

CHEMISTRY 115



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Chemistry 115

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Licensing

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1: The Properties of Oxygen Gas (Experiment)

Objectives

- To generate (and collect) oxygen gas via the decomposition of hydrogen peroxide.
- To investigate the properties of oxygen, particularly as an agent of combustion.

Oxygen is one of the most abundant elements on this planet. Our atmosphere is 21% free elemental oxygen. Oxygen is also extensively combined in compounds in the earth's crust, such as water (89%) and in mineral oxides. Even the human body is 65% oxygen by mass.

Free elemental oxygen occurs naturally as a gas in the form of diatomic molecules, $O_2(g)$. Oxygen exhibits many unique physical and chemical properties. For example, oxygen is a colorless and odorless gas, with a density greater than that of air, and a very low solubility in water. In fact, the latter two properties greatly facilitate the collection of oxygen in this lab. Among the unique chemical properties of oxygen are its ability to support respiration in plants and animals, and its ability to support combustion.

In this lab, oxygen will be generated as a product of the decomposition of hydrogen peroxide. A catalyst is used to speed up the rate of the decomposition reaction (without being consumed by the reaction), which would otherwise be too slow to use as a source of oxygen. The particular catalyst used here is active yeast. 

The reaction for generating oxygen gas from hydrogen peroxide is shown below:



The oxygen gas produced will be collected in bottles by a method known as the downward displacement of water (see figure below). Once collected, several tests will be performed in order to investigate the role of oxygen in a variety of combustion reactions.

A combustion reaction is commonly referred to as "burning". During a combustion reaction, oxygen reacts chemically with the substance being burned. Note that since our atmosphere is roughly 21% oxygen, many substances readily burn in air. Both oxygen and the substance being burned (the reactants) are consumed during the combustion reaction, while new substances (the products) and heat energy are generated. Since heat is produced, this is an exothermic reaction.

Combustion reactions can be described by the following chemical equation:



The actual products of a combustion reaction depend on what substance is burned and how much oxygen is present. In general, however, when a pure element burns in oxygen the product is called an oxide. An oxide is a compound containing both the element and oxygen chemically combined together.

Some examples of element combustion are shown below. Several such reactions will be performed using the oxygen gas collected in this lab.

Combustion of an Element can be described using the following chemical equation:



Two examples are:



Procedure

Materials and Equipment

Materials: hydrogen peroxide solution, active yeast, wooden splints, candle, sulfur, steel wool, magnesium ribbon, aluminum pellets, 6 M hydrochloric acid

Equipment: 250-mL Erlenmeyer flask, five wide-mouth bottles, four glass 'cover' plates, pneumatic trough, "stopper + thistle tube + tubing" apparatus, utility clamp, ring stand, deflagration spoon, crucible tongs, small beaker, medium test tube and plastic test tube rack.

Safety

Exercise caution and/or use gloves when using the hydrogen peroxide (H_2O_2) and the hydrochloric acid (HCl) as they can cause chemical burns and skin irritation. If either of these chemicals comes into contact with your skin, immediately rinse with water for a minimum of fifteen minutes and notify your instructor.

Also, do not look directly at the burning magnesium. In addition to being very bright, it emits harmful UV radiation that could cause damage to the retina of the eye.

Personal protective equipment (PPE) required: lab coat, safety goggles, closed-toe shoes, gloves

Part A: Generating and Collecting Oxygen Gas

1. Obtain the following equipment:

- 250-mL Erlenmeyer flask (locker)
- "two-hole stopper + thistle tube + glass tubing + rubber tubing" apparatus (stockroom)
- five wide mouth 'gas-collecting' bottles (under sink)
- four glass 'cover' plates (front desk)
- pneumatic trough (under sink) filled with water to 1/2 inch above the metal shelf

2. Fill four of the five wide-mouth bottles **to the brim** with water (the fifth will be used later). Then gently slide a glass plate over the mouth of each bottle. *Make sure that there are no air bubbles at the top of the glass plate.*

3. While holding the glass plate, gently invert a bottle and lower it into the water in the trough. Remove the glass plate when the mouth of the bottle is below the water level in the trough. Repeat this for all four bottles. Place the glass plates aside, as they will be used later.

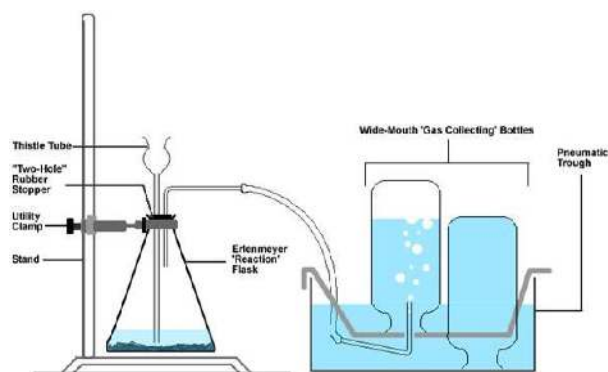
4. Place one gas-collecting bottle on the metal shelf. Make sure that the mouth of the bottle is below the water level.

5. Now focus on your reaction vessel, the Erlenmeyer flask. Add a pea-sized amount of active yeast (the catalyst) to the flask, followed by approximately 25-mL of distilled water.

6. Finally, assemble all your equipment together as demonstrated by your instructor, or as shown in the figure below. Make sure that

- the end of the thistle tube is completely covered with water at the bottom of the flask,
- the end of the glass tubing running from the Erlenmeyer flask is inserted under the opening in the bottom of the metal shelf into the gas-collecting bottle (which is full of water),

- the Erlenmeyer flask is stabilized with a utility clamp



- Obtain about 30-mL of hydrogen peroxide (H_2O_2) in your smallest beaker. Then carefully add about 10-mL of this H_2O_2 through the thistle tube. The generation of oxygen gas should begin immediately. If at any time the rate of the reaction in the Erlenmeyer flask appears to slow down, add another 10-mL portion of H_2O_2 .
- The oxygen produced will fill the inverted bottle by displacing the water in it. This is because oxygen does not dissolve in water, due to its low solubility. When the first bottle is completely filled with gas, place the second bottle on the metal rack in its place and allow it to fill in a like manner. Repeat this for the third and fourth bottles.
- As soon as each bottle is completely filled, remove it by placing a glass plate under the bottle's mouth while under water, then lifting the bottle and plate from the pneumatic trough. Place the bottle on the lab bench **mouth up** and **do not remove the glass plate**. Since oxygen is denser than air, it sinks to the bottom of the flask and will not readily leak out the top.
- Using masking tape, label each bottle of gas in the order they are collected: Bottle #1, Bottle #2, Bottle #3 and Bottle #4. Label the fifth unused bottle "Empty Bottle".
- Once all four bottles are filled with oxygen, do not add any more H_2O_2 to the Erlenmeyer flask. Set it aside and allow the reaction to go to completion. At the end of the lab, the chemicals remaining in the reaction flask and any unused H_2O_2 must be disposed of in the labeled container in the hood. In the meantime, proceed to Part B.

Part B: The Properties of Oxygen Gas

Test 1: Combustion of wood

Light a wooden splint, and then blow it out. While it is still glowing, place it in the empty wide-mouth bottle (air-filled only). Record your observations. Now re-light the same wooden splint, and again blow it out. Quickly insert the splint into Bottle #1 (oxygen-filled) while it is still glowing. Record your observations.

Test 2: Combustion of candle wax

Place a tealight candle on a glass plate and light it. Lower the empty bottle over the candle. Measure and record the number of seconds that the candle continues to burn. Then re-light the candle and lower Bottle #2 over it. Again, measure and record the number of seconds that the candle continues to burn.

Test 3: Combustion of sulfur

This test must be performed in the hood. Take the empty bottle and Bottle #3 to the hood your instructor directs you to. Place a small lump of sulfur in a deflagrating spoon. Light the Bunsen burner in the hood, and heat the sulfur in the spoon. The sulfur will first melt, then burn with an almost invisible blue flame. Insert the spoon with the burning sulfur in the empty bottle and record your observations. Then insert it in Bottle #3, and again record your observations. When finished, extinguish the burning sulfur in a beaker of water.

Test 4: Combustion of iron

Pour about 30-mL of tap water into Bottle #4 and replace the glass plate quickly. Take a loose (frayed out) 2 or 3 centimeter piece of steel wool and hold it in a Bunsen burner flame for a very brief instant with your crucible tongs (it will glow red). Then immediately lower the steel wool into Bottle #4. Record your observations. Repeat with the empty bottle and record your observations.

Test 5: Combustion of hydrogen

This test must be performed in air only. (Note: The hydrogen burned in this test must be first generated by a reaction between aluminum and hydrochloric acid.) To a medium test tube, add aluminum pellets, followed by about 3-mL of hydrochloric acid. Bubbles should begin to appear as hydrogen gas is produced. Place the test tube in the plastic test tube rack. After 30-60 seconds have elapsed, light a wooden splint. Do not blow it out. Hold the burning splint to the mouth of test tube (where the hydrogen gas is being evolved) and record your observations.

Test 6: Combustion of magnesium

This test is an instructor demonstration. It must be performed in air only. Hold a 1-inch piece of magnesium metal in a Bunsen burner flame with your crucible tongs until it ignites (in air). Record your observations, remembering not to look directly at the burning magnesium!

Lab Report: The Properties of Oxygen Gas

Part A: Generating and Collecting Oxygen Gas

- Write the equation for the reaction used to generate oxygen gas.
 - Word Equation:
 - Balanced Chemical Equation:
- What is the *name* of the catalyst used in this reaction?
 - What is the purpose of this catalyst? (There are two parts to this answer.)
- In addition to oxygen, what other substance is produced by this reaction? What happens to this substance?
- Two notable physical properties of oxygen are its low solubility in water and a density greater than air.
 - Which one of these properties allows the oxygen gas to be collected via the displacement of water? Explain.
 - Which one of these properties allows the oxygen gas collected to be stored in the bottles mouth up? Explain.
- What percentage of the air we breathe is oxygen gas?

Part B: The Properties of Oxygen Gas

Test	Observations
<i>Test 1</i>	<i>Observations</i>
Glowing splint in empty bottle	
Glowing splint in Bottle #1	
<i>Test 2</i>	<i>Observations</i>
Burning candle in empty bottle	
Burning candle in Bottle #2	
<i>Test 3</i>	<i>Observations</i>
Burning sulfur in empty bottle	
Burning sulfur in Bottle #3	
<i>Test 4</i>	<i>Observations</i>
Glowing steel in empty bottle	
Glowing steel in Bottle #4	
<i>Test 5</i>	<i>Observations</i>
Burning hydrogen in air	
<i>Test 6</i>	<i>Observations</i>
Burning magnesium in air	

Analysis of Combustion Results

- Why did you perform each test in two separate bottles (air-filled and oxygen-filled)?
- Are the combustion reactions of oxygen exothermic or endothermic? Support your answer with experimental evidence from the tests you performed.
- Consider your results for the first four tests you performed. In which bottles, air-filled or oxygen-filled, did the combustion reactions occur more vigorously? Why?
- Consider your Test 2 results. Although the candle burns for a longer period of time in one bottle, it eventually goes out in both the empty bottle and Bottle #2. Why does it go out?
- When an element burns in oxygen gas, the product is called an oxide.
 - The wood in the splint consists mostly of carbon. The combustion of carbon produces carbon dioxide, **CO₂**. Write the equation for the combustion of wood (carbon).
 - Balanced Chemical Equation:
 - The combustion of sulfur produces sulfur dioxide, **SO₂**. Write the equation for the combustion of sulfur.
 - Balanced Chemical Equation:
 - The combustion of hydrogen produces water, **H₂O**. Write the equation for the combustion of hydrogen.
 - Balanced Chemical Equation:
 - Steel wool consists mostly of iron. The combustion of iron produces iron(III) oxide, **Fe₂O₃**. Write the equation for the combustion of steel wool (iron).
 - Balanced Chemical Equation:



This page titled [1: The Properties of Oxygen Gas \(Experiment\)](#) is shared under a [not declared](#) license and was authored, remixed, and/or curated by [Santa Monica College](#).

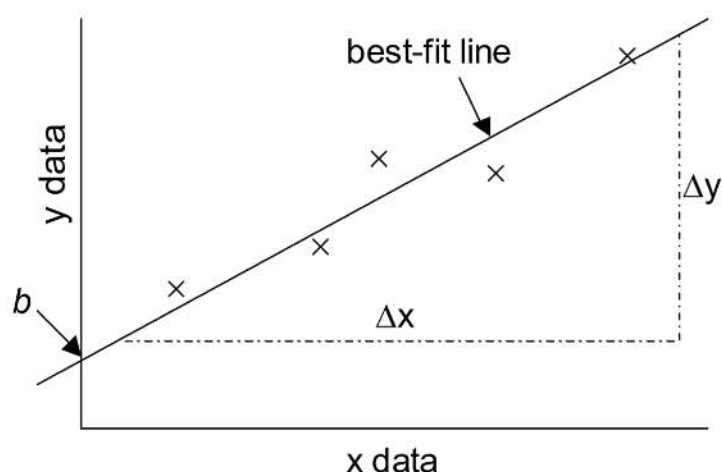
2: Using Excel for Graphical Analysis of Data (Experiment)

Objectives

- To learn to use Excel to explore a number of linear graphical relationships.

In several upcoming labs, a primary goal will be to determine the mathematical relationship between two variable physical parameters. Graphs are useful tools that can elucidate such relationships. First, plotting a graph provides a visual image of data and any trends therein. Second, via appropriate analysis, they provide us with the ability to predict the results of any changes to the system.

An important technique in graphical analysis is the transformation of experimental data to produce a straight line. If there is a direct, linear relationship between two variable parameters, the data may be fitted to the equation of line with the familiar form $y = mx + b$ through a technique known as [linear regression](#). Here m represents the slope of the line, and b represents the y-intercept, as shown in the figure below. This equation expresses the mathematical relationship between the two variables plotted, and allows for the prediction of unknown values within the parameters.



The equation for the best-fit line is

$$y = mx + b \quad (2.1)$$

where

$$b = \text{y-intercept} \quad (2.2)$$

$$m = \text{slope} \quad (2.3)$$

$$= \frac{\Delta y}{\Delta x} \quad (2.4)$$

$$= \frac{y_2 - y_1}{x_2 - x_1} \quad (2.5)$$

Computer spreadsheets are powerful tools for manipulating and graphing quantitative data. In this exercise, the spreadsheet program Microsoft Excel[®] will be used for this purpose. In particular, students will learn to use Excel in order to explore a number of linear graphical relationships. Please note that although Excel can fit curves to nonlinear data sets, this form of analysis is usually not as accurate as linear regression.

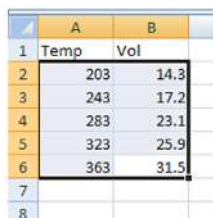
Procedure

Part 1: Simple Linear Plot

Scenario: A certain experiment is designed to measure the volume of 1 mole of helium gas at a variety of different temperatures, while keeping the gas pressure constant at 758 torr:

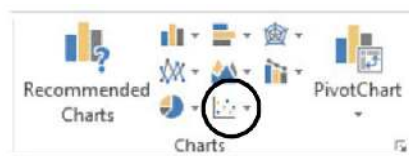
Temperature (K)	Volume of Helium (L)
203	14.3
243	17.2
283	23.1
323	25.9
363	31.5

1. Launch the program Microsoft Excel[®] (2016 version, found on all computers in all the computer centers on campus). Go to the Start button (at the bottom left on the screen), then click Programs, followed by Microsoft Excel[®].
2. Enter the above data into the first two columns in the spreadsheet.
 - Reserve the first row for column labels.
 - The x values must be entered to the left of the y values in the spreadsheet. Remember that the independent variable (the one that you, as the experimenter, have control of) goes on the x-axis while the dependent variable (the measured data) goes on the y-axis.
3. Highlight the set of data (not the column labels) that you wish to plot (Figure 1).



	A	B
1	Temp	Vol
2	203	14.3
3	243	17.2
4	283	23.1
5	323	25.9
6	363	31.5
7		
8		

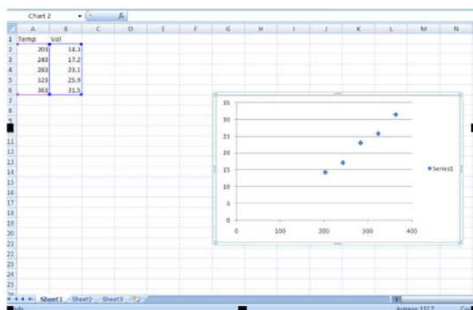
- Click on *Insert > Recommended Charts* followed by *Scatter* (Figure 2).



- Choose the scatter graph that shows data points only, with no connecting lines – the option labeled *Scatter with Only Markers* (Figure 3).

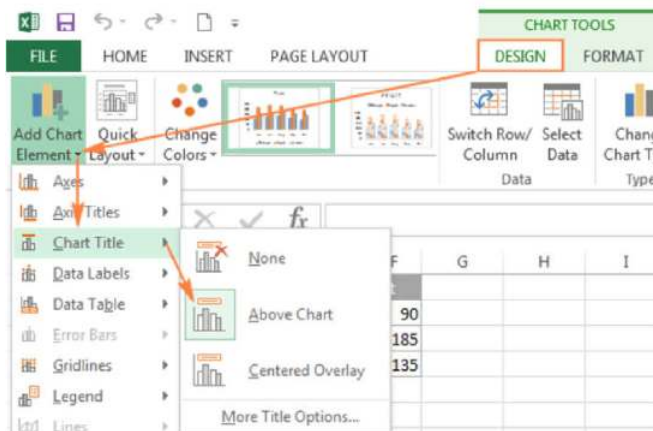


- You should now see a scatter plot on your Excel screen, which provides a preview of your graph (Figure 4).

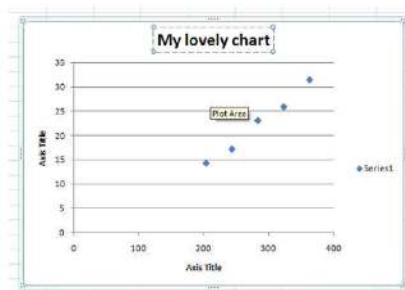


4. If all looks well, it is time to add titles and label the axes of your graph (Figure 5).

- First, click inside the chart.
 - Switch to the *Design* tab, and click *Add Chart Element* > *Chart Title* > *Above Chart*
 - The graph should be given a meaningful, explanatory title that starts out “Y versus X followed by a description of your system.
 - Click on *Axis Titles* (select *Primary Horizontal Axis Title* and *Primary Vertical Axis Title*) to add labels to the x- and y-axes.
- Note that it is important to label axes with both the measurement and the units used.



- To change the titles, click the text box for each title, highlight the text and type in your new title (Figure 6).



5. Your next step is to add a trendline to the plotted data points. A trendline represents the best possible linear fit to your data. To do this you first need to "activate" the graph. Do this by clicking on any one of the data points. When you do this, all the data points will appear highlighted.

- Click the *Chart Elements* button  next to the upper-right corner of the chart.
- Check the *Trendline* box.
- Click *More Options*. This will display the option shown in Figure 7.
- Notice that the *Linear* button is already selected. Now select the *Display Equation on Chart* box and the *Display R-squared value on Chart* box. Then click the *Close* button.

Format Trendline

TRENDLINE OPTIONS

TRENDLINE OPTIONS

☐ Exponential
☒ Linear
☐ Logarithmic
☐ Polynomial Order: 2
☐ Power
☐ Moving Average Period: 2

Trendline Name

☒ Automatic Linear (East Asia Sales 2010)
☐ Custom

Forecast

Forward: 0.0 periods
 Backward: 0.0 periods

☐ Set Intercept: 0.0
☐ Display Equation on chart
☐ Display R-squared value on chart

6. The equation that now appears on your graph is the equation of the fitted trendline. The R^2 value gives a measure of how well the data is fit by the equation. The closer the R^2 value is to 1, the better the fit. Generally, R^2 values of 0.95 or higher are considered good fits. Note that the program will always fit a trendline to the data no matter how good or awful the data is. You must judge the quality of the fit and the suitability of this type of fit to your data set.
7. Print out a full-sized copy of your prepared graph and attach it to your report. Then record the following information on your report:
 - the equation of the best-fit trendline to your data
 - the slope of the trendline
 - the y-intercept of the trendline
 - whether the fit of the line to the data is good or bad, and why.
8. By graphing the five measured values, a relationship is established between gas volume and temperature. The graph contains a visual representation of the relationship (the plot) and a mathematical expression of the relationship (the equation). It can now be used to make certain predictions.

For example, suppose the 1 mole sample of helium gas is cooled until its volume is measured to be 10.5 L. You are asked to determine the gas temperature. Note that the value 10.5 L falls outside the range of the plotted data. How can you find the temperature if it doesn't fall between the known points? There are two ways to do this.

Method (1): Extrapolate the trendline and estimate where the point on the line is.

- Click on the *Layout* tab along the top menu, then *Trendline > More Trendline Options*.
- In the section labeled *Forecast* enter a number in the box labeled *Backward*, since we want to extend the trendline the backward x direction. To decide what number to enter, look at your graph to see how far back along the x-axis you need to go in order to cover the area where volume = 10.5 L. After entering a number, click *Close*, and the line on your graph should now be extended in the backward direction.
- Now use your graph to estimate the x value by envisioning a straight line down from $y = 10.5$ L to the x-axis. Record this value on your report.

Method (2): Plug this value for volume into the equation of the trendline and solve for the unknown temperature. Do this and record your answer on your report. Note that this method is generally more precise than extrapolating and "eyeballing" from the graph.

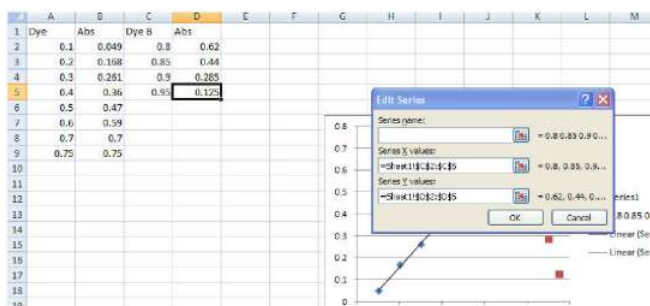
Part 2: Two Data Sets with Overlay

Scenario: In a certain experiment, a spectrophotometer is used to measure the light absorbance of several solutions containing different quantities of a red dye. The two sets of data collected are presented in the table below:

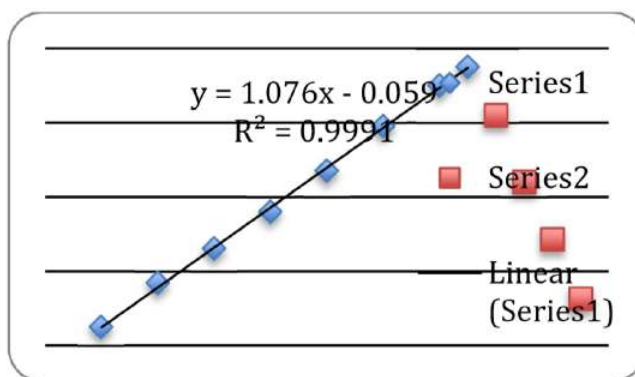
Data A		Data B	
Amount of Dye (mol)	Absorbance (unitless)	Amount of Dye (mol)	Absorbance (unitless)
0.100	0.049	0.800	0.620
0.200	0.168	0.850	0.440
0.300	0.261	0.900	0.285
0.400	0.360	0.950	0.125
0.500	0.470		
0.600	0.590		
0.700	0.700		
0.750	0.750		

You would like to see how these two sets of data relate to each other. To do this you will have to place both sets of data, as independent relationships, on the same graph. Note that this process only works when you have the same axis values and magnitudes.

1. Enter this new data on a fresh page (Sheet 2) in Excel. Be sure to label your data columns A and B. Again, remember to enter the x values to the left of the y values.
2. First, plot Data A only as an XY Scatter plot (the same way you did with the data in Part 1). Fit a trendline to this data using linear regression, and obtain the equation of this line.
3. Now you need to add Data B to this graph.
 - Activate the graph by clicking on one of the plotted data points.
 - Right-click the chart, and then choose *Select Data*. The *Select Data Source* box appears on the worksheet with the source data of the chart.
 - Click the *Add* tab and type "Data B" for the Series Name.
 - Click the little icon under *Series X values*, then highlight the x-axis values of Data B.
 - Press enter, then repeat this procedure for the *Series Y values*, highlighting the y-axis values of Data B. For each of these steps, you should see a display similar to what is shown in Figure 8. Note that slight differences may appear due to the version of Microsoft Excel[®] installed on your computer.



- Click OK twice to return to the main Excel window.
- At this point you should see the new data points (labeled as Series 2) as shown in Figure 9. You can now independently analyze this dataset by inserting a trendline as before.



4. Print out a full-sized copy of your prepared graph and attach it to your report. Then record the following information on your report:

- the equation of the best-fit trendline for Data A,
- the equation of the best-fit trendline for Data B,
- If these trendlines were extrapolated, they would intersect. Determine the values of x and y for the point of intersection using simultaneous equations.

Part 3: Statistical Analysis and Simple Scatter Plots

When many independent measurements are made for one variable, there is inevitably some scatter (noise) in the data. This is usually the result of random errors over which the experimenter has little control.

Scenario: Ten different students at two different colleges each measure the sulfate ion concentration in a sample of tap water:

College #1 dataset	35.9 ppm	43.2 ppm	33.5 ppm	35.1 ppm	32.8 ppm	37.6 ppm	31.9 ppm	36.6 ppm	35.0 ppm	32.0 ppm
College #2 dataset	45.1 ppm	34.2 ppm	36.8 ppm	31.0 ppm	40.7 ppm	29.6 ppm	35.4 ppm	32.5 ppm	43.5 ppm	38.8 ppm

Simple statistical analyses of these datasets might include calculations of the mean and median concentration, and the standard deviation. The mean ($\bar{x}$) is simply the average value, defined as the sum (Σ) of each of the measurements (x_i) in a data set divided by the number of measurements (N):

$$\bar{x} = \frac{\sum x_i}{N} \quad (2.6)$$

The median (M) is the midpoint value of a numerically ordered dataset, where half of the measurements are above the median and half are below. The median location of N measurements can be found using:

$$M = \frac{N + 1}{2} \quad (2.7)$$

When N is an odd number, the formula yields a integer that represents the value corresponding to the median location in an ordered distribution of measurements. For example, in the set of numbers (3 1 5 4 9 9 8) the median location is $(7 + 1) / 2$, or the 4th value. When applied to the numerically ordered set (1 3 4 5 8 9 9), the number 5 is the 4th value and is thus the median – three scores are above 5 and three are below 5. Note that if there were only 6 numbers in the set (1 3 4 5 8 9), the median location is $(6 + 1) / 2$, or the 3.5th value. In this case the median is half-way between the 3rd and 4th values in the ordered distribution, or 4.5.

Standard deviation (s) is a measure of the variation in a dataset, and is defined as the square root of the sum of squares divided by the number of measurements minus one:

$$s = \sqrt{\frac{\sum (x_i - \bar{x})^2}{N - 1}} \quad (2.8)$$

So to find s , subtract each measurement from the mean, square that result, add it to the results of each other difference squared, divide that sum by the number of measurements minus one, then take the square root of this result. The larger this value is, the greater the variation in the data, and the lower the precision in the measurements.

While the mean, median and standard deviation can be calculated by hand, it is often more convenient to use a calculator or computer to determine these values. Microsoft Excel® is particularly well suited for such statistical analyses, especially on large datasets.

1. Enter the data acquired by the students from **College #1** (only) into a single column of cells on a fresh page (Sheet 4) in Excel. Then in any empty cell (usually one close to the data cells), instruct the program to perform the required functions on the data. To compute the mean or average of the data entered in cells a1 through a10, for example, you must:

- click the mouse in an empty cell
- type "=average(a1:a10)"
- and press return

To obtain the median you would instead type "=median(a1:a10)". To obtain the standard deviation you would instead type "=stdev(a1:a10)".

2. Record on your report:

- The Excel calculated mean, median and standard deviation for the College #1 dataset.
- As an additional exercise, calculate the standard deviation of this dataset by hand, and compare it to the value obtained from the program.

Rejecting Outliers

Do all the measurements in the College #1 data set look equally good to you, or are there any values that do not seem to fit with the others? If so, are you allowed to reject these measurements?

Outliers are data points which lie far outside the range defined by the rest of the measurements and may skew your results to a great extent. If you determine that an outlier resulted from an obvious experimental error (e.g., you incorrectly read an instrument or prepared a solution), you may reject the point without hesitation. If, however, none of these errors is evident, you must use caution in making your decision to keep or reject a point. One rough criterion for rejecting a data point is if it lies **beyond two standard deviations from the mean or average**.

3. Using the above criteria, determine if there are any outliers in the College #1 dataset.

- Record these outlier measurements (if any) on your report.
- Then, excluding the outliers, re-calculate the mean, median and standard deviation of this data set (use Excel).

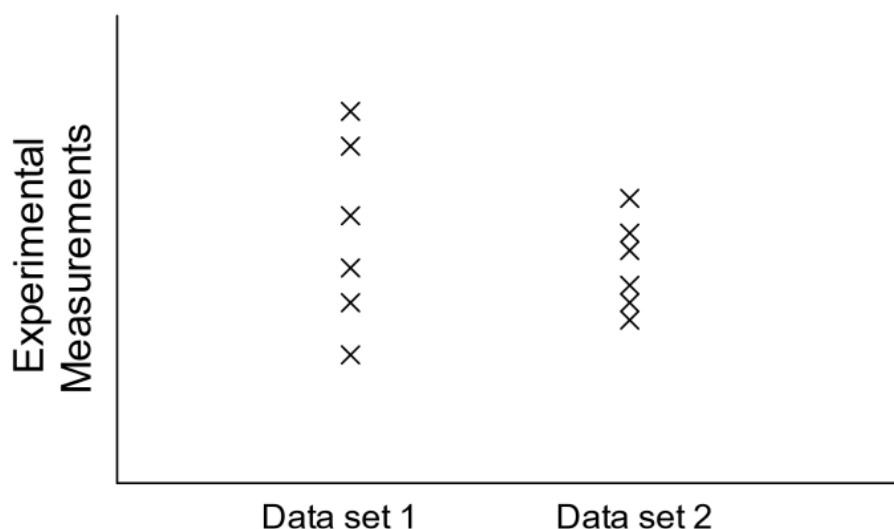
Rejecting data points cannot be done just because you want your results to look better. If you choose to reject an outlier for any reason, you must always include documentation in your lab report which clearly states:

- that you did reject a point
- which point you rejected
- why you rejected it

Failure to disclose this could constitute scientific fraud.

Graphing a Scatter Plot

Unlike the linear plots created so far, a scatter plot simply shows the variation in measurements of a single variable in a given dataset, i.e., it supplies a visual representation of the “noise” in the data. The data is plotted in a column, and there is no x-y dependence here (Figure 10). Note that datasets with a greater degree of scatter will have a higher standard deviation and consist of less precise measurements than datasets with a small degree of scatter.



To obtain such a plot using Excel, all the x values for each dataset must be identical. Thus, let the College #1 data be assigned $x = 1$, and let $x = 2$ for all the College #2 data:

Measurements by Students from College #1		Measurements by Students from College #2	
College 1	[SO ₄ ⁻²] (ppm)	College 2	[SO ₄ ⁻²] (ppm)
1	35.9	2	45.1
1	43.2	2	34.2
1	33.5	2	36.8
1	35.1	2	31.0
1	32.8	2	40.7
1	37.6	2	29.6
1	31.9	2	35.4
1	36.6	2	32.5
1	35.0	2	43.5
1	32.0	2	38.8

4. Enter the data as shown above into the first four columns of your spreadsheet.

- Plot the College #1 dataset as an XY Scatter Plot.
 - Now add the College #2 dataset to this graph applying the same steps you used to create your earlier graph in the section “Two Data Sets with Overlay” (Part 2).
 - Add appropriate axis labels and a title. You may also want to adjust the x-axis and y-axis scales to improve the final look of your graph.
5. Print out a full-sized copy of your prepared graph and attach it to your report. Then record the following information on your report:
- Which dataset (College #1 or College #2) show the least scatter? The greater standard deviation? The more precise measurements?

Lab Report: Using Excel for Graphical Analysis of Data

Name: _____ Lab Partner: None for this assignment

Date: _____ Lab Section: _____

Turn in the graphs you made for ALL three parts in this assignment

For each graph make sure the following components are in the printout:

1. Title for the graph
2. Labels for x and y axes (along with appropriate units when applicable)
3. Line equation and R^2 when appropriate.

Part 1: Simple Linear Plot

- Which set of data is plotted on the y-axis?
- Which set of data is plotted on the x-axis?
- Record the following information:
 - The equation of the fitted trendline
 - The value of the slope of this line
 - The value of the y-intercept of this line
- Is the fit of the trendline to your data good (circle one)? **Yes / No**. Explain why you think the line is a good fit to the data.
- Determine the temperature (in K) of the gas in the cold room when it has a measured volume of 10.5 L using
 - a) Extrapolation and “eyeballing”
 - b) The equation of the trendlineShow your calculations for b) below.

Part 2: Two Data Sets and Overlay

- Record the equations of the trendlines fitted to
 - Data set A:
 - Data set B:
- Perform a simultaneous equations calculation to determine the x and y values for the point of intersection between these lines. Show your work below.

Part 3: Statistical Analysis and Simple Scatter Plots

- For the College #1 data set, record the following values (determined using Excel):
 - the mean SO_4^{-2} concentration
 - the median SO_4^{-2} concentration
 - the standard deviation in the data set
- Calculate the standard deviation in the College #1 data set *by hand*. Show *all* your work below. Continue your work on an attached page if you require more space.
- Are there any outliers in the College #1 data set (circle one)? Yes / No
- If yes, which measurements are the outliers?
- Show the calculations you used to identify the outliers (or, if none, how you determined that there were none).
- Re-calculate the following values (using Excel) *excluding the outliers*:

- the mean SO_4^{-2} concentration
- the median SO_4^{-2} concentration
- the standard deviation in the data set
- Create a scatter plot showing both the College #1 and College #2 data. Attach a printout of your graph to this report. Be sure that your axes are properly labeled, and that your graph has an appropriate title.
- Examine your plotted data. Which data set:
 - has the larger standard deviation?
 - contains the more precise measurements?

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3: Lewis Structures and Molecular Shapes (Experiment)

OBJECTIVES

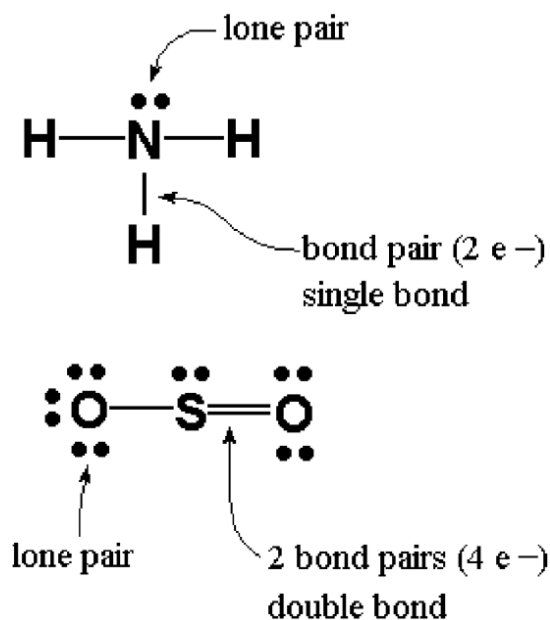
- To practice drawing Lewis Structures for various covalently bonded molecules and polyatomic ions.
- To use model kits to construct these molecules/ions in order to explore their structure and shapes.
- To practice predicting molecular shapes (using VSEPR theory) and molecular polarity.

Non-metal atoms bond covalently, resulting in the formation of either neutral molecules or polyatomic ions. A covalent bond is formed when non-metal atoms share their valence electrons, which they do in order to achieve filled valence orbitals like their nearest noble gas neighbor. This means that most bonded non-metal atoms will acquire a total of eight valence electrons via the sharing process – often referred to as the *octet rule*. A notable exception is hydrogen, which only needs to acquire two electrons to be like its nearest noble gas neighbor, helium.

Lewis Structures

A *Lewis Structure* is a representation of covalent molecules (or polyatomic ions) where all the valence electrons are shown distributed about the bonded atoms as either shared electron pairs (bond pairs) or unshared electron pairs (lone pairs). A shared pair of electrons is represented as a short line (a single bond). Sometimes atoms can share two pairs of electrons, represented by two short lines (a *double bond*). Atoms can even share three pairs of electrons, represented by three short lines (a *triple bond*). Pairs of dots are used to represent *lone pair electrons*.

Examples:



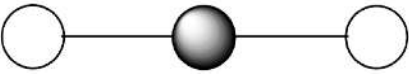
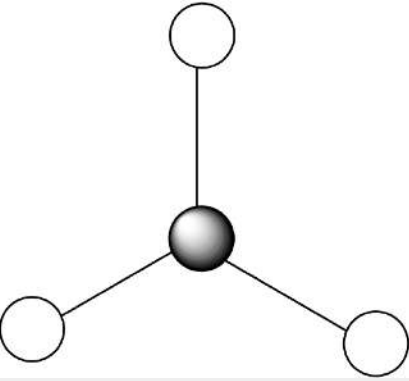
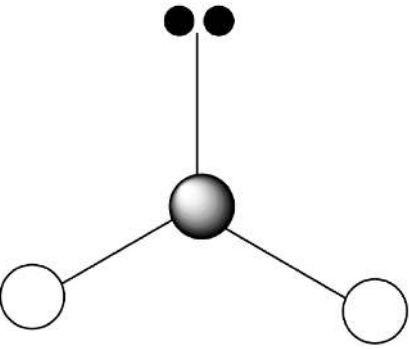
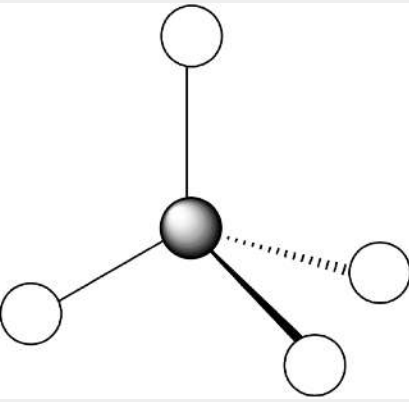
The rules for drawing Lewis structures can be found in the Procedure Section of this handout.

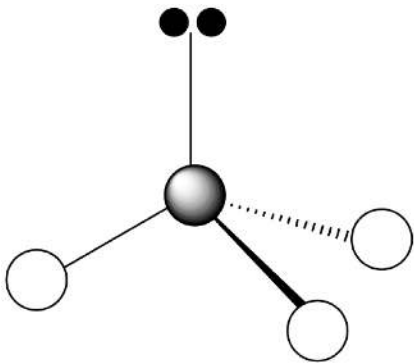
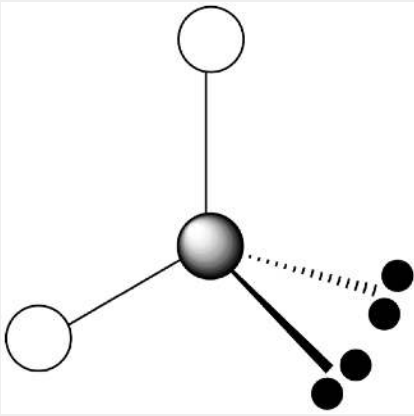
Molecular Shapes

The shape of a molecule depends on the distribution of atoms in space about the central atom, and their bond angles. Bond pair electrons and lone pair electrons repel one another, thus they will be arranged around a central atom as far apart as possible in order to minimize repulsions. This is known as:

Valence Shell Electron Pair Repulsion theory, or VSEPR theory.

The following VSEPR table supplies the names, sketches and descriptions of the most common types of molecular shapes that you will encounter. Note that several other molecular geometries do exist, however, they are beyond the scope of this course.

Shape Name	Sketch	Description
Linear		Only two outer atoms surround the central atom. There are no lone pairs on the central atom. Outer atoms are arranged opposite to each other. The bond angles are exactly 180° .
Trigonal Planar		Three outer atoms surround the central atom. There are no lone pairs on the central atom. The central and outer atoms all lie in the same plane (molecule is flat). Bond angles are exactly 120° .
Bent		Two outer atoms and one lone pair surround the central atom. Bond angles are slightly less than 120° .
Tetrahedral		Four outer atoms surround the central atom. There are no lone pairs on the central atom. The four outer atoms are evenly arranged in 3D around the central atom as if at the corners of a regular tetrahedron. The bond angles are exactly 109.5° .

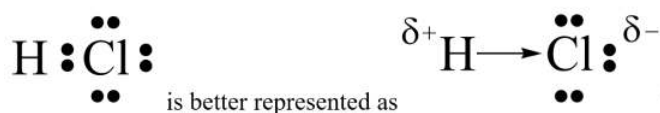
Shape Name	Sketch	Description
Trigonal Pyramidal		Three outer atoms and one lone pair surround the central atom. Here the central atom is located slightly above the three outer atoms, like a tripod. The bond angles are slightly less than 109.5° .
Bent		Two outer atoms and two lone pairs surround the central atom. Bond angles are slightly less than 109.5° .

Electronegativity and Bond Polarity

Some atoms in molecules have the ability to pull shared electrons closer to themselves than other atoms, an ability referred to as *electronegativity*. Electronegativity (χ) increases going across a period and decreases going down a group.

If two bonded atoms have different electronegativities, then the bond pair electrons will be *unequally shared*. The atom with the greater electronegativity will pull the bond electrons closer towards itself, causing it to obtain a very slight negative charge (δ^-). The atom with the lower electronegativity will have bond electrons pulled further away from it, causing it to obtain a very slight positive charge (δ^+). The result is a *polar covalent bond*. However, if two bonded atoms have the same electronegativity, then the bond pair electrons will be *equally shared*. The result is a *non-polar covalent bond*.

Example:



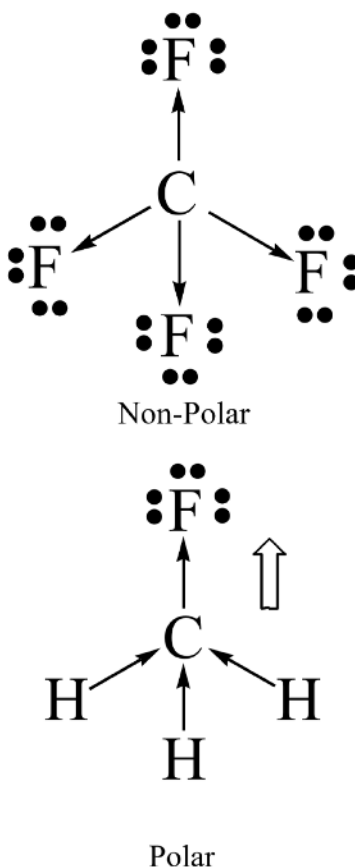
The bond electrons are pulled closer to Cl due to its greater electronegativity. Thus, HCl contains a polar covalent bond.

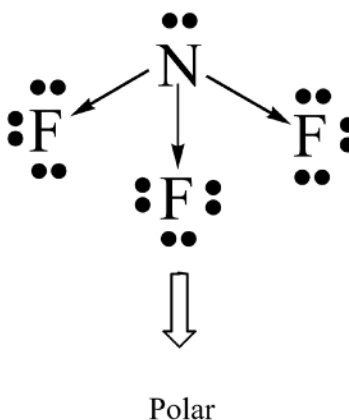
Molecular Polarity

Molecular polarity results when the *entire molecule* (not just a bond in the molecule) ends up with an unequal distribution of electrons. In general, a molecule will be *polar* if it contains polar bonds that are distributed in a *non-symmetrical* arrangement around the central atom. A polar molecule is said to have a net dipole moment. A non-symmetrical arrangement typically results when there are lone pairs on the central atom, or, when different outer atoms surround the central atom.

Not surprisingly, a molecule will be *non-polar* if it contains all non-polar bonds. It will also be non-polar if it contains polar bonds distributed in a *symmetrical* arrangement around the central atom. The symmetry causes the individual bond polarities to cancel out, resulting in a net non-polar molecule. A symmetrical arrangement typically results when there are no lone pairs on the central atom, and if all the outer atoms are identical.

Examples:





Procedure

Drawing Lewis Structures

Draw Lewis structures for each of the molecules and polyatomic ions given on your Report Form. Clearly show all bond pair electrons as lines and lone pair electrons as pairs of dots.

Rules for Drawing Lewis Structures

1. Total the number of valence electrons that each atom contributes to the molecule/polyatomic ion.
 - The quickest way is to find the group number for each atom.
 - For polyatomic anions, you must add electrons (equal to the negative charge) to the total number of valence electrons. For polyatomic cations, you must subtract electrons (equal to the positive charge) from the total number of valence electrons.
2. Draw a stick structure for the molecule.
 - Most molecules/polyatomic ions consist of one central atom bonded to 2, 3 or 4 other atoms.
 - The least electronegative atom is the central atom. Hydrogen is the only exception to this, as it forms only 1 bond. The central atom will usually need to form multiple bonds.
 - The other atoms are arranged around the central atom, and are attached to it via single bonds.
3. The octet rule must be obeyed for all elements except hydrogen (follows a “duet” rule).
 - Each bond in the stick structure contains two electrons, which must be subtracted from the total number of valence electrons.
 - Starting with the outside atoms, place the remaining electrons around each atom until it has a total of 8 electrons (except hydrogen – it only requires 2 electrons).
 - If there are not enough electrons available to obey the octet rule using single bonds, this indicates that double or triple bonds between two atoms are required in your structure. If short by two electrons, try a double bond, and if short by four electrons, try a triple bond or two double bonds.

Constructing Models, Determining Molecular Shapes and Molecular Polarity

1. Use your molecular model kit to construct a three-dimensional model of each of these molecules and polyatomic ions. Sketch a reasonably detailed picture of this model on your Report Form.

Rules for Constructing Molecules with the Model Kit

- Each colored ball in your kit corresponds to different atom or different group of atoms. Refer to the inside of the container lid for the correct color key.
 - Use the *short sticks* for *single bonds*.
 - Use two long flexible sticks for a *double bond* and three long flexible sticks for a *triple bond*.
2. Compare each model constructed to the molecular shapes described in the Theory section. Then identify and record its correct shape name and its bond angles. If the molecule has no definite central atom, then only record the bond angles.
 3. Evaluate the polarity of the bonds in each molecule as well as its overall symmetry in order to determine whether it is polar or non-polar.

Lab Report: Lewis Structures and Molecular Shapes

1. CH_4

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

2. CO_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

3. NH_4

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

4. H_2O

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

5. N_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

6. SO_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

7. O_2

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

8. O_3 - use yellow ball for central atom

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

9. CO

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

10. CO_3^{2-}

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

11. NO_3^- – use a blue ball for phosphorus when constructing this model

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

12. CF_2Cl_2 (CFC = **chlorofluorocarbon**)

Total # of Valence Electrons:

3-D Model Sketch (show dipole arrows):

Lewis Structure (show all resonance structures if applicable):

VSEPR shape name:

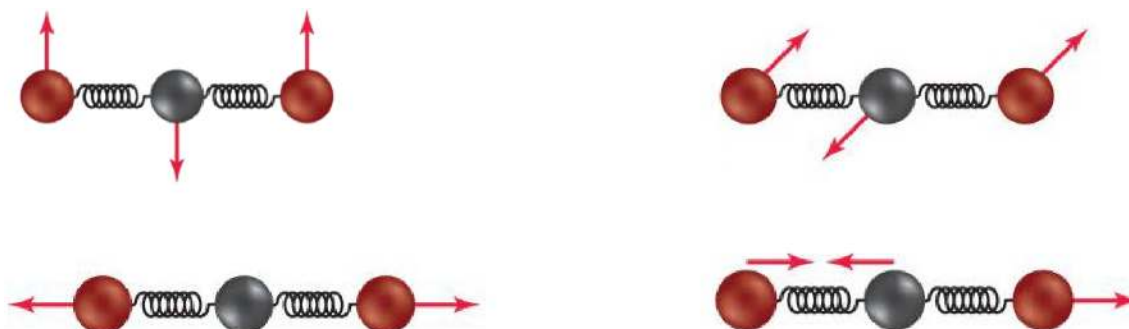
Is there a polar bond in this molecule?

Molecule Polarity: Polar/Non-Polar

Resonance: Yes/No

Questions

- When molecules and energy interact, there are different results. Atoms may dissociate, a molecule may rotate, or bonds may stretch and bend. Carbon dioxide has four possible vibrations. What is the effect of each vibration on the molecule: stretching, bending, or breaking?



- What is the definition of a greenhouse gas?
- Various atmospheric gases are listed below:

	Chemical Formula	VSEPR Shape (Sketch)	Greenhouse Gas (yes or no)
carbon dioxide			
methane			
nitrogen			
oxygen			
ozone			
sulfur dioxide			
water			

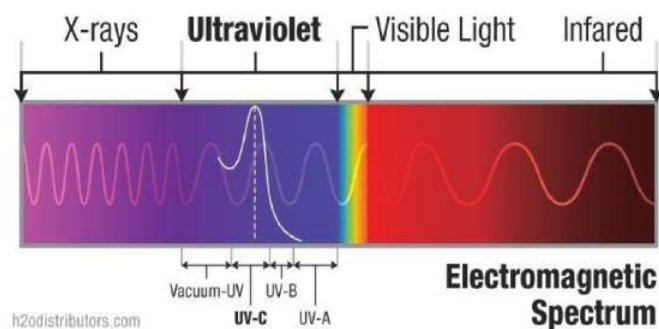
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4: Detection and Absorption of Ultraviolet Light (Experiment)

Objectives

- To understand what ultraviolet (UV) light is
- To understand the different types of UV light
- To test the abilities of materials to absorb UV light
- To understand relationship between UV light and sun protection factor (SPF)

Solar energy (sunlight) contains light we can see, and some we cannot. *Visible light* has wavelengths of 750 to 400 nm. *Ultraviolet (UV) light* has shorter wavelengths, cannot be seen, and has higher energy. Infrared (IR) radiation is the major source of heat for Earth. Though UV is a fraction of sunlight, it can be damaging to living organisms. All of these are forms of energy in the *electromagnetic spectrum*.



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Just as visible light components have names (red, orange, yellow, green, blue, indigo, violet), so do the types of UV light: UV-A, UV-B, UV-C and vacuum-UV. UV-A has lowest energy and is least damaging; UV-A is also called “black light.” UV-B and UV-C have higher energies and can cause break bonds of molecules, causing changes in DNA and thus skin cancers.

The majority of UV-B is absorbed by ozone in the stratosphere. Though UV-C is most damaging, it is totally absorbed by oxygen and ozone. In recent years, depletion of the ozone layer has allowed more UV light to reach us, resulting in more cases of skin cancers. Consequently, we have become aware of the need to protect ourselves from UV light.

What protects us from UV light? One strategy would be to avoid exposure to any type of sunlight. Since we cannot avoid sunlight while outdoors, we can physically or chemically block the sun. A wide variety of commercial sunscreens are available with *sun protection factors (SPF)* ranging from SPF 2 to SPF 100. These lotions contain organic molecules that absorb UV light. Some materials, such as glass and plastic also absorb UV light, while still allowing visible light through.

UV light type	Wavelength	Relative Energy	Comments
UV-A	320 – 400 nm	lowest energy	reaches Earth in greatest amount
UV-B	280 – 320 nm	higher energy than UV-A, but less than UV-C	most is absorbed by ozone
UV-C	200 – 280 nm	highest energy	absorbed by ozone and oxygen

Procedure

Materials and Equipment

UV-sensitive beads, sunscreen lotions with various SPF ratings, sunblock, sunglasses, clear and opaque plastic, glass plate, foil, cloth, small plastic bags, laboratory UV light, UV intensity meter card

Safety

Your eyes should not be directly exposed to the indoor lab UV light source. No materials used in this experiment should be ingested.

Personal protective equipment (PPE) required: lab coat, safety goggles

Part A: Detecting UV light

1. Place 3-5 UV-sensitive beads in two small plastic bags. These beads will turn color in the presence of UV light. The higher the intensity of UV light the stronger the color change.
2. Make two bags: one labeled “Control” and one labeled “Experiment”. Record the color of the beads when exposed to *indoor light* (lighting in the lab). The control bag will remain at your bench, and should be used for comparison to color changes in experiment.
3. Take the experiment bag outside, in an area of direct sunlight. Record the color of the beads after 10 seconds of exposure to outside *sunlight*.
4. Find a shaded area (e.g. under a tree or in a corridor). Record the color of the beads in *shade*.

Part B: Absorption of Outdoor UV light by various materials

1. Place on a tray: control and experiment bags from Part A and one of each of the materials from the front bench (clear plastic, opaque plastic, sunglasses, glass plate, foil and cloth). Take the tray outside to an area of direct sunlight.
2. Use each type of material to block the sunlight by holding directly above the experiment bag. Record the color of the beads when shielded by each sample. Note the colors of both control and experiment beads.

Part C: Absorption of UV Light by lotions

1. Place 3-5 UV-sensitive beads into a small plastic bag; make four bags. Label with the SPF numbers, and one sunblock. Note the color of the beads under indoor lighting.
2. Add one small drop/dab of each lotion onto each bag. Lightly coat/spread to cover one side of the bag. Allow lotion to dry. Bring the bags to a *laboratory UV light* setup. Leave bag for 10 seconds. Look into the front of the setup (but not directly at the light) to record the color of the beads under UV light.
3. Empty the beads into beaker at front bench. Discard coated plastic bags.

Part D: UV Light at SMC

1. Obtain a UV Intensity Meter card from the front bench. Go outside to the same area of sunlight you had been working in previously. Record the reading on your report sheet.

Lab Report: Detection and Absorption of Ultraviolet Light

Experimental Data and Observations

Part A: Detecting UV Light

Condition	Color of Control Beads	Color of Experiment Beads
indoor light		
sunlight		
shade		

Part B: Absorption of UV Light by various materials

Condition	Color of Control Beads	Color of Experiment Beads
translucent (clear) plastic		
opaque plastic		
cloth		
foil		
glass plate		
sunglasses		

Part C: Absorption of UV Light by lotions

Condition	Color of Beads Before UV	Color of Beads Under UV
sunblock		
SPF 50		
SPF 15		
SPF 5		

Part D: UV Light at SMC

- Record the intensity warning value, for the UV energy, using the UV Intensity Meter Card:
- Record the warning value and protection advisory (refer to Table 2.5 in Chemistry in Context) that corresponds to the value:

Questions

- What is the difference between UV-A, UV-B, and UV-C light?
- Compare the energies of UV light to IR and visible light. Explain why UV light is potentially more dangerous than IR or visible light.
- Opacity is the opposite of translucent (clear); for a completely opaque substance, you cannot see through it at all. Considering results from the plastics, glass, foil, and cloth pieces, did the opacity of each substance affect ability to block UV light? How might you use this knowledge to protect yourself when outdoors in sunlight?
- What is the purpose of an experimental control? Give one specific example of a control used in this experiment.
- Most sunscreen lotions claim to protect against UV-A and UV-B. Why don't they mention UV-C light?
- From your data in Part C, do higher SPF ratings provide more protection from UV light?

7. If the L.A. Times reports a UV Index of 6.5, and taking into account your skin pigmentation, how might this affect your plans for the day?

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5: Acid, Bases and pH (Experiment)

Objectives

- To set up and show how to use a pH indicator
- To determine the pH of common solutions
- To understand pH differences of acids and bases
- To learn to use a laboratory pH meter
- To understand relationship between pH and H^+ ion concentration

A pH value is a number, usually between 0 and 14, that represents the acidity or basicity of a solution. The “pH” is always written with a lowercase “p” and an uppercase “H”, which stands for “power of hydrogen.” pH values are related to hydrogen ion (H^+) concentrations.

The mathematical relationship between pH and H^+ is described by the equation

$$pH = -\log(H^+) \quad (5.1)$$

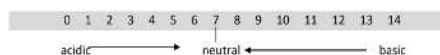
There is an *inverse* relationship between pH and H^+ ion concentration (in brackets, expressed in units of *molarity*, M). As the H^+ concentration decreases, the pH value increases, and vice versa. When the pH value is a whole number (e.g. pH 7), the number is equal to the negative exponent of the H^+ ion concentration.

$$pH \text{ value} = X \rightarrow [H^+] = 10^{-X} M \quad (5.2)$$

So for pH 7, the H^+ ion concentration is 10^{-7} M.

The pH values of everyday chemicals typically range from pH 0 to pH 14. Values between 0 and 7 indicate an acidic solution. Values between 7 and 14 indicate a basic solution. A pH of exactly 7 indicates that a solution is neutral, neither acidic or basic. Pure water is usually pH 7.

The *pH scale* is shown below.



The lower the pH value, the more *acidic* the solution; the higher the pH value, the more *basic* the solution. Basic solutions are also called *alkaline* solutions. It should be noted that the pH scale does extend beyond 0 and 14. Strong laboratory acids typically have pH values less than 0 (negative pH values) and strong laboratory bases typically have pH values greater than 14. Thus, they are considerably more dangerous.

The concept of pH is widely used in all areas of science including agriculture, biology, engineering and medicine. Many commercial products use pH as an advertisement tool, such as shampoo and water; more recently, food and drink of certain pH has been touted as more healthful.

A *pH indicator* is a substance that, when a small amount of it is added to a solution of unknown pH, will change its color. This is a way to determine pH of a solution visually. The indicator used in this lab will be obtained from a natural source, red cabbage. Cabbage indicator yields a particular color depending on the pH of the solution. pH indicators are a good way to easily and quickly show the approximate pH by color when compared to a standard. An everyday example where a pH indicator is used is for testing a water sample from a swimming pool.

While pH indicators are useful for qualitative purposes, when an exact quantitative value is needed, a *pH meter* is used. A laboratory pH meter typically has a special probe capped with a membrane that is sensitive to H^+ ion concentrations. The meter reading shows an exact pH value of the solution probed.



pH meters are used to measure pH values of water samples, such as determining acidity of rainwater samples. Rain water contains dissolved carbon dioxide that produces a weakly acidic solution. Rain naturally has a pH between 5 and 6. The pH of rain in parts of the U.S. is less than pH 5, which is harmful to aquatic life and human health. This is acid rain.

Living organisms are very sensitive to the effects of acids and bases in their environment. An excess of H^+ or OH^- can interfere with the functioning of biological molecules, especially proteins. Thus, in order to maintain homeostasis and survive, organisms must maintain a stable internal pH.

A *buffer* is a solution whose pH resists change on addition of small amounts of either an acid or a base. To be a good buffer, a solution should have a component that acts as a base (takes H^+ out of solution) and a component that acts as an acid (puts more H^+ into solution when there is an excess of OH^-).

The buffering capacity of a solution is tested by adding small amounts of acid (for example, HCl) and base (for example, $NaOH$) and checking the pH after each addition. If the pH changes only slightly, the solution is a good buffer. Eventually its buffering capacity will be exhausted, however, and the pH will change dramatically.

Procedure

Materials and Equipment

400-mL beaker, ring stand, wire gauze, Bunsen burner, large test tubes, dropper pipet, stirring rod, wash bottle with distilled water, laboratory pH meter, 0.1 M acetic acid, 0.1 M $NaC_2H_3O_2$, 0.1 M acetic acid ($HC_2H_3O_2$), 0.1 M hydrochloric acid (HCl), 0.1 M sodium hydroxide ($NaOH$), pH paper

safety

Exercise appropriate caution when using the [Bunsen burner](#).

Personal Protective Equipment (PPE) required: safety goggles, lab coat, closed-toe shoes Materials and Equipment

Experimental Procedure

Part A: Preparing pH indicator and pH standards - INSTRUCTOR DEMO

- Tear a few leaves of red cabbage into small pieces and place the leaves into a 500-mL beaker. Add about 500-mL of distilled water to this beaker. Make sure that all of the leaf pieces are completely submerged.
- Gently boil the mixture on heating plate until it appears dark purple in color (5-10 min). Turn off the heat and allow to cool (5 min).
- Add cabbage indicator solution to pH standard solutions, labeled 1-13. Students will record the colors of the pH standards.
- Each group will bring a small beaker to the front and take ~50 mL of the cabbage indicator back to their bench.

Part B: Qualitative Analysis for pH Values of Everyday Chemicals

1. Obtain 10 large test tubes (clean, but may be wet). Label each test tube with the solutions to be tested.
2. Pour about 3-mL of each solution into the appropriately labeled test tube.
3. Using a dropper pipet, add an equal volume of cabbage indicator solution. If necessary, stir to mix with a clean stirring rod (rinse with distilled water between uses).
4. Record the resulting color of the sample after mixed with the cabbage indicator. Compare this color with pH standards at the front of the laboratory to determine the pH of the sample. The color may be between the pH standard colors (e.g. green-blue instead of green or blue alone). For these, record the pH to 0.5 values (e.g. pH = 9.5 instead of 9 or 10).

Note

Do not discard the contents in these test tubes as they will be used in the next section.

Part C: Quantitative Analysis for pH Values of Everyday Chemicals

1. Plug the probe into one port on the side of the pH meter. Plug the AC adapter into the other side of the pH meter; plug the adapter into an electrical outlet. You should see a pH value reading.
2. Prepare the probe to make pH measurements: remove from the storage bottle and thoroughly rinse the lower section of the probe with distilled water/wash bottle.

Note: Do not completely submerge the probe. The handle is not waterproof.

3. Use the same ten test tubes containing samples from Part B. Or, complete steps 1 and 2 of Part B to obtain ten samples for analysis. Insert the pH probe directly into each test tube. **SPECIAL CARE IS NEEDED WHEN INSERTING THE PROBES INTO THE TEST TUBES. The probes must NOT touch the glass rim of the test tubes or the pH blub can easily be broken and the probe destroyed.**
4. Record the pH value (to 0.01 pH) shown on the pH meter screen.
5. After each pH measurement, the probe must be thoroughly rinsed with distilled water.
6. When you are finished making measurements, rinse the probe with distilled water. Slide the cap onto the probe, and then screw the cap onto the storage bottle so the tip of the probe is immersed in the storage solution.

Part D: Effect of Buffers on pH

1. Obtain 4 large test tubes. Label the test tubes A, B, C, and D.
2. Add 10-mL of distilled water to tubes A and C.
3. Add a 5-mL quantity of both 0.1 M $\text{H}_2\text{C}_2\text{H}_3\text{O}_2$ (acetic acid) and 0.1 M $\text{NaC}_2\text{H}_3\text{O}_2$ (sodium acetate) to tubes B and D. This mixture of acetic acid and sodium acetate is a buffer solution. Stir to mix completely.
4. Using pH paper, determine the pH of the contents of each test tube (A-D). Use the stirring rod to dab a small drop of the solution to be tested onto a piece of pH paper. Then compare the color obtained to the pH scale on the instructor's desk to determine the pH value. Record these pH values to 0.1
5. Add 5 drops of 0.1 M HCl (hydrochloric acid) to test tubes A and B. Record the pH using pH paper.
6. Add 5 drops of 0.1 M NaOH (sodium hydroxide) to test tubes C and D. Record the pH using pH paper.

Lab Report: Acids, Bases and pH

Part A: Color of Red Cabbage Indicator with pH standards

pH standard	Color with Cabbage Indicator
1	
2	
3	
4	
5	
6	
7	
8	
9	
10	
11	
12	

Parts B and C: pH of Everyday Chemicals

Chemical	Color with Indicator	Qualitative pH (to 0.5)	Acidic, Basic, or Neutral?	Quantitative pH (to 0.01)
soda				
shampoo				
ammonia cleaner				
bleach				
laundry detergent				
lemon juice				
vinegar				
bottled water				
0.1 M HCl				
0.1 M NaOH				

Part D: Effect of Buffers on pH

T u b e	Contents	Initial pH	Chemical Added	New pH	pH change
A	Water		HCl		
B	Buffer solution (acetic acid and sodium acetate)		HCl		
C	Water		NaOH		
D	Buffer solution (acetic acid and sodium acetate)		NaOH		

Questions

- What is an acidic solution? What is a basic solution?
- What is an alkaline solution?
- What is a pH indicator? What are common uses of pH indicators?
- Write the mathematical equation that relates pH value and H^+ ion concentration:

- Circle correct choice:

Acids have (high OR low) pH, and (high OR low) H^+ ion concentration.

Bases have (high OR low) pH, and (high OR low) H^+ ion concentration.

- When the H^+ ion concentration is expressed in brackets $[\text{H}^+]$, what are the units of the for H^+ ion concentration?
- Does a solution with pH 10 have equal, less or more H^+ ions than of a solution with a pH 6? Calculate the $[\text{H}^+]$ for both solutions, include units in your answer:

- For pH 10, $[\text{H}^+]$ = _____
 - For pH 6, $[\text{H}^+]$ = _____
7. The two methods of determining pH values (pH indicator versus pH meter) should show similar pH values for those solutions. What was different?
8. Explain why rain is naturally acidic, but not all rain is classified as “acid rain.”
9. Here are examples of what an individual might do to reduce acid rain. For each, explain the connection to production of acid rain.
- avoid running a washing machine with a small load
 - add additional insulation on a home hot water heater
 - walking instead of driving to work

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6: Synthetic Polymers and Plastics (Experiment)

Objectives

The objectives of this laboratory are to:

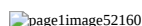
- Understand properties and uses of synthetic thermoplastics
- Compare the physical properties of “Big Six” plastics
- Identify everyday plastics by their physical properties
- Draw basic structures of polymers when given monomer structure

The word “polymer” means “many units”. A polymer can be made up of many repeating units, which are small monomer molecules that have been covalently bonded. Figure 1 (from Chemistry in Context) shows a single monomer, and a polymer made of identical monomers linked together. A polymer can contain hundreds of monomers, totaling thousands of atoms.

Examples of naturally-occurring polymers are silk, cotton, wood, cotton, starch, natural rubber, skin, hair and DNA. In the early 1900s, chemists began to replicate natural polymers, and create synthetic polymers, beginning with nylon which mimics silk in its strength and flexibility.

Plastic is a type of synthetic polymer. Currently, more than 60,000 plastics are manufactured for industrial and commercial purposes. Roughly 75% of the plastics used in this country can be categorized as one of six types, or “The Big Six”. These polymers are listed in the table below.

No.	Name	Abbreviation 	Uses
1	polyethylene terephthalate	PET	clear bottles and containers, fleece, carpet
2	high-density polyethylene	HDPE 	opaque bottles and containers, buckets, crates
3	polyvinyl chloride	PVC	rigid form: pipes & credit cards; soft form: tubing
4	low-density polyethylene	LDPE	bags, films, sheets, bubble wrap, toys
5	polypropylene	PP	bottle caps, yogurt containers, furniture
6	polystyrene	PS	expandable form: styrofoam; crystal form: CD cases



These six polymers are thermoplastic: they can be melted and reshaped, or recycled. The numbers are used to ease identification of the plastics, so that they can be separated for recycling. The symbol used on plastic packaging materials is three arrows in a triangle shape, with the number of plastic in the middle. This symbol makes recycling easier by making the identification of plastics easier.

The Big Six thermoplastics have these general attributes:

- recyclable
- insoluble in water
- resistant to most chemicals
- lightweight yet strong
- can be shaped
- can be colored with pigments
- usually made from petroleum

- used to make items that have no alternatives from other materials

The most common of the Big Six plastics is high-density polyethylene (HDPE). It is composed of repeating units of the monomer ethylene



Monomers are linked together in an addition polymerization reaction. Each new monomer adds to one end with a covalent bond; the total number of monomers in the polymer is represented by the subscript, n . The resulting polymer is a chain of monomers linked together. The figure at right depicts one part of the polymer chain. How many monomers are present?

In this experiment, you will be qualitatively analyzing plastic polymers for physical characteristics of opacity, flexibility, durability, and breakability. You'll also analyze the density of each plastic by checking whether pellet samples float or sink in three liquids of different densities.

In this lab, you will make a polymer bouncy ball using a chemical reaction between borax and glue. Glue contains the polymer polyvinyl acetate, which cross-links to itself when reacted with borax (diagram below). After cross-linking, the glue is no longer fluid, but more solid. Adding cornstarch helps to bind the molecules together so that they hold their shape.

Procedure

Safety

No materials used in this experiment should be ingested.

Personal Protective Equipment (PPE) required: safety goggles, lab coat, closed-toe shoes Materials and Equipment

samples of Big Six plastics (marked with recycling symbols), pellets of Big Six plastics, 1:1 95% ethanol/water solution, distilled water, 10% NaCl solution, 3 small test tubes, glass stirring rod, wash bottle with distilled water, 3 small beakers, 3 plastic spoons, paper cup, Elmer's glue, borax, cornstarch, ruler

Part A: Physical Characteristics of Plastic Polymers

Samples of different plastics are available at the front bench. Identify the samples by looking for the number/recycling symbol. Use these samples to analyze physical characteristics of each type of plastic: recyclability, opacity, durability/hardness, and flexibility.

Part B: Density Tests of Big Six Plastics

1. Three solutions of different densities will be used:
 - Solution A = 1:1 95% ethanol/water, density = 0.94 g/cm³
 - Solution B = distilled water, density = 1.0 g/cm³
 - Solution C = 10% NaCl, density = 1.08 g/cm³
2. Obtain and label three small test tubes: Solution A, B and C. Add about 3 mL (two full dropper squirts) to each test tube.
3. Place one piece of each plastic into each of the three test tubes. Push each piece under the liquid surface with a glass stirring rod. Surface tensions will cause all of the plastic to float until each is "wetted" and submerged using the rod.

Record whether the sample sinks rapidly, sinks slowly, floats on the surface, or floats below the surface (but doesn't sink to bottom).

If the sample floats, it has a density lower than that of the solution. This may be relative to another sample that floats. If the sample sinks, it has a density greater than that of the liquid. The sample may also sink rapidly or slowly relative to other samples.
4. Test each of the six plastic types accordingly, to complete the table in the lab report.

Part C: Polymer Bouncy Balls

1. Obtain front bench: paper cup with approximately 100 mL of Elmer's glue (sample cup will be marked to 100 mL), ruler and 3 plastic spoons. Obtain a wash bottle with distilled water.
2. Obtain from your locker: 3 small beakers, stirring rod, small graduated cylinder

Making Polymer Bouncy Ball #1:

1. In a glass beaker, add:
 - 3 level spoonfuls of glue
 - 5 mL distilled water
 - 1 level spoonful of borax powder
2. DO NOT STIR. Allow the ingredients to interact for 10-15 seconds. Then use stirring rod to mix. Once the mixture becomes impossible to stir, take it out of the beaker and mold the ball with your hands. The ball will start out sticky and messy, but will solidify as it is kneaded.
3. Record physical observations about the ball in the table: is the ball stretchy? goopy? slimy?
4. Use the ruler and hold the ball at a height of 30 cm (=12 in) above the bench. Drop the ball and record how high it bounces.

Making Polymer Bouncy Ball #2:

1. In a glass beaker, add:
 - 3 level spoonfuls of glue
 - 5 mL distilled water
 - 1 level spoonful of cornstarch
 - 1 level spoonful of borax
2. Repeat steps 2-4 from previous.

Making Polymer Bouncy Ball #3:

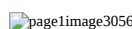
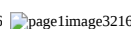
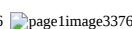
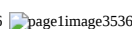
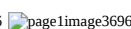
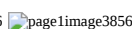
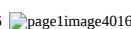
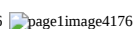
1. In a glass beaker, add:
 - 3 level spoonfuls of glue
 - 1 level spoonful of cornstarch
 - 1 level spoonful of borax
2. Repeat steps 2-4 from previous.
3. You and your lab partner may take home these bouncy balls, as all of the materials are non- toxic. However, please remember that they're not edible!

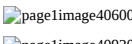
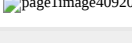
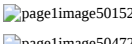
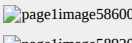
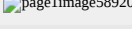
Report

Synthetic Polymers and Plastics

Part A: Physical Characteristics

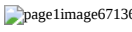
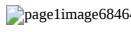
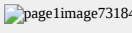
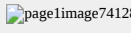
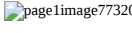
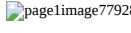
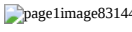
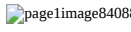
Find or choose one type of each of the following plastic polymers, and report the following characteristics:

Plastic number	Short Name (HDPE, LDPE, etc)	Clear (yes or no)	Opaque (yes or no)	Flexibility (can be bent?)	Durability (hard or soft)	Breakability (can cracked?)	Recyclable (yes or no)
1							
2							
3							
4			 				
5							
6			 				
7			 				

Part B: Density Tests

Report for plastic samples in each liquid: sinks rapidly, sinks slowly, floats on top, floats below surface

Plastic number	1:1 ethanol/water density = 0.94 g/cm ³	 Water density = 1.0 g/cm ³	 10% NaCl solution density = 1.08 g/cm ³
1			
2			
3			
4			
5			
6			

Relative Plastic Densities:

Less than 0.94 g/cm ³	Less than 1.0 g/cm ³	Less than 1.08 g/cm ³	More than 1.08 g/cm ³

Ranking of densities:

(lowest) _____ (highest)

Part C: Polymer Bouncy Balls

Polymer Ball composition	Approximate height bounced	Physical characteristics
Ball #1:		
Ball #2		
Ball #3		

Questions

- Which of the Big Six plastics was the most flexible?
- Which of the Big Six plastics would be the best material for each of the following examples? Use short names to identify each plastic (e.g. HDPE).
 - a replacement for a glass window ?
 - a take-out container for food?
 - a flexible, expandable bag for carrying items? a lightweight bottle cap?
- An unknown plastic floats in a 10% NaCl solution but sinks in water. What is the range of possible density values this plastic may have? Suggest the composition of this plastic.
- Why is it important to dislodge any adhering bubbles in the density tests?
- PET plastic (number 1) is the most valuable waste plastic at the present time. Suggest a way to separate it commercially from other waste plastics.
- Sometimes plastic containers are made from two polymers and not just one. What would happen to the water density test if HDPE and PVC were mixed?
- Why are plastic recyclers very concerned about identifying the different polymers and not mixing them together?
- The figure below depicts polymerization of polystyrene (PS). Circle the original monomers and determine how many monomers are present.
- Polyvinylchloride (PVC) is composed of the vinyl chloride monomer. The monomer structure and general reaction are shown at right.

Draw a polyvinyl chloride polymer composed of five monomers arranged in a head-to-tail pattern.

- For the bouncy balls you made, what is the name of the monomer?
What is the role of each of the following in the formation of the polymer?

glue

borax

cornstarch

- Which ball bounced the highest? Based on your data in the table, which compound was most likely responsible for this?

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SECTION OVERVIEW

Unit 1

1: Matter

- 1.1: What is Matter?
- 1.2: Classifying Matter According to Its State: Solid, Liquid, and Gas
- 1.3: Classifying Matter According to Its Composition
- 1.4: Differences in Matter: Physical and Chemical Properties
- 1.5: Changes in Matter - Physical and Chemical Changes
- 1.6: Cutting Aluminum until you get Atoms
- 1.7: Indivisible - The Atomic Theory
- 1.8: The Nuclear Atom
- 1.9: The Properties of Protons, Neutrons, and Electrons
- 1.10: Elements: Defined by Their Number of Protons
- 1.11: Looking for Patterns: The Periodic Table
- 1.12: Ions - Losing and Gaining Electrons
- 1.13: Compounds Display Constant Composition
- 1.14: Chemical Formulas: How to Represent Compounds
- 1.15: Naming Molecular Compounds
- 1.16: Structure and Composition of the Atmosphere
- 1.17: Conservation of Mass - There is No New Matter
- 1.18: Evidence of a Chemical Reaction
- 1.19: Chemical Equations
- 1.20: How to Write Balanced Chemical Equations

A: Preliminary

- 2.1: The Scope of Chemistry
- 2.2: Hypothesis, Theories, and Laws
- 2.3: Scientific Notation - Writing Large and Small Numbers
- 2.4: Significant Figures - Writing Numbers to Reflect Precision
- 2.5: Significant Figures in Calculations
- 2.6: The Basic Units of Measurement

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CHAPTER OVERVIEW

1: Matter

- 1.1: What is Matter?
- 1.2: Classifying Matter According to Its State: Solid, Liquid, and Gas
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1.1: What is Matter?

Learning Objectives

- Define matter and explain how it is composed of building blocks known as "atoms".

We are all familiar with matter. The definition of **Matter** is anything that has mass and volume (takes up space). For most common objects that we deal with every day, it is fairly simple to demonstrate that they have mass and take up space. You might be able to imagine, however, the difficulty for people several hundred years ago to demonstrate that air had mass and volume. Air (and all other gases) are invisible to the eye, have very small masses compared to equal amounts of solids and liquids, and are quite easy to compress (change volume). Without sensitive equipment, it would have been difficult to convince people that gases are matter. Today, we can measure the mass of a small balloon when it is deflated and then blow it up, tie it off, and measure its mass again to detect the additional mass due to the air inside. The mass of air, under room conditions, that occupies a one quart jar is approximately 0.0002 pounds. This small amount of mass would have been difficult to measure in times before balances were designed to accurately measure very small masses. Later, scientists were able to compress gases into such a small volume that the gases turned into liquids, which made it clear that gases are matter.



Figure 1.1.1: Everything from an ant, to a truck, to the earth, and even the entire galaxy is composed of matter. Images used with permission from Wikipedia (CC_SA-BY-3.0; credit [High Contrast](#)).

Even though the universe consists of "things" as wildly different as ants and galaxies, the matter that makes up all of these "things" is composed of a very limited number of building blocks. These building blocks are known as **atoms**, and so far, scientists have discovered or created a grand total of 118 different types of atoms. Scientists have given a name to each different type of atom. A substance that is composed of only one type of atom is called an element. At this point, what should amaze you is that all forms of matter in our universe are made with only 118 different building blocks. In some ways, it's sort of like cooking a gourmet, five-course meal using only three ingredients! How is it possible? To answer that question, you have to understand the ways in which different elements are put together to form matter.

The most important method that nature uses to organize atoms into matter is the formation of **molecules**. Molecules are groups of two or more atoms that have been bonded together. There are millions of different ways to bond atoms together, which means that

there are millions of different possible molecules. Each of these molecules has its own set of chemical properties, and it's these properties with which chemists are most concerned. You will learn a lot more about atoms and molecules, including how they were discovered, in a later part of the textbook.

Summary

All matter has mass and occupies space. All physical objects are made of matter. Matter itself is composed of tiny building blocks known as "atoms". There are only 118 different types of atoms known to man. Frequently, atoms are bonded together to form "molecules".

Contributions & Attributions

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1.2: Classifying Matter According to Its State: Solid, Liquid, and Gas

Learning Objectives

- To describe the solid, liquid and gas phases.

Water can take many forms. At low temperatures (below 0°C), it is a solid. When at "normal" temperatures (between 0°C and 100°C), it is a liquid. While at temperatures above 100°C , water is a gas (steam). The state that water is in depends upon the temperature. Each state has its own unique set of physical properties. Matter typically exists in one of three states: **solid**, **liquid**, or **gas**.

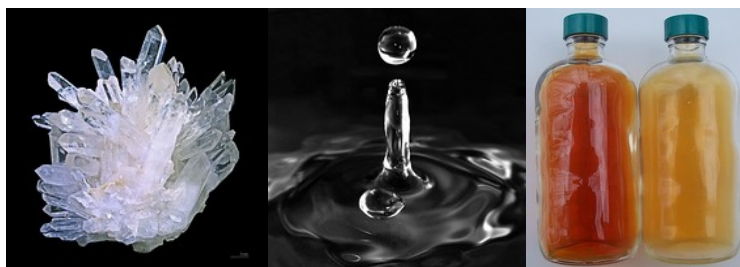


Figure 1.2.1: Matter is usually classified into three classical states, with plasma sometimes added as a fourth state. From left to right: quartz (solid), water (liquid), nitrogen dioxide (gas).

The state that a given substance exhibits is also a physical property. Some substances exist as gases at room temperature (oxygen and carbon dioxide), while others, like water and mercury metal, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states. Figure 1.2.2 shows the differences among solids, liquids, and gases at the molecular level. A solid has definite volume and shape, a liquid has a definite volume but no definite shape, and a gas has neither a definite volume nor shape.

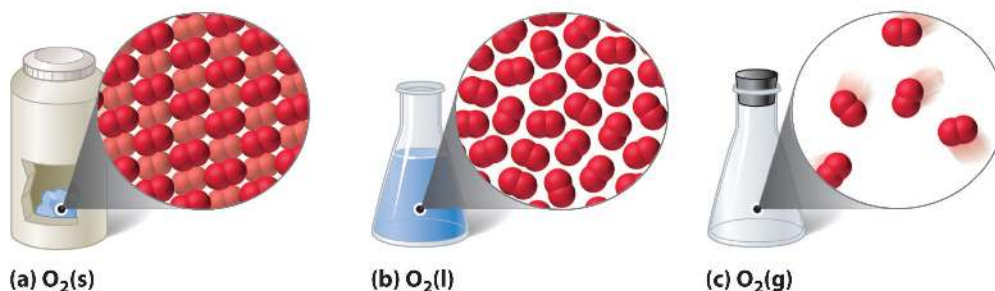
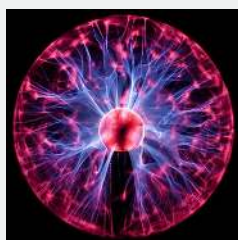


Figure 1.2.2: A Representation of the Solid, Liquid, and Gas States. (a) Solid O_2 has a fixed volume and shape, and the molecules are packed tightly together. (b) Liquid O_2 conforms to the shape of its container but has a fixed volume; it contains relatively densely packed molecules. (c) Gaseous O_2 fills its container completely—regardless of the container's size or shape—and consists of widely separated molecules.

Plasma: A Fourth State of Matter

Technically speaking, a fourth state of matter called plasma exists, but it does not naturally occur on earth, so we will omit it from our study here.



Solids

In the solid state, the individual particles of a substance are in fixed positions with respect to each other because there is not enough thermal energy to overcome the intermolecular interactions between the particles. As a result, solids have a definite shape and volume. Most solids are hard, but some (like waxes) are relatively soft. Many solids composed of ions can also be quite brittle.

Solids are defined by the following characteristics:

- Definite shape (rigid)
- Definite volume
- Particles vibrate around fixed axes

If we were to cool liquid mercury to its freezing point of -39°C , and under the right pressure conditions, we would notice all of the liquid particles would go into the solid state. Mercury can be solidified when its temperature is brought to its freezing point. However, when returned to room temperature conditions, mercury does not exist in solid state for long, and returns back to its more common liquid form.

Solids usually have their constituent particles arranged in a regular, three-dimensional array of alternating positive and negative ions called a **crystal**. The effect of this regular arrangement of particles is sometimes visible macroscopically, as shown in Figure 1.2.3. Some solids, especially those composed of large molecules, cannot easily organize their particles in such regular crystals and exist as amorphous (literally, “without form”) solids. Glass is one example of an amorphous solid.

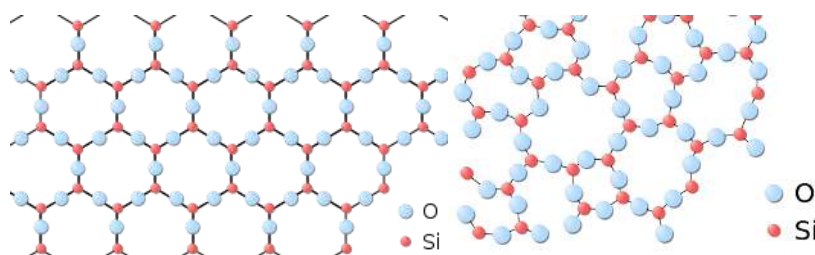


Figure 1.2.3: (left) The periodic crystalline lattice structure of quartz SiO_2 in two-dimensions. (right) The random network structure of glassy SiO_2 in two-dimensions. Note that, as in the crystal, each Silicon atom is bonded to 4 oxygen atoms, where the fourth oxygen atom is obscured from view in this plane. Images used with permission (public domain).

Liquids

If the particles of a substance have enough energy to partially overcome intermolecular interactions, then the particles can move about each other while remaining in contact. This describes the liquid state. In a liquid, the particles are still in close contact, so liquids have a definite volume. However, because the particles can move about each other rather freely, a liquid has no definite shape and takes a shape dictated by its container.

Liquids have the following characteristics:

- No definite shape (takes the shape of its container).
- Has definite volume.
- Particles are free to move over each other, but are still attracted to each other.

A familiar liquid is mercury metal. Mercury is an anomaly. It is the only metal we know of that is liquid at room temperature. Mercury also has an ability to stick to itself (surface tension)—a property that all liquids exhibit. Mercury has a relatively high surface tension, which makes it very unique. Here you see mercury in its common liquid form.



Video 1.2.1: *Mercury boiling to become a gas.*

If we heat liquid mercury to its boiling point of 357°C under the right pressure conditions, we would notice all particles in the liquid state go into the gas state.

Gases

If the particles of a substance have enough energy to completely overcome intermolecular interactions, then the particles can separate from each other and move about randomly in space. This describes the gas state, which we will consider in more detail elsewhere. Like liquids, gases have no definite shape, but unlike solids and liquids, gases have no definite volume either. The change from solid to liquid usually does not significantly change the volume of a substance. However, the change from a liquid to a gas significantly increases the volume of a substance, by a factor of 1,000 or more. Gases have the following characteristics:

- No definite shape (takes the shape of its container)
- No definite volume
- Particles move in random motion with little or no attraction to each other
- Highly compressible

Table 1.2.1: Characteristics of the Three States of Matter

Characteristics	Solids	Liquids	Gases
shape	definite	indefinite	indefinite
volume	definite	definite	indefinite
relative intermolecular interaction strength	strong	moderate	weak
relative particle positions	in contact and fixed in place	in contact but not fixed	not in contact, random positions

✓ Example 1.2.1

What state or states of matter does each statement, describe?

- This state has a definite volume, but no definite shape.
- This state has no definite volume.
- This state allows the individual particles to move about while remaining in contact.

Solution

- This statement describes the liquid state.
- This statement describes the gas state.
- This statement describes the liquid state.

? Exercise 1.2.1

What state or states of matter does each statement describe?

- This state has individual particles in a fixed position with regard to each other.
- This state has individual particles far apart from each other in space.
- This state has a definite shape.

Answer a:

solid

Answer b:

gas

Answer c:

solid

Summary

- Three states of matter exist—solid, liquid, and gas.
- Solids have a definite shape and volume.
- Liquids have a definite volume, but take the shape of the container.
- Gases have no definite shape or volume.

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1.3: Classifying Matter According to Its Composition

Learning Objectives

- Explain the difference between a pure substance and a mixture.
- Explain the difference between an element and a compound.
- Explain the difference between a homogeneous mixture and a heterogeneous mixture.

One useful way of organizing our understanding of matter is to think of a hierarchy that extends down from the most general and complex to the simplest and most fundamental (Figure 1.3.1). Matter can be classified into two broad categories: pure substances and mixtures. A **pure substance** is a form of matter that has a constant composition (meaning that it is the same everywhere) and properties that are constant throughout the sample (meaning that there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. Elements and compounds are both examples of pure 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 can be broken down into chemically simpler components (because it has more than one element) is a **compound**. For example, water is a compound composed of the elements hydrogen and oxygen. Today, there are about 118 elements in the known universe. In contrast, scientists have identified tens of millions of different compounds to date.

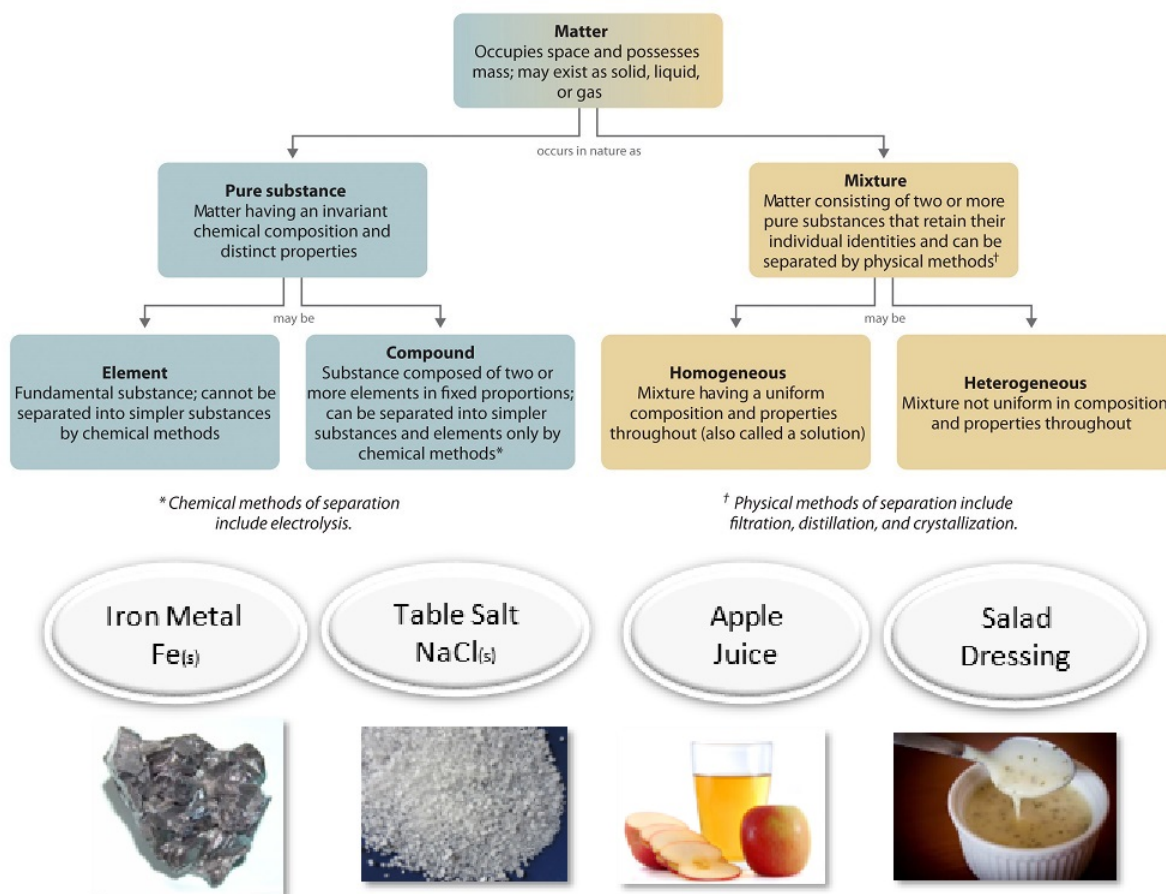


Figure 1.3.1: Relationships between the Types of Matter and the Methods Used to Separate Mixtures

Ordinary table salt is called sodium chloride. It is considered a **substance** because it has a uniform and definite composition. All samples of sodium chloride are chemically identical. Water is also a pure substance. Salt easily dissolves in water, but salt water cannot be classified as a substance because its composition can vary. You may dissolve a small amount of salt or a large amount into a given amount of water. A mixture is a physical blend of two or more components, each of which retains its own identity and

properties in the **mixture**. Only the form of the salt is changed when it is dissolved into water. It retains its composition and properties.

A **homogeneous** mixture is a mixture in which the composition is uniform throughout the mixture. The salt water described above is homogeneous because the dissolved salt is evenly distributed throughout the entire salt water sample. Often it is easy to confuse a homogeneous mixture with a pure substance because they are both uniform. The difference is that the composition of the substance is always the same. The amount of salt in the salt water can vary from one sample to another. All solutions are considered homogeneous because the dissolved material is present in the same amount throughout the solution.

A **heterogeneous mixture** is a mixture in which the composition is not uniform throughout the mixture. Vegetable soup is a heterogeneous mixture. Any given spoonful of soup will contain varying amounts of the different vegetables and other components of the soup.

Phase

A phase is any part of a sample that has a uniform composition and properties. By definition, a pure substance or a homogeneous mixture consists of a single phase. A heterogeneous mixture consists of two or more phases. When oil and water are combined, they do not mix evenly, but instead form two separate layers. Each of the layers is called a phase.

✓ Example 1.3.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

- a. filtered tea
- b. freshly squeezed orange juice
- c. a compact disc
- d. aluminum oxide, a white powder that contains a 2:3 ratio of aluminum and oxygen atoms
- e. selenium

Given: a chemical substance

Asked for: its classification

Strategy:

- A. Decide whether a substance is chemically pure. If it is pure, the substance is either an element or a compound. If a substance can be separated into its elements, it is a compound.
- B. If a substance is not chemically pure, it is either a heterogeneous mixture or a homogeneous mixture. If its composition is uniform throughout, it is a homogeneous mixture.

Solution

- a. **A)** Tea is a solution of compounds in water, so it is not chemically pure. It is usually separated from tea leaves by filtration.
B) Because the composition of the solution is uniform throughout, it is **a homogeneous mixture**.
- b. **A)** Orange juice contains particles of solid (pulp) as well as liquid; it is not chemically pure.
B) Because its composition is not uniform throughout, orange juice is **a heterogeneous mixture**.
- c. **A)** A compact disc is a solid material that contains more than one element, with regions of different compositions visible along its edge. Hence, a compact disc is not chemically pure.
B) The regions of different composition indicate that a compact disc is **a heterogeneous mixture**.
- d. **A)** Aluminum oxide is a single, chemically **pure compound**.
- e. **A)** Selenium is one of the known **elements**.

? Exercise 1.3.1

Identify each substance as a compound, an element, a heterogeneous mixture, or a homogeneous mixture (solution).

- a. white wine
- b. mercury
- c. ranch-style salad dressing
- d. table sugar (sucrose)

Answer a:

homogeneous mixture (solution)

Answer b:

element

Answer c:

heterogeneous mixture

Answer d:

compound

✓ Example 1.3.2

How would a chemist categorize each example of matter?

- a. saltwater
- b. soil
- c. water
- d. oxygen

Solution

- a. Saltwater acts as if it were a single substance even though it contains two substances—salt and water. Saltwater is a homogeneous mixture, or a solution.
- b. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
- c. Water is a substance. More specifically, because water is composed of hydrogen and oxygen, it is a compound.
- d. Oxygen, a substance, is an element.

? Exercise 1.3.2

How would a chemist categorize each example of matter?

- a. coffee
- b. hydrogen
- c. an egg

Answer a:

a homogeneous mixture (solution), assuming it is filtered coffee

Answer b:

element

Answer c:

heterogeneous mixture

Summary

Matter can be classified into two broad categories: pure substances and mixtures. A pure substance is a form of matter that has a constant composition and properties that are constant throughout the sample. Mixtures are physical combinations of two or more

elements and/or compounds. Mixtures can be classified as homogeneous or heterogeneous. Elements and compounds are both examples of pure substances. Compounds are substances that are made up of more than one type of atom. Elements are the simplest substances made up of only one type of atom.

Vocabulary

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

Contributions & Attributions

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- Stephen Lower, Professor Emeritus ([Simon Fraser U.](#)) [Chem1 Virtual Textbook](#)

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1.4: Differences in Matter: Physical and Chemical Properties

Learning Objectives

To separate physical from chemical properties.

All matter has physical and chemical properties. Physical properties are characteristics that scientists can measure without changing the composition of the sample under study, such as mass, color, and volume (the amount of space occupied by a sample). Chemical properties describe the characteristic ability of a substance to react to form new substances; they include its flammability and susceptibility to corrosion. All samples of a pure substance have the same chemical and physical properties. For example, pure copper is always a reddish-brown solid (a physical property) and always dissolves in dilute nitric acid to produce a blue solution and a brown gas (a chemical property).

Physical Property

A *physical property* is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Silver is a shiny metal that conducts electricity very well. It can be molded into thin sheets, a property called malleability. Salt is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Physical properties of matter include color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point.

For the elements, color does not vary much from one element to the next. The vast majority of elements are colorless, silver, or gray. Some elements do have distinctive colors: sulfur and chlorine are yellow, copper is (of course) copper-colored, and elemental bromine is red. However, density can be a very useful parameter for identifying an element. Of the materials that exist as solids at room temperature, iodine has a very low density compared to zinc, chromium, and tin. Gold has a very high density, as does platinum. Pure water, for example, has a density of 0.998 g/cm^3 at 25°C . The average densities of some common substances are in Table 1.4.1. Notice that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will “float.”

Table 1.4.1: Densities of Common Substances

Substance	Density at 25°C (g/cm^3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Hardness helps determine how an element (especially a metal) might be used. Many elements are fairly soft (silver and gold, for example) while others (such as titanium, tungsten, and chromium) are much harder. Carbon is an interesting example of hardness. In graphite, (the “lead” found in pencils) the carbon is very soft, while the carbon in a diamond is roughly seven times as hard.



Figure 1.4.1: Pencil (left) and Diamond ring (right). Both are a form of carbon, but exhibit very different physical properties.

Melting and boiling points are somewhat unique identifiers, especially of compounds. In addition to giving some idea as to the identity of the compound, important information can be obtained about the purity of the material.

Chemical Properties

Chemical properties of matter describe its potential to undergo some chemical change or reaction by virtue of its composition. The elements, electrons, and bonds that are present give the matter potential for chemical change. It is quite difficult to define a chemical property without using the word "change". Eventually, after studying chemistry for some time, you should be able to look at the formula of a compound and state some chemical property. For example, hydrogen has the potential to ignite and explode given the right conditions—this is a chemical property. Metals in general have the chemical property of reacting with an acid. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property.

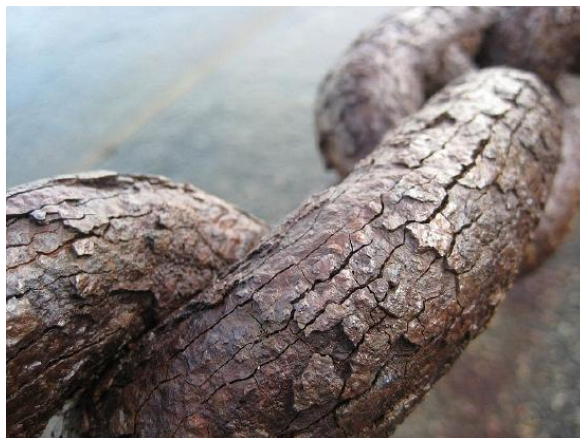


Figure 1.4.2: Heavy rust on the links of a chain near the Golden Gate Bridge in San Francisco; it was continuously exposed to moisture and salt spray, causing surface breakdown, cracking, and flaking of the metal. (CC BY-SA 3.0; Marlith).

A chemical property of iron is its capability of combining with oxygen to form iron oxide, the chemical name of rust (Figure 1.4.2). The more general term for rusting and other similar processes is corrosion. Other terms that are commonly used in descriptions of chemical changes are burn, rot, explode, decompose, and ferment. Chemical properties are very useful in identifying substances. However, unlike physical properties, chemical properties can only be observed as the substance is in the process of being changed into a different substance.

Table 1.4.2: Contrasting Physical and Chemical Properties

Physical Properties	Chemical Properties
Gallium metal melts at 30 °C.	Iron metal rusts.
Mercury is a very dense liquid.	A green banana turns yellow when it ripens.
Gold is shiny.	A dry piece of paper burns.

✓ Example 1.4.1

Which of the following is a chemical property of iron?

- a. Iron corrodes in moist air.
- b. Density = 7.874 g/cm^3
- c. Iron is soft when pure.
- d. Iron melts at 1808 K.

Solution

"Iron corrodes in moist air" is the only chemical property of iron from the list.

? Exercise 1.4.1A

Which of the following is a physical property of matter?

- a. corrosiveness
- b. pH (acidity)
- c. density
- d. flammability

Answer

c

? Exercise 1.4.1B

Which of the following is a chemical property?

- a. flammability
- b. melting point
- c. boiling point
- d. density

Answer

a

Summary

A physical property is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties include color, density, hardness, and melting and boiling points. A chemical property describes the ability of a substance to undergo a specific chemical change. To identify a chemical property, we look for a chemical change. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed.

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1.5: Changes in Matter - Physical and Chemical Changes

Learning Objectives

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

Change is happening all around us all of the time. Just as chemists have classified elements and compounds, they have also classified types of changes. Changes are classified as either physical or chemical changes. Chemists learn a lot about the nature of matter by studying the changes that matter can undergo. Chemists make a distinction between two different types of changes that they study—physical changes and chemical changes.

Physical Change

Physical changes are changes in which no bonds are broken or formed. This means that the same types of compounds or elements that were there at the beginning of the change are there at the end of the change. Because the ending materials are the same as the beginning materials, the properties (such as color, boiling point, etc.) will also be the same. Physical changes involve moving molecules around, but not changing them. Some types of physical changes include:

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

As an ice cube melts, its shape changes as it acquires the ability to flow. However, its composition does not change. Melting is an example of a physical change. A physical change is a change to a sample of matter in which some properties of the material change, but the identity of the matter does not. When liquid water is heated, it changes to water vapor. However, even though the physical properties have changed, the molecules are exactly the same as before. We still have each water molecule containing two hydrogen atoms and one oxygen atom covalently bonded. When you have a jar containing a mixture of pennies and nickels and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of the pennies or the nickels—you've merely separated them into two groups. This would be an example of a physical change. Similarly, if you have a piece of paper, you don't change it into something other than a piece of paper by ripping it up. What was paper before you started tearing is still paper when you are done. Again, this is an example of a physical change.



Figure 1.5.1: Ice melting is a physical change. When liquid water (H_2O) freezes into a solid state (ice), it appears changed; however, this change is only physical, as the composition of the constituent molecules is the same: 11.19% hydrogen and 88.81% oxygen by mass. (Public Domain; Moussa).

Physical changes can further be classified as reversible or irreversible. The melted ice cube may be refrozen, so melting is a reversible physical change. Physical changes that involve a change of state are all reversible. Other changes of state include **vaporization** (liquid to gas), **freezing** (liquid to solid), and **condensation** (gas to liquid). Dissolving is also a reversible physical change. When salt is dissolved into water, the salt is said to have entered the aqueous state. The salt may be regained by boiling off the water, leaving the salt behind.

Chemical Change

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

One good example of a chemical change is burning a candle. The act of burning paper actually results in the formation of new chemicals (carbon dioxide and water) from the burning of the wax. Another example of a chemical change is what occurs when natural gas is burned in your furnace. This time, on the left there is a molecule of methane, CH_4 , and two molecules of oxygen, O_2 ; on the right are two molecules of water, H_2O , and one molecule of carbon dioxide, CO_2 . In this case, not only has the appearance changed, but the structure of the molecules has also changed. The new substances do not have the same chemical properties as the original ones. Therefore, this is a chemical change.



Figure 1.5.2: Burning of wax to generate water and carbon dioxide is a chemical reaction. (CC-SA-BY-3.0; Andrikkos)

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

Observations that help to indicate chemical change include:

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

✓ Example 1.5.1

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

- Boiling water.
- A nail rusting.
- A green solution and colorless solution are mixed. The resulting mixture is a solution with a pale green color.
- Two colorless solutions are mixed. The resulting mixture has a yellow precipitate.

Solution

- Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written: $\text{H}_2\text{O} (l) \rightarrow \text{H}_2\text{O} (g)$

- b. Chemical: The dark grey nail changes color to form an orange flaky substance (the rust); this must be a chemical change. Color changes indicate chemical change. The following reaction occurs: $\text{Fe} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3$
- c. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been spread together. No color *change* occurred or other evidence of chemical change.
- d. Chemical: the formation of a precipitate and the color change from colorless to yellow indicate a chemical change.

? Exercise 1.5.1

Label each of the following changes as a physical or chemical change.

- a. A mirror is broken.
- b. An iron nail corroded in moist air
- c. Copper metal is melted.
- d. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer a:

physical change

Answer b:

chemical change

Answer c:

physical change

Answer d:

chemical change

Separating Mixtures Through Physical Changes

Homogeneous mixtures (solutions) can be separated into their component substances by physical processes that rely on differences in some physical property, such as differences in their boiling points. Two of these separation methods are distillation and crystallization. [Distillation](#) makes use of differences in volatility, a measure of how easily a substance is converted to a gas at a given temperature. A simple distillation apparatus for separating a mixture of substances, at least one of which is a liquid. The most volatile component boils first and is condensed back to a liquid in the water-cooled condenser, from which it flows into the receiving flask. If a solution of salt and water is distilled, for example, the more volatile component, pure water, collects in the receiving flask, while the salt remains in the distillation flask.

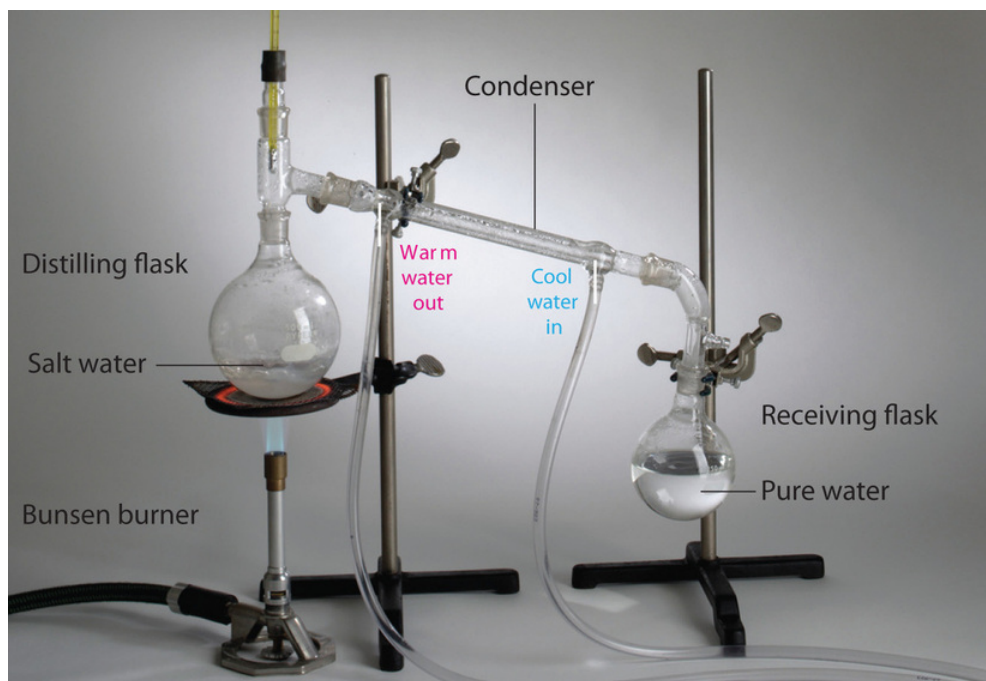


Figure 1.5.3: *The Distillation of a Solution of Table Salt in Water. The solution of salt in water is heated in the distilling flask until it boils. The resulting vapor is enriched in the more volatile component (water), which condenses to a liquid in the cold condenser and is then collected in the receiving flask.*

Parts of a distillation setup: Bunsen burner, salt water in distilling flask, condenser with cool water in and warm water out, pure water in receiving flask

Mixtures of two or more liquids with different boiling points can be separated with a more complex distillation apparatus. One example is the refining of crude petroleum into a range of useful products: aviation fuel, gasoline, kerosene, diesel fuel, and lubricating oil (in the approximate order of decreasing volatility). Another example is the distillation of alcoholic spirits such as brandy or whiskey. This relatively simple procedure caused more than a few headaches for federal authorities in the 1920s during the era of Prohibition, when illegal stills proliferated in remote regions of the United States.

Another example for using physical properties to separate mixtures is filtration (Figure 1.5.4). Filtration is any mechanical, physical or biological operation that separates solids from fluids (liquids or gases) by adding a medium through which only the fluid can pass. The fluid that passes through is called the filtrate. There are many different methods of filtration; all aim to attain the separation of substances. Separation is achieved by some form of interaction between the substance or objects to be removed and the filter. The substance that is to pass through the filter must be a fluid, i.e. a liquid or gas. Methods of filtration vary depending on the location of the targeted material, i.e. whether it is dissolved in the fluid phase or suspended as a solid.



Figure 1.5.4: Filtration for the separation of solids from a hot solution. (CC BY-SA 4.0; Suman6395).

Summary

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

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1.6: Cutting Aluminum until you get Atoms

Take some aluminum foil. Cut it in half. Now there are two smaller pieces of aluminum foil. Cut one of the pieces in half again. Cut one of those smaller pieces in half again. Continue cutting, making smaller and smaller pieces of aluminum foil. It should be obvious that the pieces are still aluminum foil; they are just becoming smaller and smaller. But how far can this exercise be taken, at least in theory? Can one continue cutting the aluminum foil into halves forever, making smaller and smaller pieces? Or is there some limit, some absolute smallest piece of aluminum foil? Thought experiments like this—and the conclusions based on them—were debated as far back as the fifth century BC by Democritus and other ancient Greek philosophers (Figure 1.6.1).

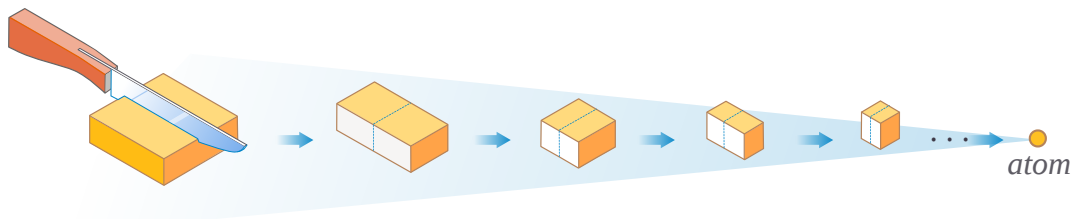


Figure 1.6.1: Democritus argued that matter, like an aluminum block, cannot be repeatedly cut in half perpetually. At some point, a limit is reached before the substance can no longer be called "aluminum"; this is the atomic limit. (CC BY-NC; Ümit Kaya via LibreTexts)

Most elements in their pure form exist as individual atoms. For example, a macroscopic chunk of iron metal is composed, microscopically, of individual atoms. Some elements, however, exist as groups of atoms called molecules. Several important elements exist as two-atom combinations and are called diatomic molecules. In representing a diatomic molecule, we use the symbol of the element and include the subscript 2 to indicate that two atoms of that element are joined together. The elements that exist as diatomic molecules are hydrogen (H_2), oxygen (O_2), nitrogen (N_2), fluorine (F_2), chlorine (Cl_2), bromine (Br_2), and iodine (I_2).

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1.7: Indivisible - The Atomic Theory

Learning Objectives

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

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

Democritus and the Greek Philosophers

About 2,500 years ago, early Greek philosophers believed the entire universe was a single, huge, entity. In other words, "everything was one." They believed that all objects, all matter, and all substances were connected as a single, big, unchangeable "thing." One of the first people to propose "atoms" was a man known as Democritus. As an alternative to the beliefs of the Greek philosophers, he suggested that **atomos**, or atomon—tiny, indivisible, solid objects—make up all matter in the universe.



Figure 1.7.1: (left) Democritus by Hendrick ter Brugghen, 1628. Democritus was known as the "laughing philosopher." It was a good thing he liked to laugh, because most other philosophers were laughing at his theories. (right) British physicist and chemist John Dalton (1766-1844). Unlike the Greek philosophers, John Dalton believed in both logical thinking and experimentation.

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

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

📌 Greeks: "All Thought and No Action"

Greek philosophers were "all thought and no action" and did not feel the need to test their theories with reality. In contrast, Dalton's efforts were based on experimentation and testing ideas against reality.

While it must be assumed that many more scientists, philosophers, and others studied composition of matter after Democritus, a major leap forward in our understanding of the composition of matter took place in the 1800's with the work of the British

scientists John Dalton. He started teaching school at age twelve, and was primarily known as a teacher. In his twenties, he moved to the growing city of Manchester, where he was able to pursue some scientific studies. His work in several areas of science brought him a number of honors. When he died, over 40,000 people in Manchester marched at his funeral.

The modern atomic theory, proposed about 1803 by the English chemist John Dalton (Figure 1.7.1), is a fundamental concept that states that all elements are composed of atoms. Previously, we defined an atom as the smallest part of an element that maintains the identity of that element. Individual atoms are extremely small; even the largest atom has an approximate diameter of only 5.4×10^{-10} m. With that size, it takes over 18 million of these atoms, lined up side by side, to equal the width of your little finger (about 1 cm).

Dalton studied the weights of various elements and compounds. He noticed that matter always combined in fixed ratios based on weight, or volume in the case of gases. Chemical compounds always contain the same proportion of elements by mass, regardless of amount, which provided further support for Proust's law of definite proportions. Dalton also observed that there could be more than one combination of two elements.

From his experiments and observations, as well as the work from peers of his time, Dalton proposed a new theory of the **atom**. This later became known as Dalton's **atomic theory**. The general tenets of this theory were as follows:

- All matter is composed of extremely small particles called atoms.
- Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
- Atoms cannot be subdivided, created, or destroyed.
- Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
- In chemical reactions, atoms are combined, separated, or rearranged.

Dalton's atomic theory has been largely accepted by the scientific community, with the exception of three changes. We know now that (1) an atom can be further subdivided, (2) all atoms of an element are not identical in mass, and (3) using nuclear fission and fusion techniques, we can create or destroy atoms by changing them into other atoms.

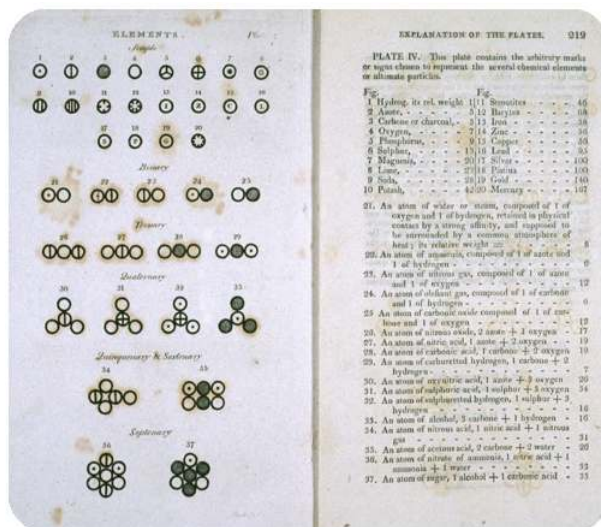


Figure 1.7.2: Dalton's symbols from his text "A New System of Chemical Philosophy."

The evidence for atoms is so great that few doubt their existence. In fact, individual atoms are now routinely observed with state-of-the-art technologies. Moreover, they can even be used for making pretty images; or as IBM research demonstrates in Video 1.7.1, control of individual atoms can be used to create animations.



Video 1.7.1: A Boy And His Atom: The World's Smallest Movie.

A Boy and His Atom is a 2012 stop-motion animated short film released by IBM Research. The movie tells the story of a boy and a wayward atom who meet and become friends. It depicts a boy playing with an atom that takes various forms. It was made by moving carbon monoxide molecules viewed with a scanning tunneling microscope, a device that magnifies them 100 million times. These molecules were moved to create images, which were then saved as individual frames to make the film.

Summary

- 2,500 years ago, Democritus suggested that all matter in the universe was made up of tiny, indivisible, solid objects he called "atomos." However, other Greek philosophers disliked Democritus' "atomos" theory because they felt it was illogical.
- Dalton's Atomic Theory is the first scientific theory to relate chemical changes to the structure, properties, and behavior of the atom. The general tenets of this theory are:
 - All matter is composed of extremely small particles called atoms.
 - Atoms of a given element are identical in size, mass, and other properties. Atoms of different elements differ in size, mass, and other properties.
 - Atoms cannot be subdivided, created, or destroyed.
 - Atoms of different elements can combine in simple whole number ratios to form chemical compounds.
 - In chemical reactions, atoms are combined, separated, or rearranged.

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1.8: The Nuclear Atom

Learning Objectives

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

Dalton's Atomic Theory held up well to a lot of the different chemical experiments that scientists performed to test it. In fact, for almost 100 years, it seemed as if Dalton's Atomic Theory was the whole truth. However, in 1897, a scientist named J. J. Thomson conducted some research that suggested that Dalton's Atomic Theory was not the entire story. He suggested that the small, negatively charged particles making up the cathode ray were actually pieces of atoms. He called these pieces "corpuscles," although today we know them as **electrons**. Thanks to his clever experiments and careful reasoning, J. J. Thomson is credited with the discovery of the electron.

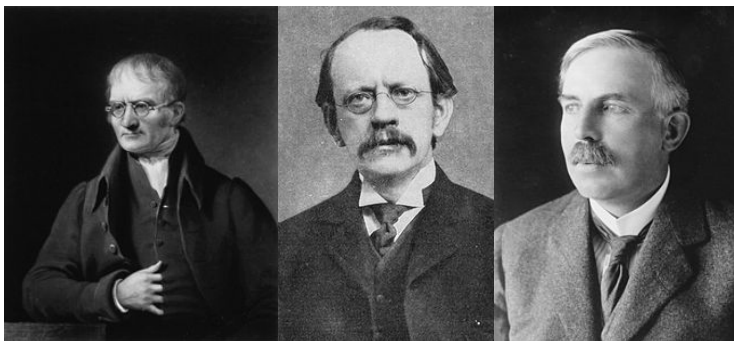


Figure 1.8.1: J. J. Thomson (center) concluded experiments that suggested that Dalton's (left) atomic theory was not telling the entire story. Ernest Rutherford (right) later built on Thomson's work to argue for the model nuclear atom.

Electrons and Plums

The electron was discovered by J. J. Thomson in 1897. The existence of protons was also known, as was the fact that atoms were neutral in charge. Since the intact atom had no net charge and the electron and proton had opposite charges, the next step after the discovery of subatomic particles was to figure out how these particles were arranged in the atom. This was a difficult task because of the incredibly small size of the atom. Therefore, scientists set out to design a model of what they believed the atom could look like. The goal of each **atomic model** was to accurately represent all of the experimental evidence about atoms in the simplest way possible.

Following the discovery of the electron, J.J. Thomson developed what became known as the "**plum pudding**" model in 1904. Plum pudding is an English dessert similar to a blueberry muffin. In Thomson's plum pudding model of the atom, the electrons were embedded in a uniform sphere of positive charge like blueberries stuck into a muffin. The positive matter was thought to be jelly-like or similar to a thick soup. The electrons were somewhat mobile. As they got closer to the outer portion of the atom, the positive charge in the region was greater than the neighboring negative charges, and the electron would be pulled back more toward the center region of the atom.

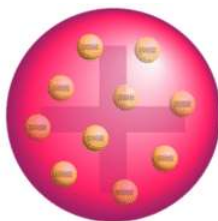


Figure 1.8.2: The "plum pudding" model.

However, this model of the atom soon gave way to a new model developed by New Zealander Ernest Rutherford (1871-1937) about five years later. Thomson did still receive many honors during his lifetime, including being awarded the Nobel Prize in Physics in 1906 and a knighthood in 1908.

Atoms and Gold

In 1911, Rutherford and coworkers Hans Geiger and Ernest Marsden initiated a series of groundbreaking experiments that would completely change the accepted model of the atom. They bombarded very thin sheets of gold foil with fast moving **alpha particles**. Alpha particles, a type of natural radioactive particle, are positively charged particles with a mass about four times that of a hydrogen atom.

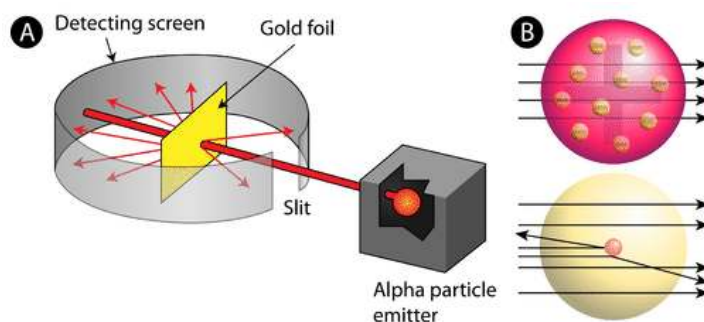


Figure 1.8.3: (A) The experimental setup for Rutherford's gold foil experiment: A radioactive element that emitted alpha particles was directed toward a thin sheet of gold foil that was surrounded by a screen which would allow detection of the deflected particles. (B) According to the plum pudding model (top), all of the alpha particles should have passed through the gold foil with little or no deflection. Rutherford found that a small percentage of alpha particles were deflected at large angles, which could be explained by an atom with a very small, dense, positively-charged nucleus at its center (bottom).

According to the accepted atomic model, in which an atom's mass and charge are uniformly distributed throughout the atom, the scientists expected that all of the alpha particles would pass through the gold foil with only a slight deflection or none at all. Surprisingly, while most of the alpha particles were indeed not deflected, a very small percentage (about 1 in 8000 particles) bounced off the gold foil at very large angles. Some were even redirected back toward the source. No prior knowledge had prepared them for this discovery. In a famous quote, Rutherford exclaimed that it was "as if you had fired a 15-inch [artillery] shell at a piece of tissue and it came back and hit you."

Rutherford needed to come up with an entirely new model of the atom in order to explain his results. Because the vast majority of the alpha particles had passed through the gold, he reasoned that most of the atom was empty space. In contrast, the particles that were highly deflected must have experienced a tremendously powerful force within the atom. He concluded that all of the positive charge and the majority of the mass of the atom must be concentrated in a very small space in the atom's interior, which he called the nucleus. The **nucleus** is the tiny, dense, central core of the atom and is composed of protons and neutrons.

Rutherford's atomic model became known as the **nuclear model**. In the nuclear atom, the protons and neutrons, which comprise nearly all of the mass of the atom, are located in the nucleus at the center of the atom. The electrons are distributed around the nucleus and occupy most of the volume of the atom. It is worth emphasizing just how small the nucleus is compared to the rest of the atom. If we could blow up an atom to be the size of a large professional football stadium, the nucleus would be about the size of a marble.

Rutherford's model proved to be an important step towards a full understanding of the atom. However, it did not completely address the nature of the electrons and the way in which they occupy the vast space around the nucleus. It was not until some years later that a full understanding of the electron was achieved. This proved to be the key to understanding the chemical properties of elements.

Atomic Nucleus

The **nucleus** (plural, nuclei) is a positively charged region at the center of the atom. It consists of two types of subatomic particles packed tightly together. The particles are protons, which have a positive electric charge, and neutrons, which are neutral in electric charge. Outside of the nucleus, an atom is mostly empty space, with orbiting negative particles called electrons whizzing through it. The figure below shows these parts of the atom.

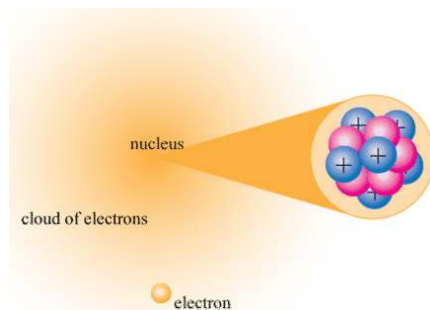


Figure 1.8.4: The nuclear atom.

The nucleus of the atom is extremely small. Its radius is only about 1/100,000 of the total radius of the atom. Electrons have virtually no mass, but protons and neutrons have a lot of mass for their size. As a result, the nucleus has virtually all the mass of an atom. Given its great mass and tiny size, the nucleus is very dense. If an object the size of a penny had the same density as the nucleus of an atom, its mass would be greater than 30 million tons!

Holding it all Together

Particles with opposite electric charges attract each other. This explains why negative electrons orbit the positive nucleus. Particles with the same electric charge repel each other. This means that the positive protons in the nucleus push apart from one another. So why doesn't the nucleus fly apart? An even stronger force—called the **strong nuclear force**—holds protons and neutrons together in the nucleus.

Summary

- Atoms are the ultimate building blocks of all matter.
- The modern atomic theory establishes the concepts of atoms and how they compose matter.
- Bombardment of gold foil with alpha particles showed that some particles were deflected.
- The nuclear model of the atom consists of a small and dense positively charged interior surrounded by a cloud of electrons.

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1.9: The Properties of Protons, Neutrons, and Electrons

Learning Objectives

- Describe the locations, charges, and masses of the three main subatomic particles.
- Determine the number of protons and electrons in an atom.
- Define atomic mass unit (amu).

Dalton's Atomic Theory explained a lot about matter, chemicals, and chemical reactions. Nevertheless, it was not entirely accurate, because contrary to what Dalton believed, atoms can, in fact, be broken apart into smaller subunits or subatomic particles. We have been talking about the electron in great detail, but there are two other particles of interest to us: protons and neutrons. We already learned that J. J. Thomson discovered a negatively charged particle, called the **electron**. Rutherford proposed that these electrons orbit a positive nucleus. In subsequent experiments, he found that there is a smaller positively charged particle in the nucleus, called a **proton**. There is also a third subatomic particle, known as a neutron.

Electrons

Electrons are one of three main types of particles that make up atoms. Unlike protons and neutrons, which consist of smaller, simpler particles, electrons are fundamental particles that do not consist of smaller particles. They are a type of fundamental particle called leptons. All leptons have an electric charge of -1 or 0 . Electrons are extremely small. The mass of an electron is only about $1/2000$ the mass of a proton or neutron, so electrons contribute virtually nothing to the total mass of an atom. Electrons have an electric charge of -1 , which is equal but opposite to the charge of a proton, which is $+1$. All atoms have the same number of electrons as protons, so the positive and negative charges "cancel out", making atoms electrically neutral.

Unlike protons and neutrons, which are located inside the nucleus at the center of the atom, electrons are found outside the nucleus. Because opposite electric charges attract one another, negative electrons are attracted to the positive nucleus. This force of attraction keeps electrons constantly moving through the otherwise empty space around the nucleus. The figure below is a common way to represent the structure of an atom. It shows the electron as a particle orbiting the nucleus, similar to the way that planets orbit the sun. However, this is an incorrect perspective, as quantum mechanics demonstrates that electrons are more complicated.



Figure 1.9.1: Electrons are much smaller than protons or neutrons. If an electron was the mass of a penny, a proton or a neutron would have the mass of a large bowling ball!

Protons

A proton is one of three main particles that make up the atom. Protons are found in the nucleus of the atom. This is a tiny, dense region at the center of the atom. Protons have a positive electrical charge of one ($+1$) and a mass of 1 atomic mass unit (**amu**), which is about 1.67×10^{-27} kilograms. Together with neutrons, they make up virtually all of the mass of an atom.

Neutrons

Atoms of all elements—except for most atoms of hydrogen—have neutrons in their nucleus. Unlike protons and electrons, which are electrically charged, neutrons have no charge—they are electrically neutral. That's why the neutrons in the diagram above are labeled n^0 . The zero stands for "zero charge". The mass of a neutron is slightly greater than the mass of a proton, which is 1 atomic

mass unit (**amu**). (An atomic mass unit equals about 1.67×10^{-27} kilograms.) A neutron also has about the same diameter as a proton, or 1.7×10^{-15} meters.

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

Before we move on, we must discuss how the different types of subatomic particles interact with each other. When it comes to neutrons, the answer is obvious. Since neutrons are neither attracted to nor repelled from objects, they don't really interact with protons or electrons (beyond being bound into the nucleus with the protons).

Even though electrons, protons, and neutrons are all types of subatomic particles, they are not all the same size. When you compare the masses of electrons, protons, and neutrons, what you find is that electrons have an extremely small mass, compared to either protons or neutrons. On the other hand, the masses of protons and neutrons are fairly similar, although technically, the mass of a neutron is slightly larger than the mass of a proton. Because protons and neutrons are so much more massive than electrons, almost all of the mass of any atom comes from the nucleus, which contains all of the neutrons and protons.

Table 1.9.1: Properties of Subatomic Particles

Particle	Symbol	Mass (amu)	Relative Mass (proton = 1)	Relative Charge	Location
proton	p^+	1	1	+1	inside the nucleus
electron	e^-	5.45×10^{-4}	0.00055	-1	outside the nucleus
neutron	n^0	1	1	0	inside the nucleus

Table 1.9.1 gives the properties and locations of electrons, protons, and neutrons. The third column shows the masses of the three subatomic particles in "atomic mass units." An **atomic mass unit (amu)** is defined as one-twelfth of the mass of a carbon-12 atom. Atomic mass units (**amu**) are useful, because, as you can see, the mass of a proton and the mass of a neutron are almost exactly 1 in this unit system.

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

Summary

- Electrons are a type of subatomic particle with a negative charge.
- Protons are a type of subatomic particle with a positive charge. Protons are bound together in an atom's nucleus as a result of the strong nuclear force.
- Neutrons are a type of subatomic particle with no charge (they are neutral). Like protons, neutrons are bound into the atom's nucleus as a result of the strong nuclear force.
- Protons and neutrons have approximately the same mass, but they are both much more massive than electrons (approximately 2,000 times as massive as an electron).
- The positive charge on a proton is equal in magnitude to the negative charge on an electron. As a result, a neutral atom must have an equal number of protons and electrons.
- The atomic mass unit (amu) is a unit of mass equal to one-twelfth the mass of a carbon-12 atom

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1.10: Elements: Defined by Their Number of Protons

Learning Objectives

- Define atomic number.
- Define mass number.
- Determine the number of protons, neutrons, and electrons in an atom.

It's important to be able to distinguish atoms of one element from atoms of another element. Elements are pure substances that make up all other matter, so each one is given a unique name. The names of elements are also represented by unique one- or two-letter symbols, such as **H** for hydrogen, **C** for carbon, or **He** for helium. However, it would more powerful if these names could be used to identify the numbers of protons and neutrons in the atoms. That's where atomic number and mass number are useful.



Figure 1.10.1: It is difficult to find qualities that differ between each element, and to distinguish one element from another. Each element, however, does have a unique number of protons. Sulfur has 16 protons, silicon has 14 protons, and gold has 79 protons. Images used with permission (public domain for sulfur and silicon, gold is licensed by CC-BY-SA-NC-ND; [Alchemist-hp](#)).

Atomic Number

Scientists distinguish between different elements by counting the number of protons in the nucleus (Table 1.10.1). If an atom has only one proton, we know that it's a hydrogen atom. An atom with two protons is always a helium atom. If scientists count four protons in an atom, they know it's a beryllium atom. An atom with three protons is a lithium atom, an atom with five protons is a boron atom, an atom with six protons is a carbon atom . . . the list goes on.

Since an atom of one element can be distinguished from an atom of another element by the number of protons in its nucleus, scientists are always interested in this number, and how this number differs between different elements. The number of protons in an atom is called its **atomic number** (Z). This number is very important because it is unique for atoms of a given element. All atoms of an element have the same number of protons, and every element has a different number of protons in its atoms. For example, all helium atoms have two protons, and no other elements have atoms with two protons.

Table 1.10.1: Atoms of the First Six Elements

Name	Protons	Neutrons	Electrons	Atomic Number (Z)	Mass Number (A)
Hydrogen	1	0	1	1	1
Helium	2	2	2	2	4
Lithium	3	4	3	3	7
Beryllium	4	5	4	4	9
Boron	5	6	5	5	11
Carbon	6	6	6	6	12

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

Neutral Atoms

Atoms are neutral in electrical charge because they have the same number of negative electrons as positive protons (Table 1.10.1). Therefore, the atomic number of an atom also tells you how many electrons the atom has. This, in turn, determines many of the atom's chemical properties.

Mass Number

The **mass number** (A) of an atom is the total number of protons and neutrons in its nucleus. The mass of the atom is a unit called the atomic mass unit (**amu**). One atomic mass unit is the mass of a proton, or about 1.67×10^{-27} kilograms, which is an extremely small mass. A neutron has just a tiny bit more mass than a proton, but its mass is often assumed to be one atomic mass unit as well. Because electrons have virtually no mass, just about all the mass of an atom is in its protons and neutrons. Therefore, the total number of protons and neutrons in an atom determines its mass in atomic mass units (Table 1.10.1).

Consider helium again. Most helium atoms have two neutrons in addition to two protons. Therefore the mass of most helium atoms is 4 atomic mass units (**2 amu** for the protons + **2 amu** for the neutrons). However, some helium atoms have more or less than two neutrons. Atoms with the same number of protons but different numbers of neutrons are called isotopes. Because the number of neutrons can vary for a given element, the mass numbers of different atoms of an element may also vary. For example, some helium atoms have three neutrons instead of two (these are called **isotopes** and are discussed in detail later on).

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

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

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

✓ Example 4.5.1

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

Solution

(number of protons) = 2 (Remember that an atom of helium always has 2 protons.)

(number of neutrons) = 2

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

$$\text{mass number} = 2 + 2 = 4$$

A **chemical symbol** is a one- or two-letter designation of an element. Some examples of chemical symbols are **O** for oxygen, **Zn** for zinc, and **Fe** for iron. The first letter of a symbol is always capitalized. If the symbol contains two letters, the second letter is lower case. The majority of elements have symbols that are based on their English names. However, some of the elements that have been known since ancient times have maintained symbols that are based on their Latin names, as shown in Table 1.10.2.

Table 1.10.2: Symbols and Latin Names for Elements

Chemical Symbol	Name	Latin Name
Na	Sodium	Natrium
K	Potassium	Kalium
Fe	Iron	Ferrum
Cu	Copper	Cuprum
Ag	Silver	Argentum
Sn	Tin	Stannum
Sb	Antimony	Stibium
Au	Gold	Aurum
Pb	Lead	Plumbum

Summary

- Elements are pure substances that make up all matter, so each one is given a unique name.
- The names of elements are also represented by unique one- or two-letter symbols.
- Each element has a unique number of protons. An element's atomic number is equal to the number of protons in the nuclei of any of its atoms.
- The mass number of an atom is the sum of the protons and neutrons in the atom.
- Isotopes are atoms of the same element (same number of protons) that have different numbers of neutrons in their atomic nuclei.

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- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.

In 1864, Julius Lothar Meyer, a German chemist, organized the elements by atomic mass and grouped them according to their chemical properties. Later that decade, Dmitri Mendeleev, a Russian chemist, organized all the known elements according to similar properties. He left gaps in his table for what he thought were undiscovered elements, and he made some bold predictions regarding the properties of those undiscovered elements. When elements were later discovered whose properties closely matched Mendeleev's predictions, his version of the table gained favor in the scientific community. Because certain properties of the elements repeat on a regular basis throughout the table (that is, they are periodic), it became known as the periodic table.

The **periodic table** is one of the cornerstones of chemistry because it organizes all of the known elements on the basis of their chemical properties. A modern version is shown in Figure 1.11.1. Most periodic tables provide additional data (such as atomic mass) in a box that contains each element's symbol. The elements are listed in order of atomic number.

Figure 1.11.1: A Modern Periodic Table. A modern periodic table lists elements left to right by atomic number. An interactive Periodic table can be found [Periodic Table of Elements](#). (Public Domain; PubChem via NIH)

Features of the Periodic Table

Elements that have similar chemical properties are grouped in columns called groups (or families). As well as being numbered, some of these groups have names—for example, *alkali metals* (the first column of elements), *alkaline earth metals* (the second column of elements), *halogens* (the next-to-last column of elements), and *noble gases* (the last column of elements).

The word halogen comes from the Greek for “salt maker” because these elements combine with other elements to form a group of compounds called salts.

To Your Health: Radon

Radon is an invisible, odorless noble gas that is slowly released from the ground, particularly from rocks and soils whose uranium content is high. Because it is a noble gas, radon is not chemically reactive. Unfortunately, it is radioactive, and increased exposure to it has been correlated with an increased lung cancer risk.

Because radon comes from the ground, we cannot avoid it entirely. Moreover, because it is denser than air, radon tends to accumulate in basements, which if improperly ventilated can be hazardous to a building’s inhabitants. Fortunately, specialized ventilation minimizes the amount of radon that might collect. Special fan-and-vent systems are available that draw air from below the basement floor, before it can enter the living space, and vent it above the roof of a house.

After smoking, radon is thought to be the second-biggest *preventable* cause of lung cancer in the United States. The American Cancer Society estimates that 10% of all lung cancers are related to radon exposure. There is uncertainty regarding what levels of exposure cause cancer, as well as what the exact causal agent might be (either radon or one of its breakdown products, many of which are also radioactive and, unlike radon, not gases). The US Environmental Protection Agency recommends testing every floor below the third floor for radon levels to guard against long-term health effects.

Each row of elements on the periodic table is called a period. Periods have different lengths; the first period has only 2 elements (hydrogen and helium), while the second and third periods have 8 elements each. The fourth and fifth periods have 18 elements each, and later periods are so long that a segment from each is removed and placed beneath the main body of the table.

Certain elemental properties become apparent in a survey of the periodic table as a whole. Every element can be classified as either a metal, a nonmetal, or a metalloid (or semi metal), as shown in Figure 1.11.2. A **metal** is a substance that is shiny, typically (but not always) silvery in color, and an excellent conductor of electricity and heat. Metals are also malleable (they can be beaten into thin sheets) and ductile (they can be drawn into thin wires). A **nonmetal** is typically dull and a poor conductor of electricity and heat. Solid nonmetals are also very brittle. As shown in Figure 1.11.2, metals occupy the left three-fourths of the periodic table, while nonmetals (except for hydrogen) are clustered in the upper right-hand corner of the periodic table. The elements with properties intermediate between those of metals and nonmetals are called **metalloids** (or **semi-metals**). Elements adjacent to the bold line in the right-hand portion of the periodic table have semimetal properties.

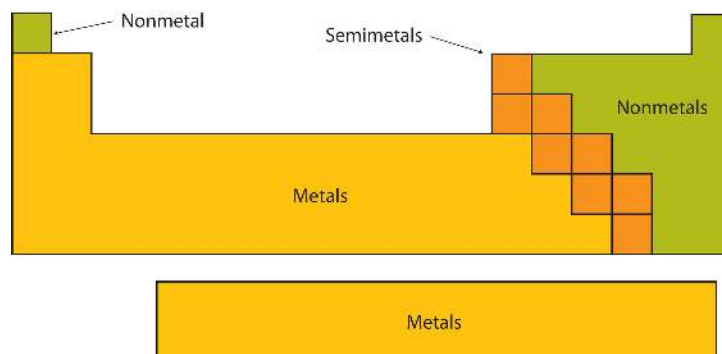


Figure 1.11.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table. (CC BY-NC-SA; Anonymous by request)

✓ Example 1.11.1

Based on its position in the periodic table, classify each element below as metal, a nonmetal, or a metalloid.

- Se
- Mg
- Ge

Solution

- In Figure 1.11.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
- Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
- Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

? Exercise 1.11.1

Based on its location in the periodic table, do you expect indium to be a nonmetal, a metal, or a metalloid?

Answer

Indium is a metal.

Another way to categorize the elements of the periodic table is shown in Figure 1.11.3. The first two columns on the left and the last six columns on the right are called the main group elements. The ten-column block between these columns contains the **transition metals**. The two rows beneath the main body of the periodic table contain the **inner transition metals**. The elements in these two rows are also referred to as, respectively, the **lanthanide metals** and the **actinide metals**.

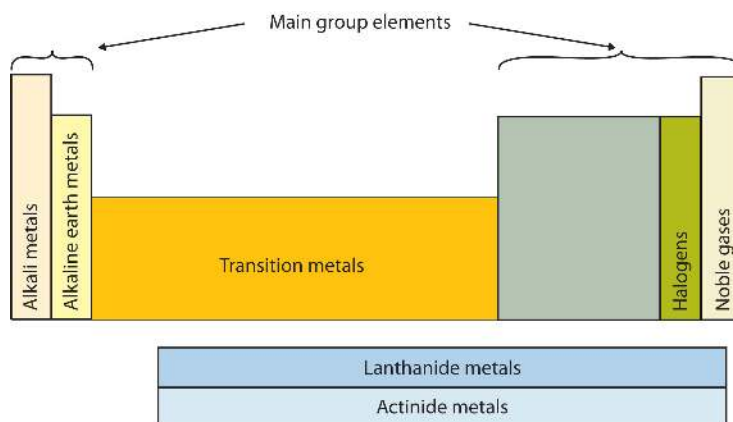


Figure 1.11.3: Special Names for Sections of the Periodic Table. (CC BY-NC-SA; Anonymous by request)

Sections of period table: Alkali metals, alkaline earth metals, transition metals, halogens, noble gases, lanthanide metals, actinide metals.

Descriptive Names

As previously noted, the periodic table is arranged so that elements with similar chemical behaviors are in the same group. Chemists often make general statements about the properties of the elements in a group using descriptive names with historical origins.

Group 1: The Alkali Metals

The alkali metals are lithium, sodium, potassium, rubidium, cesium, and francium. Hydrogen is unique in that it is generally placed in Group 1, but it is not a metal.

The compounds of the alkali metals are common in nature and daily life. One example is table salt (sodium chloride); lithium compounds are used in greases, in batteries, and as drugs to treat patients who exhibit manic-depressive, or bipolar, behavior. Although lithium, rubidium, and cesium are relatively rare in nature, and francium is so unstable and highly radioactive that it exists in only trace amounts, sodium and potassium are the seventh and eighth most abundant elements in Earth's crust, respectively.



Video: Alkali metals in water - Chemical elements: properties and reactions. (The Open University via <https://youtu.be/6ZY6d6jrj-0>)

Group 2: The Alkaline Earth Metals

The **alkaline earth metals** are beryllium, magnesium, calcium, strontium, barium, and radium. Beryllium, strontium, and barium are rare, and radium is unstable and highly radioactive. In contrast, calcium and magnesium are the fifth and sixth most abundant elements on Earth, respectively; they are found in huge deposits of limestone and other minerals.

Group 17: The Halogens

The **halogens** are fluorine, chlorine, bromine, iodine, and astatine. The name halogen is derived from the Greek words for “salt forming,” which reflects that all of the halogens react readily with metals to form compounds, such as sodium chloride and calcium chloride (used in some areas as road salt).

Compounds that contain the fluoride ion are added to toothpaste and the water supply to prevent dental cavities. Fluorine is also found in Teflon coatings on kitchen utensils. Although chlorofluorocarbon propellants and refrigerants are believed to lead to the depletion of Earth's ozone layer and contain both fluorine and chlorine, the latter is responsible for the adverse effect on the ozone layer. Bromine and iodine are less abundant than chlorine, and astatine is so radioactive that it exists in only negligible amounts in nature.

Group 18: The Noble Gases

The noble gases are helium, neon, argon, krypton, xenon, and radon. Because the noble gases are composed of only single atoms, they are called monatomic. At room temperature and pressure, they are unreactive gases. Because of their lack of reactivity, for many years they were called inert gases or rare gases. However, the first chemical compounds containing the noble gases were prepared in 1962. Although the noble gases are relatively minor constituents of the atmosphere, natural gas contains substantial amounts of helium. Because of its low reactivity, argon is often used as an unreactive (inert) atmosphere for welding and in light bulbs. The red light emitted by neon in a gas discharge tube is used in neon lights.

✓ Example 1.11.2: Groups

Provide the family or group name of each element.

- a. Li
- b. Ar
- c. Cl

Solution

- a. Lithium is an alkali metal (Group 1)
- b. Argon is a noble gas (Group 18)
- c. Chlorine is a halogen (Group 17)

? Exercise 1.11.2: Groups

Provide the family or group name of each element.

- a. F
- b. Ca
- c. Kr

Answer a:

Fluorine is a halogen (Group 17).

Answer b:

Calcium is an alkaline earth metal (Group 2).

Answer c:

Krypton is a noble gas (Group 18).

✓ Example 1.11.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

- a. Li
- b. Ar
- c. Am
- d. Fe

Solution

- a. Lithium is a metal.
- b. Argon is a non metal.
- c. Americium is an inner transition metal.
- d. Iron is a transition metal.

? Exercise 1.11.3: Classification of Elements

Classify each element as metal, non metal, transition metal or inner transition metal.

- a. F
- b. U
- c. Cu

Answer a:

Fluorine is a nonmetal.

Answer b:

Uranium is a metal (and an inner transition metal too).

Answer c:

Copper is a metal (and a transition metal too).

Summary

The periodic table is an arrangement of the elements in order of increasing atomic number. Elements that exhibit similar chemistry appear in vertical columns called groups (numbered 1–18 from left to right); the seven horizontal rows are called periods. Some of the groups have widely-used common names, including the alkali metals (Group 1) and the alkaline earth metals (Group 2) on the far left, and the halogens (Group 17) and the noble gases (Group 18) on the far right.

The elements can be broadly divided into metals, nonmetals, and semi metals. Semi metals exhibit properties intermediate between those of metals and nonmetals. Metals are located on the left of the periodic table, and nonmetals are located on the upper right. They are separated by a diagonal band of semi metals. Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors. Other important groupings of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).

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2. Sisler, Harry H. Electronic structure, properties, and the periodic law. New York; Reinhold publishing corporation, 1963.
3. Petrucci, Ralph H., Carey Bissonnette, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. Custom Edition for CHEM 2. Pearson Learning Solutions, 2010.

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1.12: Ions - Losing and Gaining Electrons

Learning Objectives

- Define the two types of ions.

Atoms tend to lose, gain, or share some valence electrons, making bonds to acquire the electron configuration of the nearest noble gas, i.e., usually eight valence electrons for the main group elements, called an octet. Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while others lack only one or two electrons to have an octet. In cases where an atom has three or fewer valence electrons, the atom may lose those valence electrons quite easily until what remains is a lower shell that contains an octet. Atoms that lose electrons acquire a positive charge because they are left with fewer negatively charged electrons to balance the positive charges of the protons in the nucleus. Positively charged ions are called **cations**. Most metals become cations when they make ionic compounds.

Cations

A neutral sodium atom is likely to achieve an octet in its outermost shell by losing its one valence electron.



The cation produced in this way, Na^+ , is called the sodium ion to distinguish it from the element. The outermost shell of the sodium ion is the second electron shell, which has eight electrons in it. The octet rule has been satisfied. Figure 1.12.1 is a graphical depiction of this process.

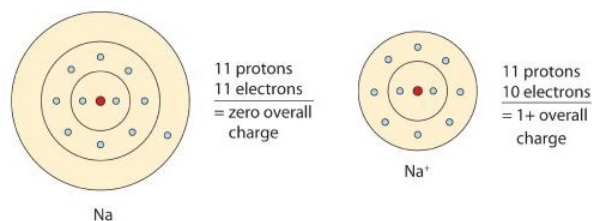
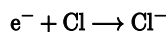


Figure 1.12.1: The Formation of a Sodium Ion. On the left, a sodium atom has 11 electrons. On the right, the sodium ion only has 10 electrons and a 1+ charge. Neutral sodium atom on left has 11 protons and 11 electrons. Sodium ion on right has 11 protons and 10 electrons, with a +1 overall charge.

Anions

Some atoms have nearly eight electrons in their valence shell and can gain additional valence electrons until they have an octet. When these atoms gain electrons, they acquire a negative charge because they now possess more electrons than protons. Negatively charged ions are called **anions**. Most nonmetals become anions when they make ionic compounds.

A neutral chlorine atom has seven electrons in its outermost shell. Only one more electron is needed to achieve an octet in chlorine's valence shell. (In table salt, this electron comes from the sodium atom.)



In this case, the ion has the same outermost shell as the original atom, but now that shell has eight electrons in it. Once again, the octet rule has been satisfied. The resulting anion, Cl^- , is called the chloride ion; note the slight change in the suffix (*-ide* instead of *-ine*) to create the name of this anion. Figure 1.12.2 is a graphical depiction of this process.

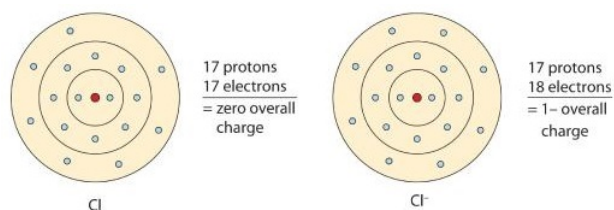


Figure 1.12.2: The Formation of a Chlorine Ion. On the left, the chlorine atom has 17 electrons. On the right, the chloride ion has 18 electrons and has a 1⁻ charge. Neutral chlorine atom on left has 17 protons and 17 electrons. Sodium ion on right has 17 protons and 18 electrons, with a -1 overall charge.

The names for positive and negative ions are pronounced CAT-eye-ons and ANN-eye-ons, respectively.

In many cases, elements that belong to the same group (vertical column) on the periodic table form ions with the same charge because they have the same number of valence electrons. Thus, the periodic table becomes a tool for remembering the charges on many ions. For example, all ions made from alkali metals, the first column on the periodic table, have a 1⁺ charge. Ions made from alkaline earth metals, the second group on the periodic table, have a 2⁺ charge. On the other side of the periodic table, the next-to-last column, the halogens, form ions having a 1⁻ charge. Figure 1.12.3 shows how the charge on many ions can be predicted by the location of an element on the periodic table. Note the convention of first writing the number and then the sign on a ion with multiple charges. The barium cation is written Ba²⁺, not Ba⁺².

1A							8A
H ⁺	2A						
Li ⁺				N ³⁻	O ²⁻	F ⁻	
Na ⁺	Mg ²⁺			P ³⁻	S ²⁻	Cl ⁻	
K ⁺	Ca ²⁺				Se ²⁻	Br ⁻	
Rb ⁺	Sr ²⁺					I ⁻	

Figure 1.12.3: Predicting Ionic Charges. 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.

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1.13: Compounds Display Constant Composition

When building a house, the starting point is a blueprint of what the house will look like. The plan states how many windows and what kind, how many doors and what style, how many rooms and what type (bedroom, kitchen, other). The blueprint shows how the different pieces will go together to make the house. As long as the blueprint is followed and exactly the same items are used, the house will be identical to its blueprint.

Compounds

A **compound** is a substance that contains two or more elements chemically combined in a fixed proportion. The elements carbon and hydrogen combine to form many different compounds. One of the simplest is called methane, in which there are always four times as many hydrogen particles as carbon particles. Methane is a pure substance because it always has the same composition. However, it is not an element because it can be broken down into simpler substances—carbon and hydrogen.

Recall that the components of a mixture can be separated from one another by physical means. This is not true for a compound. Table salt is a compound consisting of equal parts of the elements sodium and chlorine. Salt cannot be separated into its two elements by filtering, distillation, or any other physical process. Salt and other compounds can only be decomposed into their elements by a chemical process. A **chemical change** is a change that produces matter with a different composition. Many compounds can be decomposed into their elements by heating. When sugar is heated, it decomposes into carbon and water. Water is still a compound, but one which cannot be broken down into hydrogen and oxygen by heating. Instead, the passage of an electrical current through water will produce hydrogen and oxygen gases.

The properties of compounds are generally very different than the properties of the elements from which the compound is formed. Sodium is an extremely reactive soft metal that cannot be exposed to air or water. Chlorine is a deadly gas. The compound sodium chloride is a white solid which is essential for all living things (see below).



Figure 1.13.1: (A) Sodium is so reactive that it must be stored under oil. (B) Chlorine is a poisonous yellow-green gas. (C) Salt crystals, a compound of sodium and chlorine.

Summary

- A compound is a substance that contains two or more elements chemically combined in a fixed proportion.
- A chemical change is a change that produces matter with a different composition.

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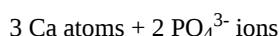
1.14: Chemical Formulas: How to Represent Compounds

Learning Objectives

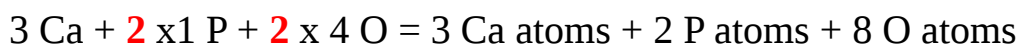
- Determine the number of different atoms in a formula.
- Define chemical formula, molecular formula, and empirical formula.

A **chemical formula** is an expression that shows the elements in a compound and the relative proportions of those elements. Water is composed of hydrogen and oxygen in a 2:1 ratio. The chemical formula for water is H_2O . Sulfuric acid is one of the most widely produced chemicals in the United States and is composed of the elements hydrogen, sulfur, and oxygen. The chemical formula for sulfuric acid is H_2SO_4 .

Certain groups of atoms are bonded together to form what is called a polyatomic ion that acts as a single unit. Polyatomic ions are discussed in more detail in [Section 5.5](#). Polyatomic ions are enclosed in parenthesis followed by a subscript if more than one of the same ion exist in a chemical formula. The formula $\text{Ca}_3(\text{PO}_4)_2$ represents a compound with the following:



To count the total number of atoms for formulas with polyatomic ions enclosed in parenthesis, use the subscript as a multiplier for each atom or number of atoms.



Molecular Formula

A **molecular formula** is a chemical formula of a molecular compound that shows the kinds and numbers of atoms present in a molecule of the compound. Ammonia is a compound of nitrogen and hydrogen as shown below:

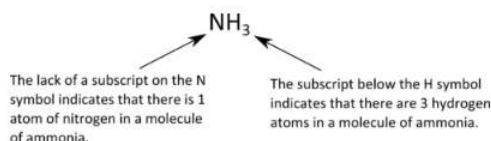


Figure 1.14.1: The molecular formula for ammonia, NH_3 . There is one atom of nitrogen and 3 atoms of hydrogen in a molecule of ammonia.

Note from the example that there are some standard rules to follow in writing molecular formulas. The arrangements of the elements depend on the particular structure, which is not of concern at this point. The number of atoms of each kind is indicated by a subscript following the atom. If there is only one atom, no number is written. If there is more than one atom of a specific kind, the number is written as a subscript following the atom. We would not write N_3H for ammonia, because that would mean that there are three nitrogen atoms and one hydrogen atom in the molecule, which is incorrect.

Empirical Formula

An **empirical formula** is a formula that shows the elements in a compound in their lowest whole-number ratio. Glucose is an important simple sugar that cells use as their primary source of energy. Its molecular formula is $\text{C}_6\text{H}_{12}\text{O}_6$. Since each of the subscripts is divisible by 6, the empirical formula for glucose is CH_2O . When chemists analyze an unknown compound, often the first step is to determine its empirical formula.

- molecular formula: $\text{C}_6\text{H}_{12}\text{O}_6$
- empirical formula: CH_2O

There are a great many compounds whose molecular and empirical formulas are the same. If the molecular formula cannot be simplified into a smaller whole-number ratio, as in the case of H_2O or P_2O_5 , then the empirical formula is also the molecular formula.

Summary

- A chemical formula is an expression that shows the elements in a compound and the relative proportions of those elements.
- If only one atom of a specific type is present, no subscript is used.
- For atoms that have two or more of a specific type of atom present, a subscript is written after the symbol for that atom.
- Polyatomic ions in chemical formulas are enclosed in parentheses followed by a subscript if more than one of the same type of polyatomic ion exist.
- Molecular formulas do not indicate how the atoms are arranged in the molecule.
- The empirical formula tells the lowest whole-number ratio of elements in a compound. The empirical formula does not show the actual number of atoms.

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1.15: Naming Molecular Compounds

Learning Objectives

- Determine the name of a simple molecular compound from its chemical formula.

Molecular Compounds

Molecular compounds are inorganic compounds that take the form of discrete **molecules**. Examples include such familiar substances as water (H_2O) and carbon dioxide (CO_2). These compounds are very different from ionic compounds like sodium chloride (NaCl). Ionic compounds are formed when metal atoms lose one or more of their electrons to nonmetal atoms. The resulting cations and anions are electrostatically attracted to each other.

So what holds the atoms of a molecule together? Rather than forming ions, the atoms of a molecule share their **electrons** in such a way that a **bond** forms between a pair of atoms. In a carbon dioxide molecule, there are two of these bonds, each occurring between the carbon atom and one of the two oxygen atoms.

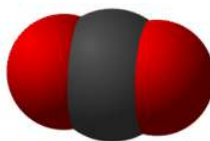


Figure 1.15.1: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Larger molecules can have many, many bonds that serve to keep the molecule together. In a large sample of a given molecular compound, all of the individual molecules are identical.

Naming Binary Molecular Compounds

Recall that a molecular formula shows the number of atoms of each element that a molecule contains. A molecule of water contains two hydrogen atoms and one oxygen atom, so its formula is H_2O . A molecule of octane, which is a component of gasoline, contains 8 atoms of carbon and 18 atoms of hydrogen. The molecular formula of octane is C_8H_{18} .



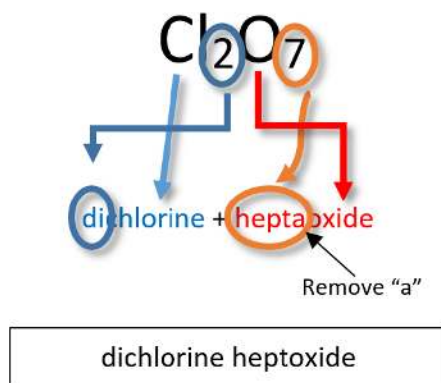
Figure 1.15.2: Nitrogen dioxide (NO_2) is a reddish-brown toxic gas that is a prominent air pollutant produced by internal combustion engines.

Naming *binary* (two-element) molecular compounds is similar to naming simple ionic compounds. The first element in the formula is simply listed using the name of the element. The second element is named by taking the stem of the element name and adding the suffix *-ide*. A system of numerical prefixes is used to specify the number of atoms in a molecule. Table 1.15.1 lists these numerical prefixes.

Table 1.15.1: Numerical Prefixes for Naming Binary Covalent Compounds

Number of Atoms in Compound	Prefix on the Name of the Element
1	mono-*
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

*This prefix is not used for the first element's name.



Note

- Generally, the less electronegative element is written first in the formula, though there are a few exceptions. Carbon is always first in a formula and hydrogen is after nitrogen in a formula such as NH_3 . The order of common nonmetals in binary compound formulas is **C, P, N, H, S, I, Br, Cl, O, F**.
- The *a* or *o* at the end of a prefix is usually dropped from the name when the name of the element begins with a vowel. As an example, four oxygen atoms, is tetraoxide instead of tetraoxide.
- The prefix "mono" is not added to the first element's name if there is only one atom of the first element in a molecule.

Some examples of molecular compounds are listed in Table 1.15.2.

Table 1.15.2

Formula	Name
NO	nitrogen monoxide
N_2O	dinitrogen monoxide
S_2Cl_2	disulfur dichloride
Cl_2O_7	dichlorine heptoxide

Notice that the *mono-* prefix is not used with the nitrogen in the first compound, but is used with the oxygen in both of the first two examples. The S_2Cl_2 emphasizes that the formulas for molecular compounds are not reduced to their lowest ratios. The *o* of the *mono-* and the *a* of *hepta-* are dropped from the name when paired with oxide.

? Exercise 1.15.1

Write the name for each compound.

- a. CF_4
- b. SeCl_2
- c. SO_3

Answer a:

carbon tetrafluoride

Answer b:

selenium dichloride

Answer c:

sulfur trioxide

Simple molecular compounds with common names

For some simple covalent compounds, we use common names rather than systematic names. We have already encountered these compounds, but we list them here explicitly:

- H_2O : water
- NH_3 : ammonia
- CH_4 : methane
- H_2O_2 : hydrogen peroxide

Methane is the simplest organic compound. Organic compounds are compounds with carbon atoms and are named by a separate nomenclature system.

Some Compounds Have Both Covalent and Ionic Bonds

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

Summary

- A molecular compound is usually composed of two or more nonmetal elements.
- Molecular compounds are named with the first element first and then the second element by using the stem of the element name plus the suffix *-ide*. Numerical prefixes are used to specify the number of atoms in a molecule.

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1.16: Structure and Composition of the Atmosphere

Life as we know it on the Earth is entirely dependent on the tenuous layer of gas that clings to the surface of the globe, adding about 1% to its diameter and an insignificant amount to its total mass. And yet the atmosphere serves as the earth's window and protective shield, as a medium for the transport of heat and water, and as source and sink for exchange of carbon, oxygen, and nitrogen with the biosphere. The atmosphere acts as a compressible fluid tied to the earth by gravitation; as a receptor of solar energy and a thermal reservoir, it constitutes the working fluid of a heat engine that transports and redistributes matter and energy over the entire globe. The atmosphere is also a major temporary repository of a number of chemical elements that move in a cyclic manner between the hydrosphere, atmosphere, and the upper lithosphere. Finally, the atmosphere is a site for a large variety of complex photochemically initiated reactions involving both natural and anthropogenic substances.

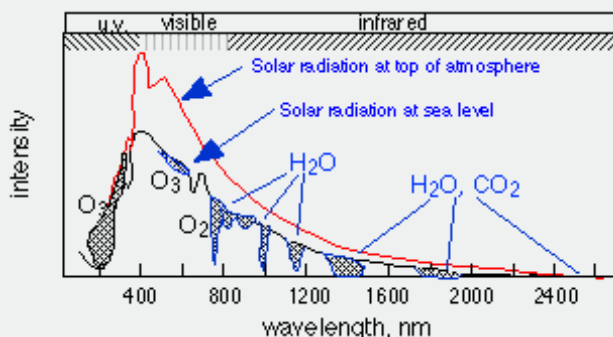
On the scale of cubic meters the air is a homogeneous mixture of its constituent gases, but on a larger scale the atmosphere is anything but uniform. Variations of temperature, pressure, and moisture content in the layers of air near the earth's surface give rise to the dynamic effects we know as the weather.

Although the density of the atmosphere decreases without limit with increasing height, for most practical purposes one can roughly place its upper boundary at about 500 km. However, half the mass of the atmosphere lies within 5 km, and 99.99% within 80 km of the surface. The average atmospheric pressure at sea level is 1.01×10^5 pascals, or 1010 millibars. A 1-cm² cross section of the earth's surface supports a column weighing 1030 g; the total mass of the atmosphere is about 5.27×10^{21} g.

About 80% of the mass of the atmosphere resides in the first 10 km; this well-mixed region of fairly uniform composition is known as the *troposphere*.

Solar irradiation of the Earth

The gases ozone, water vapor, and carbon dioxide are only minor components of the atmosphere, but they exert a huge effect on the Earth by absorbing radiation in the ranges indicated by the shading. Ozone in the upper atmosphere filters out the ultraviolet light below about 360 nm that is destructive of life. O₂, H₂O, CO₂ and CH₄ are "greenhouse" gases that trap some of the heat absorbed from the Sun and prevent it from re-radiating back into space.

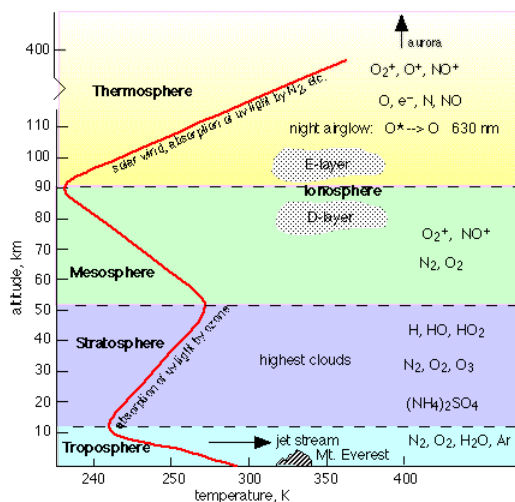


Structure of the atmosphere

We commonly think of gas molecules as moving about in a completely random manner, but the Earth's gravitational field causes downward motions to be very slightly favored so that the molecules in any thin layer of the air collide more frequently with those in the layer below. This gives rise to a pressure gradient that is the most predictable and well known structural characteristic of the atmosphere. This gradient is described by an exponential law which predicts that the atmospheric pressure should decrease by 50% for every 6 km increase in altitude. This law also predicts that the composition of a gas mixture will change with altitude, the lower-molecular weight components being increasingly favored at higher altitudes. However, this gravitational fractionation effect is completely obliterated below about 160 km owing to turbulence and convective flows (winds).

The atmosphere is divided vertically into several major regions which are distinguished by the sign of the temperature gradient. In the lowermost region, the troposphere, the temperature falls with increasing altitude. The major source of heat input into this part of the atmosphere is long-wave radiation from the earth's surface, while the major loss is radiation into space.

At higher elevations the temperature begins to rise with altitude as we move into a region in which heat is produced by exothermic chemical reactions, mainly the decomposition of ozone that is formed photochemically from dioxygen in the stratosphere. At still higher elevations the ozone gives out and the temperature begins to drop; this is the mesosphere, which is finally replaced by the thermosphere which consists largely of a plasma (gaseous ions). This outer section of the atmosphere which extends indefinitely to perhaps 2000 km is heated by absorption of intense u.v. radiation from the Sun and also from the solar wind, a continual rain of electrons, protons, and other particles emitted from the Sun's surface.



Structure of the atmosphere. The main divisions of the atmosphere are defined by the elevations at which the sign of the temperature gradient changes. The chemical formulas at the right show the major species of interest in the various regions. The shaded D- and E-layers are regions of high ion concentrations that reflect radio waves and are important in long-distance communication.

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. 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 Ar^{40} which is a decay product of K^{40} in the mantle and crust.

Table 1: Major components

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, about which more will be said below. 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 present 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 (a small amount is produced by biological action), 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 2: Minor components

water	H ₂ O	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 ₂ O	0.3 ppm
carbon monoxide	CO	0.05-0.2 ppm
ozone	O ₃	0.02 - 10 ppm
xenon	Xe	0.08 ppm

Table 3: Trace components

ammonia	NH ₃	4 ppb
nitrogen oxide	NO	1 ppb
sulfur dioxide	SO ₂	1 ppb
hydrogen sulfide	H ₂ S	0.05 ppb

Ozone

Ozone is formed by the reaction of O₂ with oxygen atoms produced photochemically. As a consequence the abundance of ozone varies with the time of day, the concentration of O atoms from other sources (photochemical smog, for example), and particularly with altitude; at 30 km, the ozone concentration reaches a maximum of 12 ppm.

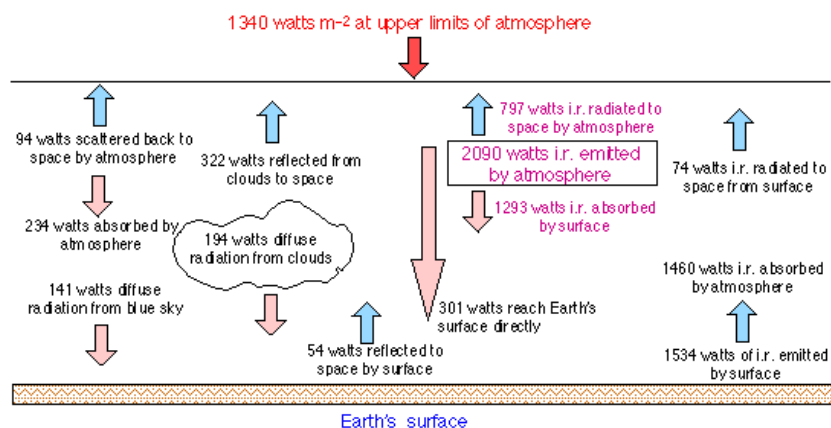
Carbon dioxide

The concentration of atmospheric carbon dioxide, while fairly uniform globally, is increasing at a rate of 0.2-0.7% per year as a result of fossil fuel burning. The present CO₂ content of the atmosphere is about 129  10¹⁸ g. Most of the CO₂, however, is of natural origin, and represents the smallest part of the total carbonate reservoir that includes oceanic CO₂, HCO₃⁻, and carbonate sediments. The latter contain about 600 times as much CO₂ as the atmosphere, and the oceans contain about 50 times as much. These relative amounts are controlled by the rates of the reactions that interconvert the various forms of carbonate.

The surface conditions on the earth are sensitively dependent on the atmospheric CO₂ concentration. This is due mainly to the strong infrared absorption of CO₂, which promotes the absorption and trapping of solar heat (see below). Since CO₂ acts as an acid in aqueous solution, the pH of the oceans is also dependent on the concentration of CO₂ in the atmosphere; it has been estimated that if only 1% of the carbonate presently in sediments were still in the atmosphere, the pH of the oceans would be 5.9, instead of the present 8.2.

Energy balance of the atmosphere and earth

The amount of energy (the solar flux) impinging on the outer part of the atmosphere is 1367 watts m⁻². About 30% of this is reflected or scattered back into space by clouds, dust, and the atmospheric gas molecules themselves, and by the earth's surface. About 19% of the radiation is absorbed by clouds or the atmosphere (mainly by H₂O and O₃, but not CO₂), leaving 51% of the incident energy available for absorption by the earth's surface. If one takes into account the uneven illumination of the earth's surface and the small flux of internal heat to the surface, the assumption of thermal equilibrium requires that the earth emit about 240 watts m⁻². This corresponds to the power that would be emitted by a black body at 255 K, or -18°C, which is the average temperature of the atmosphere at an altitude of 5 km. The observed mean global surface temperature of the earth is 13°C, and is presumably the temperature required to maintain thermal equilibrium between the earth and the atmosphere.



The greenhouse effect

The energy radiated by the earth has a longer wavelength (maximum 12μm) than the incident radiation. Most gases absorb radiation in this range quite efficiently, including those gases such as CO₂ and N₂O that do not absorb the incident radiation. The energy absorbed by atmospheric gases is re-radiated in all directions; some of it therefore escapes into space, but a portion returns to the earth and is reabsorbed, thus raising its temperature. This is commonly called the greenhouse effect. If the amount of an infrared-absorbing gas such as carbon dioxide increases, a larger fraction of the incident solar radiation is trapped, and the mean temperature of the earth will increase.

Any significant increase in the temperature of the oceans would increase the atmospheric concentrations of both water and CO₂, producing the possibility of a runaway process that would be catastrophic from a human perspective. Fossil fuel combustion and deforestation during the last two hundred years have increased the atmospheric CO₂ concentration by 25%, and this increase is continuing. The same combustion processes responsible for the increasing atmospheric CO₂ concentration also introduce considerable quantities of particulate materials into the upper atmosphere. The effect of these would be to scatter more of the incoming solar radiation, reducing the amount that reaches and heats the earth's surface. The extent to which this process counteracts the greenhouse effect is still a matter of controversy; all that is known for sure is that the average temperature of the Earth is increasing.

Nitrous oxide

Carbon dioxide is not the only atmospheric gas of anthropogenic origin that can affect the heat balance of the earth; other examples are SO₂ and N₂O. Nitrous oxide is of particular interest, since its abundance is fairly high, and is increasing at a rate of about 0.5% per year. It is produced mainly by bacteria, and much of the increase is probably connected with introduction of increased nitrate into the environment through agricultural fertilization and sewage disposal. Besides being a strong infrared absorber, N₂O is

photochemically active, and can react with ozone. Any significant depletion of the ozone content of the upper atmosphere would permit more ultraviolet radiation to reach the earth. This would have numerous deleterious effects on present life forms, as well as contributing to a temperature increase. The warming effect attributed to anthropogenic additions of greenhouse gases to the atmosphere is estimated to be about 2 watts per m^2 , or about 1.5% of the 150 watts per m^2 trapped by clouds and atmospheric gases. This is a relatively large perturbation compared to the maximum variation in solar output of 0.5 watts per m^2 that has been observed during the past century. Continuation of greenhouse gas emission at present levels for another century could increase the atmospheric warming effect by 6-8 watts per m^2 .

A less-appreciated side effect of the increase in atmospheric carbon dioxide (and of other plant nutrients such as nitrates) may be reduction in plant species diversity by selectively encouraging the growth of species which are ordinarily held in check by other species that are able to grow well with fewer nutrients. This effect, for which there is already some evidence, could be especially pronounced when the competing species utilize the C_3 and C_4 photosynthetic pathways that differ in their sensitivity to CO_2 .

Contributors and Attributions

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1.17: Conservation of Mass - There is No New Matter

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Look at Figure 1.17.1 below. It shows that when wood burns, it combines with oxygen and changes not only to ashes, but also to carbon dioxide and water vapor. The gases float off into the air, leaving behind just the ashes. Suppose you had measured the mass of the wood before it burned and the mass of the ashes after it burned. Also suppose you had been able to measure the oxygen used by the fire and the gases produced by the fire. What would you find? The total mass of matter after the fire would be the same as the total mass of matter before the fire.

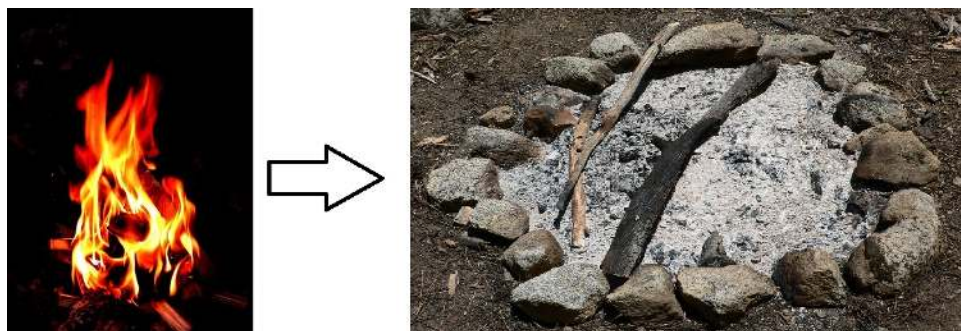


Figure 1.17.1: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). (CC BY-SA 2.5; Einar Helland Berger for fire and Walter Siegmund for ash).

Law of Conservation of Mass

The law of conservation of mass was created in 1789 by a French chemist, Antoine Lavoisier. The **law of conservation of mass** states that matter cannot be created or destroyed in a chemical reaction. For example, when wood burns, the mass of the soot, ashes, and gases equals the original mass of the charcoal and the oxygen when it first reacted. So the mass of the product equals the mass of the reactant. A reactant is the chemical reaction of two or more elements to make a new substance, and a product is the substance that is formed as the result of a chemical reaction (Video 1.17.1). Matter and its corresponding mass may not be able to be created or destroyed, but can change forms to other substances like liquids, gases, and solids.



Video 1.17.1: This is a nice little demonstration showing the Conservation of Mass in action.

If you witness a 300 kg tree burn to the ground, there are only ashes left after the burn, and all of them together weigh 10 kg. It may make you wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke, so the only thing left that you can see is the 10 kg of ash. If you know the law of conservation of mass, then you know that the other 290 kg has to go somewhere, because it has to equal the mass of the tree before it burnt down.

✓ Example 1.17.1

If heating 10.0 grams of calcium carbonate (CaCO_3) produces 4.4 g of carbon dioxide (CO_2) and 5.6 g of calcium oxide (CaO), show that these observations are in agreement with the law of conservation of mass.

Solution

Mass of the reactants = Mass of the products

10.0 g of CaCO_3 = 4.4 g of CO_2 + 5.6 g of CaO

10.0 g of reactant = 10.0 g of products

Because the mass of the reactant is equal to the mass of the products, the observations are in agreement with the law of conservation of mass.

? Exercise 1.17.1

Potassium hydroxide (KOH) readily reacts with carbon dioxide (CO_2) to produce potassium carbonate (K_2CO_3) and water (H_2O). How many grams of potassium carbonate are produced if 224.4 g of KOH reacts with 88.0 g of CO_2 ? The reaction also produces 36.0 g of water.

Answer

276.4 g of potassium carbonate

The Law is also applicable to both chemical and physical changes. For example, if you have an ice cube that melts into a liquid and you heat that liquid up, it becomes a gas. It will appear to have disappeared, but is still there.

Summary

- Burning and other changes in matter do not destroy matter.
- The mass of matter is always the same before and after the changes occur.
- The law of conservation of mass states that matter cannot be created or destroyed.

Contributions & Attributions

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1.18: Evidence of a Chemical Reaction

Learning Objectives

- Identify the evidence for chemical reactions.

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

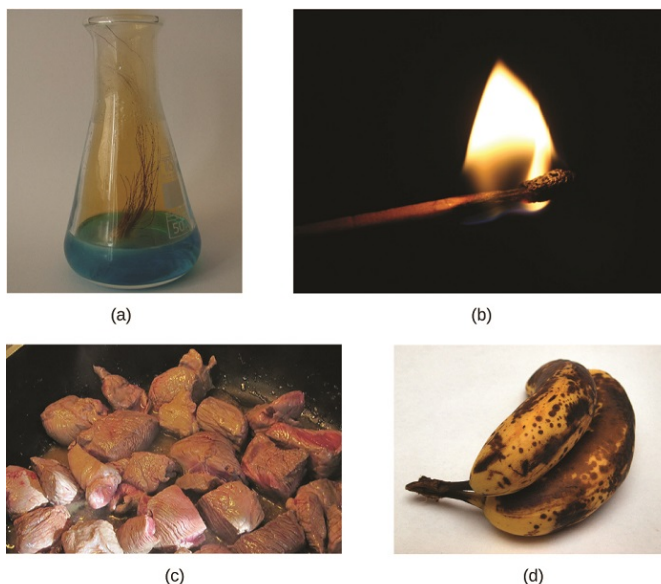


Figure 1.18.1: (a) Copper and nitric acid undergo a chemical change to form copper nitrate and brown, gaseous nitrogen dioxide. (b) During the combustion of a match, cellulose in the match and oxygen from the air undergo a chemical change to form carbon dioxide and water vapor. (c) Cooking red meat causes a number of chemical changes, including the oxidation of iron in myoglobin that results in the familiar red-to-brown color change. (d) A banana turning brown is a chemical change as new, darker (and less tasty) substances form. (Credit b: modification of work by Jeff Turner; credit c: modification of work by Gloria Cabada-Leman; credit d: modification of work by Roberto Verzo.)

To identify a chemical reaction, we look for a **chemical change**. A chemical change always produces one or more types of matter that differ from the matter present before the change. The formation of rust is a chemical change because rust is a different kind of matter than the iron, oxygen, and water present before the rust formed. The explosion of nitroglycerin is a chemical change because the gases produced are very different kinds of matter from the original substance. Other examples of chemical changes include: reactions that are performed in a lab (such as copper reacting with nitric acid), all forms of combustion (burning), and food being cooked, digested, or rotting (Figure 1.18.1).



Video1.18.1: Evidence of a Chemical Reaction

✓ **Example 1.18.1: Evidence of a Chemical Reaction**

Which of the following is a chemical reaction?

- a. Freezing liquid mercury.
- b. Adding yellow to blue to make green.
- c. Cutting a piece of paper into two pieces.
- d. Dropping a sliced orange into a vat of sodium hydroxide.
- e. Filling a balloon with natural air.

Solution

A, B, C, & E involve only physical changes. A sliced orange has acid (citric acid) that can react with sodium hydroxide, so the answer is D.

? **Exercise 1.18.1**

Which of the following is a chemical reaction?

- a. Painting a wall blue.
- b. A bicycle rusting.
- c. Ice cream melting.
- d. Scratching a key across a desk.
- e. Making a sand castle.

Answer

B

✓ **Example 1.18.2: Evidence of a Chemical Reaction**

Which of the following is not a chemical reaction?

- a. Shattering glass with a baseball.
- b. Corroding metal.
- c. Fireworks exploding.
- d. Lighting a match.
- e. Baking a cake.

Solution

Shattering glass with a baseball results in glass broken into many pieces but no chemical change happens, so the answer is A.

? Exercise 1.18.2

Which of the following is NOT a chemical reaction?

- a. Frying an egg.
- b. Slicing carrots.
- c. A Macbook falling out of a window.
- d. Creating ATP in the human body.
- e. Dropping a fizzy tablet into a glass of water.

Answer

B and C

Summary

Chemical reactions can be identified via a wide range of different observable factors including change in color, energy change (temperature change or light produced), gas production, formation of precipitate and change in properties.

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1.19: Chemical Equations

Learning Objectives

- Identify the reactants and products in any chemical reaction.
- Convert word equations into chemical equations.
- Use the common symbols, *(s)*, *(l)*, *(g)*, *(aq)*, and $\rightarrow$ appropriately when writing a chemical reaction.

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

Reactants and Products

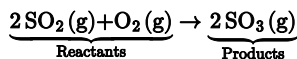
To describe a chemical reaction, we need to indicate what substances are present at the beginning and what substances are present at the end. The substances that are present at the beginning are called **reactants** and the substances present at the end are called **products**.

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

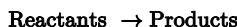


Writing Chemical Equations

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



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



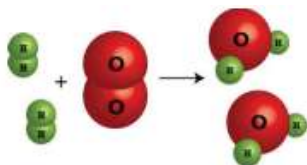
There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. Table 1.19.1 shows a listing of symbols used in chemical equations.

Table 1.19.1: Symbols Used in Chemical Equations

Symbol	Description	Symbol	Description
+	used to separate multiple reactants or products	(s)	reactant or product in the solid state
→	yield sign; separates reactants from products	(l)	reactant or product in the liquid state
⇌	replaces the yield sign for reversible reactions that reach equilibrium	(g)	reactant or product in the gas state
$\xrightarrow{\text{Pt}}$	formula written above the arrow is used as a catalyst in the reaction	(aq)	reactant or product in an aqueous solution (dissolved in water)
$\xrightarrow{\Delta}$	triangle indicates that the reaction is being heated		

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

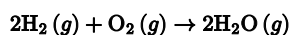
1. They could draw a picture of the chemical reaction.



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

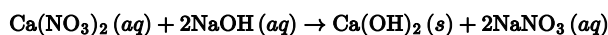
"Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor."

3. They could write the equation in chemical shorthand.

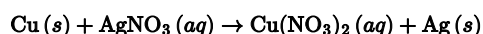


In the symbolic equation, chemical formulas are used instead of chemical names for reactants and products, while symbols are used to indicate the phase of each substance. It should be apparent that the chemical shorthand method is the quickest and clearest method for writing chemical equations.

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



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



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

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

1. Identify the reactants and products. This will help you know which symbols go on each side of the arrow and where the + signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 5 (including making all ionic compounds charge balanced).

3. Write the correct formulas for all elements. Usually this is given straight off of the periodic table. However, there are seven elements that are considered diatomic, meaning that they are always found in pairs in nature. They include those elements listed in the table.

Table 1.19.1: Diatomic Elements

Element Name	Hydrogen	Nitrogen	Oxygen	Fluorine	Chlorine	Bromine	Iodine
Formula	H_2	N_2	O_2	F_2	Cl_2	Br_2	I_2

✓ Example 1.19.1

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

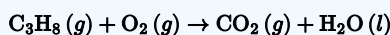
- $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
- Gaseous propane, C_3H_8 , burns in oxygen gas to produce gaseous carbon dioxide and liquid water.
- Hydrogen fluoride gas reacts with an aqueous solution of potassium carbonate to produce an aqueous solution of potassium fluoride, liquid water, and gaseous carbon dioxide.

Solution

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

b. Reactants: propane (C_3H_8) and oxygen (O_2)

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



c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

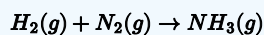


? Exercise 1.19.1

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

- Hydrogen gas reacts with nitrogen gas to produce gaseous ammonia.
- $HCl(aq) + LiOH(aq) \rightarrow LiCl(aq) + H_2O(l)$
- Copper metal is heated with oxygen gas to produce solid copper(II) oxide.

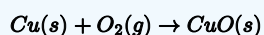
Answer a



Answer b

An aqueous solution of hydrochloric acid reacts with an aqueous solution of lithium hydroxide to produce an aqueous solution of lithium chloride and liquid water.

Answer c



Summary

- A chemical reaction is the process by which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.

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1.20: How to Write Balanced Chemical Equations

Learning Objectives

- Explain the roles of subscripts and coefficients in chemical equations.
- Balance a chemical equation when given the unbalanced equation.
- Explain the role of the Law of Conservation of Mass in a chemical reaction.

Even though chemical compounds are broken up and new compounds are formed during a chemical reaction, atoms in the reactants do not disappear, nor do new atoms appear to form the products. In chemical reactions, atoms are never created or destroyed. The same atoms that were present in the reactants are present in the products—they are merely reorganized into different arrangements. In a complete chemical equation, the two sides of the equation must be present on the reactant and the product sides of the equation.

Coefficients and Subscripts

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

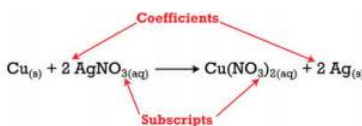
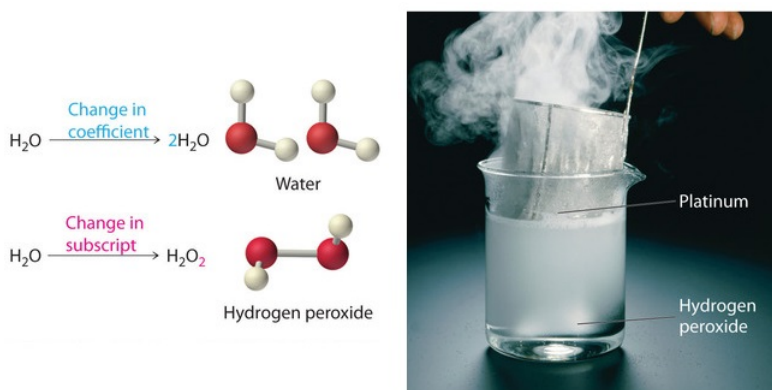


Figure 1.20.1: Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (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.

The **subscripts** are part of the formulas and once the formulas for the reactants and products are determined, the subscripts may not be changed. The **coefficients** indicate the number of each substance involved in the reaction and may be changed in order to balance the equation. The equation above indicates that one mole of solid copper is reacting with two moles of aqueous silver nitrate to produce one mole of aqueous copper (II) nitrate and two atoms of solid silver.

Balancing a Chemical 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 1.20.1.



Original molecule H_2O : if the coefficient 2 is added in front, that makes 2 water molecules; but if the subscript 2 is added to make H_2O_2 , that's hydrogen peroxide.

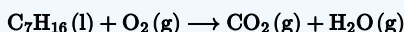
The simplest and most generally useful method for balancing chemical equations is “inspection,” better known as trial and error. The following is an efficient approach to balancing a chemical equation using this method.

Steps in Balancing a Chemical Equation

1. Identify the most complex substance.
2. Beginning with that substance, choose an element(s) that appears in only one reactant and one product, if possible. Adjust the coefficients to obtain the same number of atoms of this element(s) on both sides.
3. Balance polyatomic ions (if present on both sides of the chemical equation) 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. Count the numbers of atoms of each kind on both sides of the equation to be sure that the chemical equation is balanced.

✓ Example 1.20.1: Combustion of Heptane

Balance the chemical equation for the combustion of Heptane (C_7H_{16}).



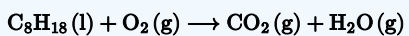
Solution

Solutions to Example 1.20.1

Steps	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.	a. 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: $\text{C}_7\text{H}_{16}(\text{l}) + \text{O}_2(\text{g}) \rightarrow 7\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$ • 7 carbon atoms on both reactant and product sides b. 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: $\text{C}_7\text{H}_{16}(\text{l}) + \text{O}_2(\text{g}) \rightarrow 7\text{CO}_2(\text{g}) + 8\text{H}_2\text{O}(\text{g})$ • 16 hydrogen atoms on both reactant and product sides
3. Balance polyatomic ions as a unit.	There are no polyatomic ions to be considered in this reaction.
4. Balance the remaining atoms.	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: $\text{C}_7\text{H}_{16}(\text{l}) + 11\text{O}_2(\text{g}) \rightarrow 7\text{CO}_2(\text{g}) + 8\text{H}_2\text{O}(\text{g})$ • 22 oxygen atoms on both reactant and product sides
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.

✓ Example 1.20.2: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

**Solution**

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. The combustion of any hydrocarbon with oxygen produces carbon dioxide and water.

Solutions to Example 7.4.2

Steps	Example
1. Identify the most complex substance.	The most complex substance is the one with the largest number of different atoms, which is C_8H_{18} . We will assume initially that the final balanced chemical equation contains 1 molecule or formula unit of this substance.
2. Adjust the coefficients.	a. 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: $\text{C}_8\text{H}_{18}(\text{l}) + \text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{g})$ 8 carbon atoms on both reactant and product sides b. 18 hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $\text{C}_8\text{H}_{18}(\text{l}) + \text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 9\text{H}_2\text{O}(\text{g})$ 18 hydrogen atoms on both reactant and product sides
3. Balance polyatomic ions as a unit.	There are no polyatomic ions to be considered in this reaction.
4. Balance the remaining atoms.	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, O_2, but because there are 2 oxygen atoms per O_2 molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $\text{C}_8\text{H}_{18}(\text{l}) + \frac{25}{2}\text{O}_2(\text{g}) \rightarrow 8\text{CO}_2(\text{g}) + 9\text{H}_2\text{O}(\text{g})$ 25 oxygen atoms on both reactant and product sides The equation 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: $2\text{C}_8\text{H}_{18}(\text{l}) + 25\text{O}_2(\text{g}) \rightarrow 16\text{CO}_2(\text{g}) + 18\text{H}_2\text{O}(\text{g})$
5. Check your work.	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side. 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 1.20.3: Precipitation of Lead (II) Chloride

Aqueous solutions of lead (II) nitrate and sodium chloride are mixed. The products of the reaction are an aqueous solution of sodium nitrate and a solid precipitate of lead (II) chloride. Write the balanced chemical equation for this reaction.

Solution

Solutions to Example 7.4.3

Steps	Example
1. Identify the most complex substance.	The most complex substance is lead (II) chloride. $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \text{NaCl}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{PbCl}_2(\text{s})$
2. Adjust the coefficients.	There are twice as many chloride ions in the reactants as there are in the products. Place a 2 in front of the NaCl in order to balance the chloride ions. $\text{Pb}(\text{NO}_3)_2(\text{aq}) + \underline{2} \text{NaCl}(\text{aq}) \rightarrow \text{NaNO}_3(\text{aq}) + \text{PbCl}_2(\text{s})$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on reactant side, 1 Na atom on product side • 2 Cl atoms on both reactant and product sides
3. Balance polyatomic ions as a unit.	The nitrate ions are still unbalanced. Place a 2 in front of the NaNO₃. The result is: $\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2 \text{NaCl}(\text{aq}) \rightarrow \underline{2} \text{NaNO}_3(\text{aq}) + \text{PbCl}_2(\text{s})$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO₃⁻ atoms on both reactant and product sides
4. Balance the remaining atoms.	There is no need to balance the remaining atoms because they are already balanced.
5. Check your work.	$\text{Pb}(\text{NO}_3)_2(\text{aq}) + 2 \text{NaCl}(\text{aq}) \rightarrow 2 \text{NaNO}_3(\text{aq}) + \text{PbCl}_2(\text{s})$ • 1 Pb atom on both reactant and product sides • 2 Na atoms on both reactant and product sides • 2 Cl atoms on both reactant and product sides • 2 NO₃⁻ atoms on both reactant and product sides

? Exercise 1.20.1

Is each chemical equation balanced?

- $2\text{Hg}(\ell) + \text{O}_2(\text{g}) \rightarrow \text{Hg}_2\text{O}_2(\text{s})$
- $\text{C}_2\text{H}_4(\text{g}) + 2\text{O}_2(\text{g}) \rightarrow 2\text{CO}_2(\text{g}) + 2\text{H}_2\text{O}(\text{g})$
- $\text{Mg}(\text{NO}_3)_2(\text{s}) + 2\text{Li}(\text{s}) \rightarrow \text{Mg}(\text{s}) + 2\text{LiNO}_3(\text{s})$

Answer a

yes

Answer b

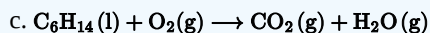
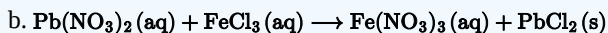
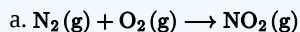
no

Answer c

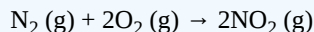
yes

? Exercise 1.20.2

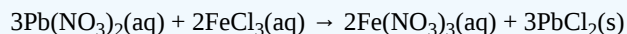
Balance the following chemical equations.



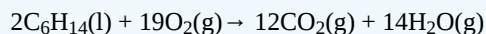
Answer a



Answer b



Answer c



Summary

- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Vocabulary

- **Chemical reaction** - The process in which one or more substances are changed into one or more new substances.
- **Reactants** - The starting materials in a reaction.
- **Products** - Materials present at the end of a reaction.
- **Balanced chemical equation** - A chemical equation in which the number of each type of atom is equal on the two sides of the equation.
- **Subscripts** - Part of the chemical formulas of the reactants and products that indicate the number of atoms of the preceding element.
- **Coefficient** - A small whole number that appears in front of a formula in a balanced chemical equation.

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CHAPTER OVERVIEW

A: Preliminary

- 2.1: The Scope of Chemistry
- 2.2: Hypothesis, Theories, and Laws
- 2.3: Scientific Notation - Writing Large and Small Numbers
- 2.4: Significant Figures - Writing Numbers to Reflect Precision
- 2.5: Significant Figures in Calculations
- 2.6: The Basic Units of Measurement

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2.1: The Scope of Chemistry

Learning Objectives

- To recognize the breadth, depth, and scope of chemistry.
- Define chemistry in relation to other sciences.
- Identify the main disciplines of chemistry.

Chemistry is the study of matter—what it consists of, what its properties are, and how it changes. Matter is anything that has mass and takes up space—that is, anything that is physically real. Some things are easily identified as matter—the screen on which you are reading this book, for example. Others are not so obvious. Because we move so easily through air, we sometimes forget that it, too, is matter. Because of this, chemistry is a science that has its fingers in just about everything. Being able to describe the ingredients in a cake and how they change when the cake is baked, for example, is chemistry!

Chemistry is one branch of science. Science is the process by which we learn about the natural universe by observing, testing, and then generating models that explain our observations. Because the physical universe is so vast, there are many different branches of science (Figure 2.1.1). Thus, chemistry is the study of matter, biology is the study of living things, and geology is the study of rocks and the earth. Mathematics is the language of science, and we will use it to communicate some of the ideas of chemistry.

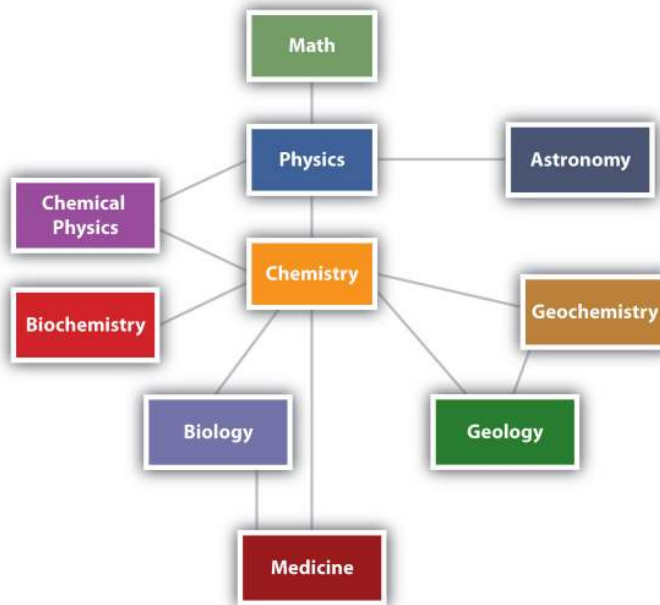


Figure 2.1.1: The Relationships between Some of the Major Branches of Science. Chemistry lies more or less in the middle, which emphasizes its importance to many branches of science.

Although we divide science into different fields, there is much overlap among them. For example, some biologists and chemists work in both fields so much that their work is called biochemistry. Similarly, geology and chemistry overlap in the field called geochemistry. Figure 2.1.1 shows how many of the individual fields of science are related. At some level, all of these fields depend on matter because they all involve "stuff"; because of this, chemistry has been called the "central science", linking them all together.

There are many other fields of science, in addition to the ones (biology, medicine, etc.) listed here.

✓ Example 2.1.1: Science Fields

Which fields of study are branches of science? Explain.

- a. sculpture
- b. astronomy

Solution

- a. Sculpture is not considered a science because it is not a study of some aspect of the natural universe.
- b. Astronomy is the study of stars and planets, which are part of the natural universe. Astronomy is therefore a field of science.

? Exercise 2.1.1

Which fields of study are branches of science?

- a. physiology (the study of the function of an animal's or a plant's body)
- b. geophysics
- c. agriculture
- d. politics

Answer a:

yes

Answer b:

yes

Answer c:

yes

Answer d:

no

Areas of Chemistry

The study of modern chemistry has many branches, but can generally be broken down into five main disciplines, or areas of study:

- **Physical chemistry:** Physical chemistry is the study of macroscopic properties, atomic properties, and phenomena in chemical systems. A physical chemist may study such things as the rates of chemical reactions, the energy transfers that occur in reactions, or the physical structure of materials at the molecular level.
- **Organic chemistry:** Organic chemistry is the study of chemicals containing carbon. Carbon is one of the most abundant elements on Earth and is capable of forming a tremendously vast number of chemicals (over twenty million so far). Most of the chemicals found in all living organisms are based on carbon.
- **Inorganic chemistry:** Inorganic chemistry is the study of chemicals that, in general, are not primarily based on carbon. Inorganic chemicals are commonly found in rocks and minerals. One current important area of inorganic chemistry deals with the design and properties of materials involved in energy and information technology.
- **Analytical chemistry:** Analytical chemistry is the study of the composition of matter. It focuses on separating, identifying, and quantifying chemicals in samples of matter. An analytical chemist may use complex instruments to analyze an unknown material in order to determine its various components.
- **Biochemistry:** Biochemistry is the study of chemical processes that occur in living things. Research may cover anything from basic cellular processes up to understanding disease states so that better treatments can be developed.



Figure 2.1.2: (left) Measurement of trace metals using atomic spectroscopy. (right) Measurement of hormone concentrations.

In practice, chemical research is often not limited to just one of the five major disciplines. A particular chemist may use biochemistry to isolate a particular chemical found in the human body such as hemoglobin, the oxygen carrying component of red blood cells. He or she may then proceed to analyze the hemoglobin using methods that would pertain to the areas of physical or analytical chemistry. Many chemists specialize in areas that are combinations of the main disciplines, such as bioinorganic chemistry or physical organic chemistry.

History of Chemistry

The **history** of chemistry is an interesting and challenging one. Very early chemists were often motivated mainly by the achievement of a specific goal or product. Making perfume or soaps did not need a lot of **theory**, just a good recipe and careful attention to detail. There was no standard way of naming materials (and no periodic table that we could all agree on). It is often difficult to figure out exactly what a particular person was using. However, the science developed over the centuries by trial and error.

Major progress was made toward putting chemistry on a solid foundation when Robert Boyle (1637-1691) began his **research** in chemistry (Figure 2.1.3). He developed the basic ideas about the behavior of gases. He could then describe gases mathematically. Boyle also helped form the idea that small particles could combine to form molecules. Many years later, John Dalton used these ideas to develop the atomic theory.



Figure 2.1.3: Robert Boyle.

The field of chemistry began to develop rapidly in the 1700's. Joseph Priestley (1733-1804) isolated and characterized several gases: oxygen, carbon monoxide, and nitrous oxide. It was later discovered that nitrous oxide ("laughing gas") worked as an anesthetic. This gas was used for that purpose for the first time in 1844 during a tooth extraction. Other gases discovered during that time were chlorine, by C.W. Scheele (1742-1786) and nitrogen, by Antoine Lavoisier (1743-1794). Lavoisier has been considered by many scholars to be the "father of chemistry". Among other accomplishments, he discovered the role of oxygen in combustion and definitively formulated the law of conservation of matter.

Chemists continued to discover new compounds in the 1800's. The science also began to develop a more theoretical foundation. John Dalton (1766-1844) put forth his atomic theory in 1807. This idea allowed scientists to think about chemistry in a much more systematic way. Amadeo Avogadro (1776-1856) laid the groundwork for a more quantitative approach to chemistry by calculating the number of particles in a given amount of a gas. A lot of effort was put forth in studying chemical reactions. These efforts led to new materials being produced. Following the invention of the battery by Alessandro Volta (1745-1827), the field of electrochemistry (both theoretical and applications) developed through major contributions by Humphry Davy (1778-1829) and Michael Faraday (1791-1867). Other areas of the discipline also progressed rapidly.

It would take a large book to cover developments in chemistry during the twentieth century and up to today. One major area of expansion was in the area of the chemistry of living processes. Research in photosynthesis in plants, the discovery and

characterization of enzymes as biochemical catalysts, elucidation of the structures of biomolecules such as insulin and DNA—these efforts gave rise to an explosion of information in the field of biochemistry.

The practical aspects of chemistry were not ignored. The work of Volta, Davy, and Faraday eventually led to the development of batteries that provided a source of electricity to power a number of devices (Figure 2.1.4).

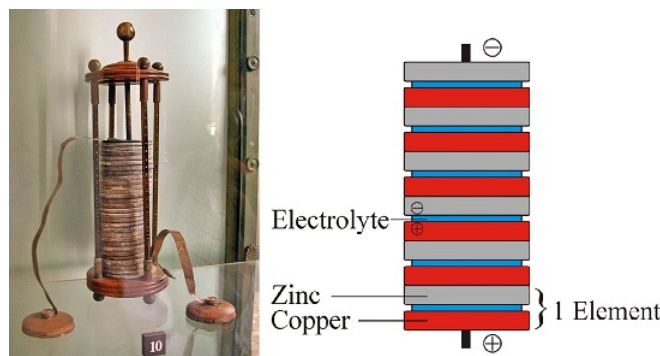


Figure 2.1.4: Battery developed by Volta. (CC BY-SA 3.0; (left) GuidoB and (right) Kkkdc).

Charles Goodyear (1800-1860) discovered the process of vulcanization, allowing a stable rubber product to be produced for the tires of all the vehicles we have today. Louis Pasteur (1822-1895) pioneered the use of heat sterilization to eliminate unwanted microorganisms in wine and milk. Alfred Nobel (1833-1896) invented dynamite (Figure 2.1.5). After his death, the fortune he made from this product was used to fund the Nobel Prizes in science and the humanities. J.W. Hyatt (1837-1920) developed the first plastic. Leo Baekeland (1863-1944) developed the first synthetic resin, widely used for inexpensive and sturdy dinnerware.



Figure 2.1.5: Dynamite explosion in Panama, Central America (1908).

Today, chemistry continues to be essential to the development of new materials and technologies, from semiconductors for electronics to powerful new medicines, and beyond.

Summary

- Chemistry is the study of matter and the changes it undergoes and considers both *macroscopic* and *microscopic* information.
- Matter is anything that has mass and occupies space.
- The five main disciplines of chemistry are physical chemistry, organic chemistry, inorganic chemistry, analytical chemistry and biochemistry.
- Many civilizations contributed to the growth of chemistry. A lot of early chemical research focused on practical uses. Basic chemistry theories were developed during the nineteenth century. New materials and batteries are a few of the products of modern chemistry.

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2.2: Hypothesis, Theories, and Laws

Learning Objectives

- Describe the difference between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.

Although many have taken science classes throughout the course of their studies, people often have incorrect or misleading ideas about some of the most important and basic principles in science. Most students have heard of hypotheses, theories, and laws, but what do these terms really mean? Prior to reading this section, consider what you have learned about these terms before. What do these terms mean to you? What do you read that contradicts or supports what you thought?

What is a Fact?

A fact is a basic statement established by experiment or observation. All facts are true under the specific conditions of the observation.

What is a Hypothesis?

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

- An educated guess: a scientific hypothesis provides a suggested solution based on evidence.
- Prediction: if you have ever carried out a science experiment, you probably made this type of hypothesis when you predicted the outcome of your experiment.
- Tentative or proposed explanation: hypotheses can be suggestions about why something is observed. In order for it to be scientific, however, a scientist must be able to test the explanation to see if it works and if it is able to correctly predict what will happen in a situation. For example, "if my hypothesis is correct, we should see ___ result when we perform ___ test."

A hypothesis is very tentative; it can be easily changed.

What is a Theory?

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

"Some scientific explanations are so well established that no new evidence is likely to alter them. The explanation becomes a scientific theory. In everyday language a theory means a hunch or speculation. Not so in science. In science, the word **theory** refers to a comprehensive explanation of an important feature of nature supported by facts gathered over time. Theories also allow scientists to make predictions about as yet unobserved phenomena."

"A scientific theory is a well-substantiated explanation of some aspect of the natural world, based on a body of facts that have been repeatedly confirmed through observation and experimentation. Such fact-supported theories are not "guesses" but reliable accounts of the real world. The theory of biological evolution is more than "just a theory." It is as factual an explanation of the universe as the atomic theory of matter (stating that everything is made of atoms) or the germ theory of disease (which states that many diseases are caused by germs). Our understanding of gravity is still a work in progress. But the phenomenon of gravity, like evolution, is an accepted fact.

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

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

| *Theories aren't likely to change.*

What is a Law?

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

For example, around the year 1800, Jacques Charles and other scientists were working with gases to, among other reasons, improve the design of the hot air balloon. These scientists found, after many, many tests, that certain patterns existed in the observations on gas behavior. If the temperature of the gas is increased, the volume of the gas increased. This is known as a natural law. A law is a relationship that exists between variables in a group of data. Laws describe the patterns we see in large amounts of data, but do not describe why the patterns exist.

What is a Belief?

A **belief** is a statement that is not scientifically provable. Beliefs may or may not be incorrect; they just are outside the realm of science to explore.

 Laws vs. Theories

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



Video 2.2.1: What's the difference between a scientific law and theory?

Summary

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

Contributions & Attributions

-
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2.3: Scientific Notation - Writing Large and Small Numbers

Learning Objectives

- Express a large number or a small number in scientific notation.
- Carry out arithmetical operations and express the final answer in 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 would require a display with at least 24 decimal places. A system called **scientific notation** avoids much of the tedium and awkwardness of manipulating numbers with large or small magnitudes. 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 from 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 2.3.1.

✓ Example 2.3.1: Expressing Numbers in 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

Solutions to Example 2.2.1

	Explanation	Answer
a	To convert 637.8 to a number from 1 to 10, we move the decimal point two places to the left: 637.8 Because the decimal point was moved two places to the left, $n = 2$.	6.378×10^2
b	To convert 0.0479 to a number from 1 to 10, we move the decimal point two places to the right: 0.0479 Because the decimal point was moved two places to the right, $n = -2$.	4.79×10^{-2}
c	This is usually expressed simply as 7.86. (Recall that $10^0 = 1$.)	7.86×10^0
d	Because the decimal point was moved four places to the left, $n = 4$.	1.2378×10^4
e	Because the decimal point was moved four places to the right, $n = -4$.	3.2×10^{-4}
f	Because the decimal point was moved one place to the left, $n = 1$.	6.106700×10^1
g	Because the decimal point was moved three places to the left, $n = 3$.	2.002080×10^3
h	Because the decimal point was moved two places to the right, $n = -2$.	1.020×10^{-2}

Addition and Subtraction

Before numbers expressed in scientific notation can be added or subtracted, they must be converted to a form in which all the exponents have the same value. The appropriate operation is then carried out on the values of N . Example 2.3.2 illustrates how to do this.

✓ Example 2.3.2: Expressing Sums and Differences in Scientific Notation

Carry out the appropriate operation and then express the answer in scientific notation.

a. $(1.36 \times 10^2) + (4.73 \times 10^3)$

b. $(6.923 \times 10^{-3}) - (8.756 \times 10^{-4})$

Solution

Solutions to Example 2.2.2.

	Explanation	Answer
a	Both exponents must have the same value, so these numbers are converted to either $(1.36 \times 10^2) + (47.3 \times 10^2) =$ $(1.36 + 47.3) \times 10^2 = 48.66 \times 10^2$ or $(0.136 \times 10^3) + (4.73 \times 10^3) =$ $(0.136 + 4.73) \times 10^3 = 4.87 \times 10^3.$ Choosing either alternative gives the same answer, reported to two decimal places. In converting 48.66×10^2 to scientific notation, n has become more positive by 1 because the value of N has decreased.	4.87×10^3
b	Converting the exponents to the same value gives either $(6.923 \times 10^{-3}) - (0.8756 \times 10^{-3}) =$ $(6.923 - 0.8756) \times 10^{-3}$ or $(69.23 \times 10^{-4}) - (8.756 \times 10^{-4}) =$ $(69.23 - 8.756) \times 10^{-4} = 60.474 \times 10^{-4}.$ In converting 60.474×10^{-4} to scientific notation, n has become more positive by 1 because the value of N has decreased.	6.047×10^{-3}

Multiplication and Division

When multiplying numbers expressed in scientific notation, we multiply the values of N and add together the values of n . Conversely, when dividing, we divide N in the dividend (the number being divided) by N in the divisor (the number by which we are dividing) and then subtract n in the divisor from n in the dividend. In contrast to addition and subtraction, the exponents do not have to be the same in multiplication and division. Examples of problems involving multiplication and division are shown in Example 2.3.3.

✓ Example 2.3.3: Expressing Products and Quotients in Scientific Notation

Perform the appropriate operation and express your answer in scientific notation.

a. $(6.022 \times 10^{23})(6.42 \times 10^{-2})$

b. $\frac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}}$

c. $\frac{(6.63 \times 10^{-34})(6.0 \times 10)}{8.52 \times 10^{-2}}$

Solution

Solution to Example 2.2.3

Explanation		
a	In multiplication, we add the exponents: $(6.022 \times 10^{23})(6.42 \times 10^{-2}) = (6.022)(6.42) \times 10^{[23+(-2)]} = 38.7 \times 10^{21}$ In converting 38.7×10^{21} to scientific notation, n has become more positive by 1 because the value of N has decreased.	3.87×10^{22}
b	In division, we subtract the exponents: $\frac{1.67 \times 10^{-24}}{9.12 \times 10^{-28}} = \frac{1.67}{9.12} \times 10^{[-24-(-28)]} = 0.183 \times 10^4$ In converting 0.183×10^4 to scientific notation, n has become more negative by 1 because the value of N has increased.	1.83×10^3
c	This problem has both multiplication and division: $\frac{(6.63 \times 10^{-34})(6.0 \times 10)}{(8.52 \times 10^{-2})} = \frac{39.78}{8.52} \times 10^{[-34+1-(-2)]}$	4.7×10^{-31}

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2.4: Significant Figures - Writing Numbers to Reflect Precision

Learning Objectives

- Identify the number of significant figures in a reported value.

The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit. In the ruler illustration below, the bottom ruler gave a length with 2 significant figures, while the top ruler gave a length with 3 significant figures. In a correctly reported measurement, the final digit is significant but not certain. Insignificant digits are not reported. With either ruler, it would not be possible to report the length at **2.553 cm** as there is no possible way that the thousandths digit could be estimated. The 3 is not significant and would not be reported.

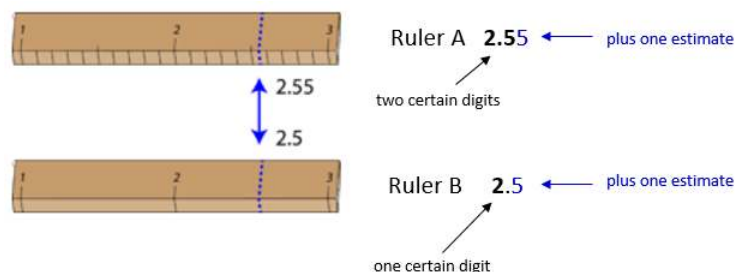


Figure 2.4.1: Measurement with two different rulers.

Ruler A's measurement can be rounded to 2.55, with 2 certain digits, while Ruler B's measurement of 2.5 has 1 certain digit

Measurement Uncertainty

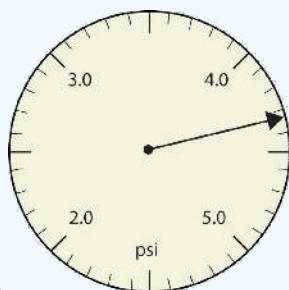
Some error or **uncertainty** always exists in any measurement. The amount of uncertainty depends both upon the skill of the measurer and upon the quality of the measuring tool. While some balances are capable of measuring masses only to the nearest **0.1 g**, other highly sensitive balances are capable of measuring to the nearest **0.001 g** or even better. Many measuring tools such as rulers and graduated cylinders have small lines which need to be carefully read in order to make a measurement. Figure 2.4.1 shows two rulers making the same measurement of an object (indicated by the blue arrow).

With either ruler, it is clear that the length of the object is between **2** and **3 cm**. The bottom ruler contains no millimeter markings. With that ruler, the tenths digit can be estimated and the length may be reported as **2.5 cm**. However, another person may judge that the measurement is **2.4 cm** or perhaps **2.6 cm**. While the 2 is known for certain, the value of the tenths digit is uncertain.

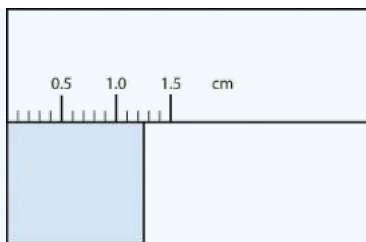
The top ruler contains marks for tenths of a centimeter (millimeters). Now the same object may be measured as **2.55 cm**. The measurer is capable of estimating the hundredths digit because he can be certain that the tenths digit is a 5. Again, another measurer may report the length to be **2.54 cm** or **2.56 cm**. In this case, there are two certain digits (the 2 and the 5), with the hundredths digit being uncertain. Clearly, the top ruler is a superior ruler for measuring lengths as precisely as possible.

✓ Example 2.4.1: Reporting Measurements to the Proper Number of Significant Figures

Use each diagram to report a measurement to the proper number of significant figures.



-
-



Ruler measuring a rectangle in units of centimeters, with the rectangle's edge between 1.2 and 1.3 cm marks

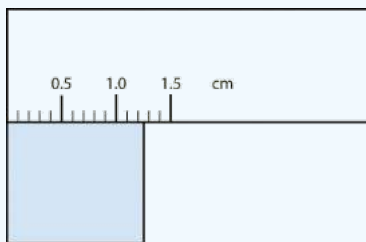
Solutions

Solutions to Example 2.3.1

	Explanation	Answer
a.	The arrow is between 4.0 and 5.0, so the measurement is at least 4.0. The arrow is between the third and fourth small tick marks, so it's at least 0.3. We will have to estimate the last place. It looks like about one-third of the way across the space, so let us estimate the hundredths place as 3. The symbol psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire. The measurement is reported to three significant figures.	4.33 psi
b.	The rectangle is at least 1.0 cm wide but certainly not 2.0 cm wide, so the first significant digit is 1. The rectangle's width is past the second tick mark but not the third; if each tick mark represents 0.1, then the rectangle is at least 0.2 in the next significant digit. We have to estimate the next place because there are no markings to guide us. It appears to be about halfway between 0.2 and 0.3, so we will estimate the next place to be a 5. Thus, the measured width of the rectangle is 1.25 cm. The measurement is reported to three significant figures.	1.25 cm

? Exercise 2.4.1

What would be the reported width of this rectangle?



Answer

1.25 cm

When you look at a reported measurement, it is necessary to be able to count the number of significant figures. The table below details the rules for determining the number of significant figures in a reported measurement. For the examples in the table, assume that the quantities are correctly reported values of a measured quantity.

Table 2.4.1: Significant Figure Rules

Rule	Examples
1. All nonzero digits in a measurement are significant.	237 has three significant figures. 1.897 has four significant figures.
2. Zeros that appear between other nonzero digits (middle zeros) are always significant.	39,004 has five significant figures. 5.02 has three significant figures.
3. Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.	0.008 has one significant figure. 0.000416 has three significant figures.
4. Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant. Use scientific notation to indicate the appropriate number of significant figures.	1400 is ambiguous. 1.4×10^3 has two significant figures. 1.40×10^3 three significant figures. 1.400×10^3 has four significant figures.
5. Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.	620.0 has four significant figures. 19.000 has five significant figures.

Exact Numbers

Integers obtained either by counting objects or from definitions are exact numbers, which are considered to have infinitely many significant figures. If we have counted four objects, for example, then the number 4 has an infinite number of significant figures (i.e., it represents 4.000...). Similarly, 1 foot (ft) is defined to contain 12 inches (in), so the number 12 in the following equation has infinitely many significant figures:

✓ Example 2.4.2

Give the number of significant figures in each. Identify the rule for each.

- 5.87
- 0.031
- 52.90
- 00.2001
- 500
- 6 atoms

Solution

Solution to Example 2.3.2

	Explanation	Answer
a	All three numbers are significant (rule 1).	5.87, three significant figures
b	The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1).	0.031, two significant figures
c	The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5).	52.90, four significant figures
d	The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2).	00.2001, four significant figures
e	The number is ambiguous. It could have one, two or three significant figures.	500, ambiguous
f	The 6 is a counting number. A counting number is an exact number.	6, infinite

? Exercise 2.4.2

Give the number of significant figures in each.

- a. 36.7 m
- b. 0.006606 s
- c. 2,002 kg
- d. 306,490,000 people
- e. 3,800 g

Answer a

three significant figures

Answer b

four significant figures

Answer c

four significant figures

Answer d

infinite (exact number)

Answer e

Ambiguous, could be two, three or four significant figures.

Accuracy and Precision

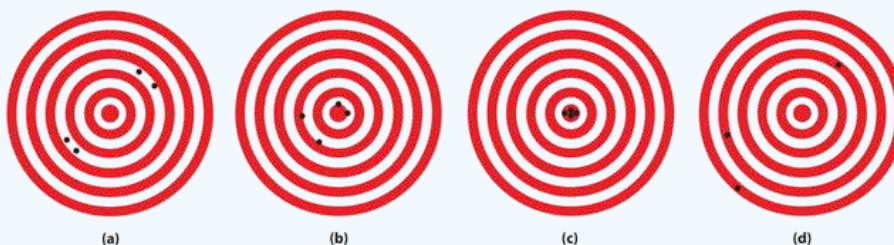
Measurements may be accurate, meaning that the measured value is the same as the true value; they may be precise, meaning that multiple measurements give nearly identical values (i.e., reproducible results); they may be both accurate and precise; or they may be neither accurate nor precise. The goal of scientists is to obtain measured values that are both accurate and precise. The video below demonstrates the concepts of accuracy and precision.



Video 2.4.1: Difference between precision and accuracy.

✓ Example 2.4.3

The following archery targets show marks that represent the results of four sets of measurements.



Which target shows

- a precise, but inaccurate set of measurements?
- a set of measurements that is both precise and accurate?
- a set of measurements that is neither precise nor accurate?

Solution

- Set **(a)** is precise, but inaccurate.
- Set **(c)** is both precise and accurate.
- Set **(d)** is neither precise nor accurate.

Summary

Uncertainty exists in all measurements. The degree of uncertainty is affected in part by the quality of the measuring tool. Significant figures give an indication of the certainty of a measurement. Rules allow decisions to be made about how many digits to use in any given situation.

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2.5: Significant Figures in Calculations

Learning Objectives

- Use significant figures correctly in arithmetical operations.

Rounding

Before dealing with the specifics of the rules for determining the significant figures in a calculated result, we need to be able to round numbers correctly. To **round** a number, first decide how many significant figures the number should have. Once you know that, round to that many digits, starting from the left. If the number immediately to the right of the last significant digit is less than 5, it is dropped and the value of the last significant digit remains the same. If the number immediately to the right of the last significant digit is greater than or equal to 5, the last significant digit is increased by 1.

Consider the measurement **207.518 m**. Right now, the measurement contains six significant figures. How would we successively round it to fewer and fewer significant figures? Follow the process as outlined in Table 2.5.1.

Table 2.5.1: Rounding examples

Number of Significant Figures	Rounded Value	Reasoning
6	207.518	All digits are significant
5	207.52	8 rounds the 1 up to 2
4	207.5	2 is dropped
3	208	5 rounds the 7 up to 8
2	210	8 is replaced by a 0 and rounds the 0 up to 1
1	200	1 is replaced by a 0

Notice that the more rounding that is done, the less reliable the figure is. An approximate value may be sufficient for some purposes, but scientific work requires a much higher level of detail.

It is important to be aware of significant figures when you are mathematically manipulating numbers. For example, dividing 125 by 307 on a calculator gives 0.4071661238... to an infinite number of digits. But do the digits in this answer have any practical meaning, especially when you are starting with numbers that have only three significant figures each? When performing mathematical operations, there are two rules for limiting the number of significant figures in an answer—one rule is for addition and subtraction, and one rule is for multiplication and division.

*In operations involving significant figures, the answer is reported in such a way that it reflects the reliability of the **least precise** operation. An answer is no more precise than the least precise number used to get the answer.*

Multiplication and Division

For multiplication or division, the rule is to count the number of significant figures in each number being multiplied or divided and then limit the significant figures in the answer to the lowest count. An example is as follows:

$$\underbrace{38.65}_{4 \text{ sig figs}} \times \underbrace{105.93}_{5 \text{ sig figs}} = \underbrace{4,094.1945}_{\text{reduce to 4 sig figs}}$$

The final answer, limited to four significant figures, is 4,094. The first digit dropped is 1, so we do not round up.

Scientific notation provides a way of communicating significant figures without ambiguity. You simply include all the significant figures in the leading number. For example, the number 450 has two significant figures and would be written in scientific notation as 4.5×10^2 , whereas 450.0 has four significant figures and would be written as 4.500×10^2 . In scientific notation, all significant figures are listed explicitly.

✓ Example 2.5.1

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

a. 23.096×90.300

b. 125×9.000

Solution

a

Explanation	Answer
The calculator answer is 2,085.5688, but we need to round it to five significant figures. Because the first digit to be dropped (in the tenths place) is greater than 5, we round up to 2,085.6.	2.0856×10^3

b

Explanation	Answer
The calculator gives 1,125 as the answer, but we limit it to three significant figures.	1.13×10^3

Addition and Subtraction

How are significant figures handled in calculations? It depends on what type of calculation is being performed. If the calculation is an addition or a subtraction, the rule is as follows: limit the reported answer to the rightmost column that all numbers have significant figures in common. For example, if you were to add 1.2 and 4.71, we note that the first number stops its significant figures in the tenths column, while the second number stops its significant figures in the hundredths column. We therefore limit our answer to the tenths column.

$$\begin{array}{r}
 1.2 \\
 4.41 \\
 \hline
 5.61 \\
 \uparrow \text{ limit final answer to the tenths column: } 5.6
 \end{array}$$

We drop the last digit—the 1—because it is not significant to the final answer.

The dropping of positions in sums and differences brings up the topic of rounding. Although there are several conventions, in this text we will adopt the following rule: the final answer should be rounded up if the first dropped digit is 5 or greater, and rounded down if the first dropped digit is less than 5.

$$\begin{array}{r}
 77.2 \\
 10.46 \\
 \hline
 87.66 \\
 \uparrow \text{ limit final answer to the tenths column and round up: } 87.7
 \end{array}$$

✓ Example 2.5.2

- a. $13.77 + 908.226$
 b. $1,027 + 611 + 363.06$

Solution

a

Explanation	Answer
The calculator answer is 921.996, but because 13.77 has its farthest-right significant figure in the hundredths place, we need to round the final answer to the hundredths position. Because the first digit to be dropped (in the thousandths place) is greater than 5, we round up to 922.00	$922.00 = 9.2200 \times 10^2$

b

Explanation	Answer
The calculator gives 2,001.06 as the answer, but because 611 and 1027 has its farthest-right significant figure in the ones place, the final answer must be limited to the ones position.	$2,001.06 = 2.001 \times 10^3$

? Exercise 2.5.2

Write the answer for each expression using scientific notation with the appropriate number of significant figures.

- a. $217 \div 903$
 b. $13.77 + 908.226 + 515$
 c. $255.0 - 99$
 d. 0.00666×321

Answer a:

$$0.240 = 2.40 \times 10^{-1}$$

Answer b:

$$1,437 = 1.437 \times 10^3$$

Answer c:

$$156 = 1.56 \times 10^2$$

Answer d:

$$2.14 = 2.14 \times 10^0$$

Remember that calculators do not understand significant figures. *You* are the one who must apply the rules of significant figures to a result from your calculator.

Calculations Involving Multiplication/Division and Addition/Subtraction

In practice, chemists generally work with a calculator and carry all digits forward through subsequent calculations. When working on paper, however, we often want to minimize the number of digits we have to write out. Because successive rounding can compound inaccuracies, intermediate rounding needs to be handled correctly. When working on paper, always round an intermediate result so as to retain at least one more digit than can be justified and carry this number into the next step in the calculation. The final answer is then rounded to the correct number of significant figures at the very end.



Video 2.5.1: Significant figures in mixed operations (<https://www.youtube.com/watch?v=yBntMndXQWA>).



Video 2.5.2: https://www.youtube.com/watch?v=__csP0NtlGI

In the worked examples in this text, we will often show the results of intermediate steps in a calculation. In doing so, we will show the results to only the correct number of significant figures allowed for that step, in effect treating each step as a separate calculation. This procedure is intended to reinforce the rules for determining the number of significant figures, but in some cases it may give a final answer that differs in the last digit from that obtained using a calculator, where all digits are carried through to the last step.

✓ Example 2.5.3

- $2(1.008 \text{ g}) + 15.99 \text{ g}$
- $137.3 \text{ s} + 2(35.45 \text{ s})$
- $\frac{118.7\text{g}}{2} - 35.5\text{g}$

Solution

a.

Explanation	Answer
$2(1.008 \text{ g}) + 15.99 \text{ g} =$ Perform multiplication first. $2(1.008 \text{ g } 4 \text{ sig figs}) = 2.016 \text{ g } 4 \text{ sig figs}$ The number with the least number of significant figures is 1.008 g; <i>the number 2 is an exact number and therefore has an infinite number of significant figures.</i> Then, perform the addition. $2.016 \text{ g thousandths place} + 15.99 \text{ g hundredths place (least precise)} = 18.006 \text{ g}$ Round the final answer. Round the final answer to the hundredths place since 15.99 has its farthest right significant figure in the hundredths place (least precise).	18.01 g (rounding up)

b.

Explanation	Answer
$137.3 \text{ s} + 2(35.45 \text{ s}) =$ Perform multiplication first. $2(35.45 \text{ s } 4 \text{ sig figs}) = 70.90 \text{ s } 4 \text{ sig figs}$ The number with the least number of significant figures is 35.45; <i>the number 2 is an exact number and therefore has an infinite number of significant figures.</i> Then, perform the addition. $137.3 \text{ s tenths place (least precise)} + 70.90 \text{ s hundredths place} = 208.20 \text{ s}$ Round the final answer. Round the final answer to the tenths place based on 137.3 s.	208.2 s

c.

Explanation	Answer
$\frac{118.7\text{g}}{2} - 35.5\text{g} =$ Perform division first. $\frac{118.7\text{g}}{2} 4 \text{ sig figs} = 59.35 \text{ g } 4 \text{ sig figs}$ The number with the least number of significant figures is 118.7 g; <i>the number 2 is an exact number and therefore has an infinite number of significant figures.</i> Perform subtraction next. $59.35 \text{ g hundredths place} - 35.5 \text{ g tenths place (least precise)} = 23.85 \text{ g}$ Round the final answer. Round the final answer to the tenths place based on 35.5 g.	23.9 g (rounding up)

? Exercise 2.5.3

Complete the calculations and report your answers using the correct number of significant figures.

- a. $5(1.008\text{s}) - 10.66\text{ s}$
- b. $99.0\text{ cm} + 2(5.56\text{ cm})$

Answer a

-5.62 s

Answer b

110.2 cm

Summary

- Rounding
 - If the number to be dropped is greater than or equal to 5, increase the number to its left by 1 (e.g. 2.9699 rounded to three significant figures is 2.97).
 - If the number to be dropped is less than 5, there is no change (e.g. 4.00443 rounded to four significant figures is 4.004).
- The rule in multiplication and division is that the final answer should have the same number of significant figures as there are in the number with the fewest significant figures.
- The rule in addition and subtraction is that the answer is given the same number of decimal places as the term with the fewest decimal places.

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2.6: The Basic Units of Measurement

Learning Objectives

- State the different measurement systems used in chemistry.
- Describe how prefixes are used in the metric system and identify how the prefixes milli-, centi-, and kilo- compare to the base unit.

How long is a yard? It depends on whom you ask and when you asked the question. Today we have a standard definition of the yard, which you can see marked on every football field. If you move the ball ten yards, you get a first down and it does not matter whether you are playing in Los Angeles, Dallas, or Green Bay. But at one time that yard was arbitrarily defined as the distance from the tip of the king's nose to the end of his outstretched hand. Of course, the problem there is simple: new king, new distance (and then you have to re-mark all of those football fields).

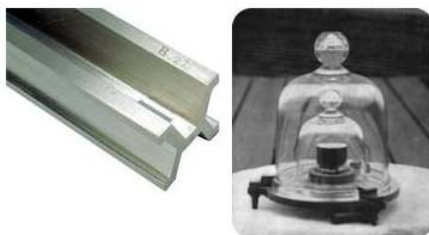


Figure 2.6.1: Meter standard (left) and Kilogram standard (right).

SI Base Units

All measurements depend on the use of units that are well known and understood. The **English system** of measurement units (inches, feet, ounces, etc.) are not used in science because of the difficulty in converting from one unit to another. The **metric system** is used because all metric units are based on multiples of 10, making conversions very simple. The metric system was originally established in France in 1795. **The International System of Units** is a system of measurement based on the metric system. The acronym **SI** is commonly used to refer to this system and stands for the French term, *Le Système International d'Unités*. The SI was adopted by international agreement in 1960 and is composed of seven base units in Table 2.6.1.

Table 2.6.1: SI Base Units of Measurement

Quantity	SI Base Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Temperature	kelvin	K
Time	second	s
Amount of a Substance	mole	mol
Electric Current	ampere	A
Luminous Intensity	candela	cd

The first units are frequently encountered in chemistry. All other measurement quantities, such as volume, force, and energy, can be derived from these seven base units.

Unfortunately, the Metric System is Not Ubiquitous

The map below shows the adoption of the SI units in countries around the world. The United States has legally adopted the metric system for measurements, but does not use it in everyday practice. Great Britain and much of Canada use a combination of metric and imperial units.

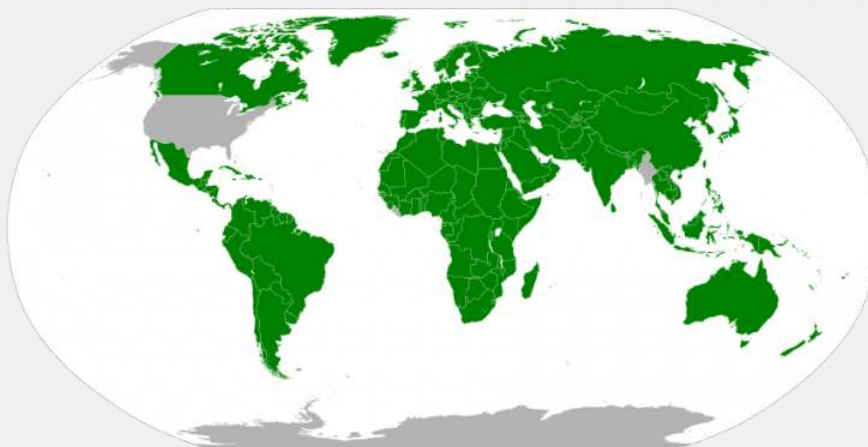


Figure 2.6.1: Areas of world using metric system (in green). Only a few countries are slow or resistant to adoption including the United States.

Prefix Multipliers

Conversions between metric system units are straightforward because the system is based on powers of ten. For example, meters, centimeters, and millimeters are all metric units of length. There are 10 millimeters in 1 centimeter and 100 centimeters in 1 meter. **Metric prefixes** are used to distinguish between units of different size. These prefixes all derive from either Latin or Greek terms. For example, *mega* comes from the Greek word *μεγας*, meaning "great". Table 2.6.2 lists the most common metric prefixes and their relationship to the central unit that has no prefix. Length is used as an example to demonstrate the relative size of each prefixed unit.

Table 2.6.2: SI Prefixes

Prefix	Unit Abbreviation	Meaning	Example
giga	G	1,000,000,000	1 gigameter (Gm) = 10^9 m
mega	M	1,000,000	1 megameter (Mm) = 10^6 m
kilo	k	1,000	1 kilometer (km) = 1,000 m
hecto	h	100	1 hectometer (hm) = 100 m
deka	da	10	1 dekameter (dam) = 10 m
		1	1 meter (m)
deci	d	1/10	1 decimeter (dm) = 0.1 m
centi	c	1/100	1 centimeter (cm) = 0.01 m
milli	m	1/1,000	1 millimeter (mm) = 0.001 m
micro	μ	1/1,000,000	1 micrometer (μ m) = 10^{-6} m
nano	n	1/1,000,000,000	1 nanometer (nm) = 10^{-9} m
pico	p	1/1,000,000,000,000	1 picometer (pm) = 10^{-12} m

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. We use "m" for meter and not "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So we would write 3.5 milliliters as **3.5 mL**.

As a practical matter, whenever possible you should express the units in a small and manageable number. If you are measuring the weight of a material that weighs **6.5 kg**, this is easier than saying it weighs **6500 g** or **0.65 dag**. All three are correct, but the **kg** units in this case make for a small and easily managed number. However, if a specific problem needs grams instead of kilograms, go with the grams for consistency.

✓ Example 2.6.1: Unit Abbreviations

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

- kiloliter
- microsecond
- decimeter
- nanogram

Solutions

Solutions to Example 2.5.1

	Explanation	Answer
a	The prefix kilo means "1,000 ×," so 1 kL equals 1,000 L.	kL
b	The prefix micro implies 1/1,000,000th of a unit, so 1 μs equals 0.000001 s.	μs
c	The prefix deci means 1/10th, so 1 dm equals 0.1 m.	dm
d	The prefix nano means 1/1000000000, so a nanogram is equal to 0.000000001 g.	ng

? Exercise 2.6.1

Give the abbreviation for each unit and define the abbreviation in terms of the base unit.

- kilometer
- milligram
- nanosecond
- centiliter

Answer a:

km

Answer b:

mg

Answer c:

ns

Answer d:

cL

Summary

- Metric prefixes derive from Latin or Greek terms. The prefixes are used to make the units manageable.
- The SI system is based on multiples of ten. There are seven basic units in the SI system. Five of these units are commonly used in chemistry.

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