic bomb, several pieces of fissionable material are held sufficiently far apart so no chain reaction can occur. When these components are brought together, the critical mass is reached and an atomic explosion results immediately.
- The two fission bombs assembled at Los Alamos, New Mexico, by J. Robert Oppenheimer, and his research group were manufactured with two types of fuel namely, U-235 and Pu-239.
 - The smaller of the two bombs contained U-235 and was code-named "Little Boy." This bomb was dropped in Hiroshima, Japan on Aug.6, 1945.
 - The larger bomb, which housed Pu-239, was named "Fat Man." This bomb was dropped in Nagasaki, Japan on Aug.9, 1945.
- Nuclear fallout refers to the radioactive material that "falls out" of the atmosphere after a nuclear explosion (reactor or bomb).

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- [Frank A. Settle](#) (Washington and Lee University)
- [Marisa Alviar-Agnew](#) (Sacramento City College)
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