SURVEY OF GENERAL CHEMISTRY



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1: FOUNDATIONS FOR CHEMISTRY

1.1: What is Chemistry?
1.2: The Scientific Method
1.3: Working with Numbers in Chemistry
1.4: Math Concepts Needed in Chemistry
1.5: Using a Scientific Calculator
1.6: Problem Solving in Chemistry

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1.1: What is Chemistry?

Learning Objectives

- Learn the definitions of science and chemistry.
- Learn why chemistry is often referred to as the central science.

Chemistry is a science, but more specifically it is a natural science. **Science** is the broad, overall system of gaining knowledge through observation and testing, whereas a **natural science** is focused on learning about the natural world. Natural sciences include physics, chemistry, biology, geology and astronomy. Science also uses mathematics and logic, which are sometimes called "formal sciences". People who study and research science are **scientists**. Scientists study things by looking at them very carefully, by measuring, and by doing experiments. Scientists try to explain why things act the way that they do, and predict what will happen.

Chemistry is the scientific discipline that involves the properties and behavior of matter, and the changes matter undergoes during a reaction with other substances. In the scope of its subject, chemistry occupies an intermediate position between physics and biology. It is sometimes called *the central science*, because it provides a foundation for understanding both basic and applied scientific disciplines at a fundamental level (Figure 1.1.1). For example, chemistry explains aspects of plant chemistry (botany), the formation of igneous rocks (geology), how atmospheric ozone is formed and how environmental pollutants are degraded (ecology), the properties of the soil on the moon (astrophysics), how medications work (pharmacology), and how to collect DNA evidence at a crime scene (forensics).

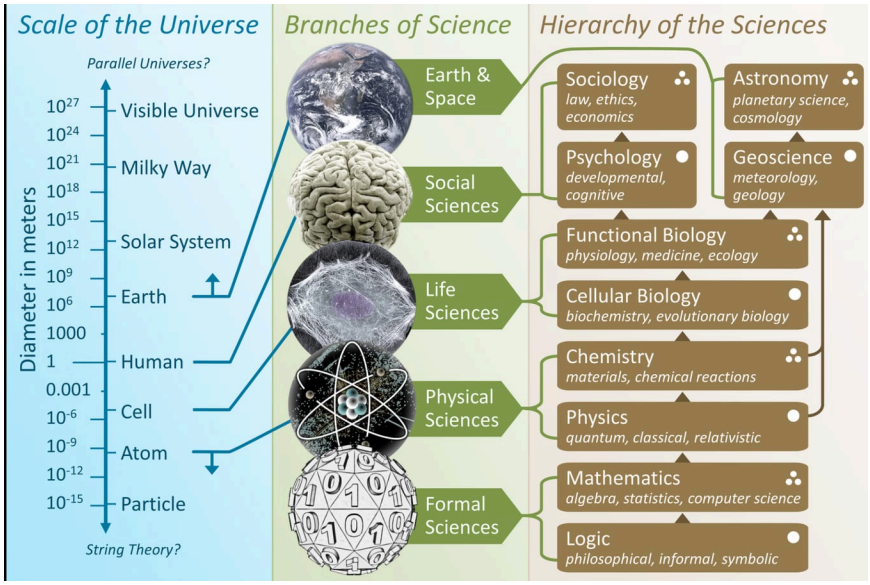


Figure 1.1.1: The scale of the universe mapped to the branches of science. (Source: Wikipedia)

Summary

- Science is the broad, overall system of gaining knowledge through observation and testing.
- Chemistry is the branch of science involved with the properties and behavior of matter, and the changes it undergoes during a reaction with other substances.
- Chemistry provides foundational understanding for both basic and applied sciences.

Contributors and Attributions

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1.2: The Scientific Method

Learning Objectives

- Describe the differences between hypothesis and theory as scientific terms.
- Describe the difference between a theory and scientific law.
- Identify the components of the scientific method.

Scientists search for answers to questions and solutions to problems by using a procedure called the **scientific method**. This procedure consists of making observations, formulating hypotheses, and designing experiments, which in turn lead to additional observations, hypotheses, and experiments in repeated cycles (Figure 1.2.1).

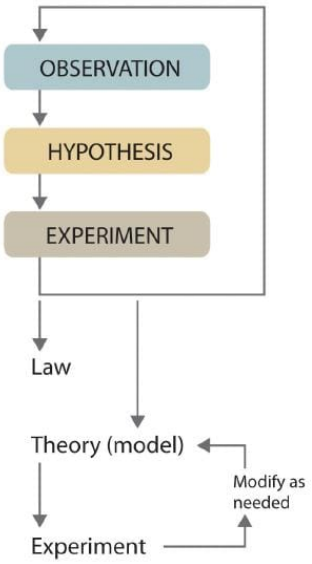


Figure 1.2.1 *The steps in the scientific method.*

The scientific method is a repetitive process used in chemistry to understand matter and how it behaves. The following steps outline the general progression of this method, but it is important to note that step 5 is only reached after a very large number of repititions of steps 1-4.

Step 1: Make observations.

Observations can be qualitative or quantitative. **Qualitative observations** describe properties or occurrences in ways that do not rely on numbers. Examples of qualitative observations include the following: "the outside air temperature is cooler during the winter season," "table salt is a crystalline solid," "sulfur crystals are yellow," and "dissolving a penny in dilute nitric acid forms a blue solution and a brown gas." **Quantitative observations** are measurements, which by definition consist of both a number and a unit. Examples of quantitative observations include the following: "the melting point of crystalline sulfur is 115.21 degrees Celsius," and "35.9 grams of table salt—the chemical name of which is sodium chloride—dissolve in 100 grams of water at 20 degrees Celsius." For the question of the dinosaurs' extinction, the initial observation was quantitative: iridium concentrations in sediments dating to 66 million years ago were 20–160 times higher than normal.

Step 2: Formulate a hypothesis.

After deciding to learn more about an observation or a set of observations, scientists generally begin an investigation by forming a **hypothesis**, a tentative explanation for the observation(s). The hypothesis may not be correct, but it puts the scientist's understanding of the system being studied into a form that can be tested. For example, the observation that we experience

alternating periods of light and darkness which correspond to observed movements of the sun, moon, clouds, and shadows, is consistent with either of two hypotheses:

1. Earth rotates on its axis every 24 hours, alternately exposing one side to the sun.
2. The sun revolves around Earth every 24 hours.

Suitable experiments can be designed to choose between these two alternatives. In the case of disappearance of the dinosaurs, the hypothesis was that the impact of a large extraterrestrial object caused their extinction. Unfortunately (or perhaps fortunately), this hypothesis does not lend itself to direct testing by any obvious experiment, but scientists can collect additional data that either supports or refutes it.

Step 3: Design and perform experiments.

After a hypothesis has been formed, scientists conduct experiments to test its validity. Experiments are systematic observations or measurements, preferably made under controlled conditions—that is, under conditions in which a single variable changes.

Step 4: Accept or modify the hypothesis.

A properly designed and executed experiment enables a scientist to determine whether the original hypothesis is valid. In the case of validity, the scientist can proceed to step 5. In other cases, experiments may demonstrate that the hypothesis is incorrect or that it must be modified, thus requiring further experimentation.

Step 5: Development of a law and/or theory.

More experimental data are then collected and analyzed, at which point a scientist may begin to think that the results are sufficiently reproducible (i.e., dependable) to merit being summarized in a **law**—a verbal or mathematical description of a phenomenon that allows for general predictions. A law simply states what happens; it does not address the question of why.

One example of a law, the law of definite proportions (discovered by the French scientist Joseph Proust [1754–1826]), states that a chemical substance always contains the same proportions of elements by mass. Thus, sodium chloride (table salt) always contains the same proportion by mass of sodium to chlorine—in this case, 39.34% sodium and 60.66% chlorine by mass. Sucrose (table sugar) is always 42.11% carbon, 6.48% hydrogen, and 51.41% oxygen by mass.

Whereas a law states only what happens, a **theory** attempts to explain why nature behaves as it does. Laws are unlikely to change greatly over time, unless a major experimental error is discovered. A theory, in contrast, is incomplete and imperfect; it evolves with time to explain new facts as they are discovered.

Because scientists can enter the cycle shown in Figure 1.2.1 at any point, the actual application of the scientific method to different topics can take many different forms. For example, a scientist may start with a hypothesis formed by reading about work done by others in the field, rather than by making direct observations.

Example 1.2.1

Classify each statement as a law, theory, experiment, hypothesis, or observation.

1. Ice always floats on liquid water.
2. Birds evolved from dinosaurs.
3. Hot air is less dense than cold air, probably because the components of hot air are moving more rapidly.
4. When 10 g of ice was added to 100 mL of water at 25°C, the temperature of the water decreased to 15.5°C after the ice melted.
5. The ingredients of Ivory soap were analyzed to see whether it really is 99.44% pure, as advertised.

Solution

1. This is a general statement of a relationship between the properties of liquid and solid water, so it is a law.

2. This is a possible explanation for the origin of birds, so it is a hypothesis.
3. This is a statement that tries to explain the relationship between the temperature and the density of air based on fundamental principles, so it is a theory.
4. The temperature is measured before and after a change is made in a system, so these are observations.
5. This is an analysis designed to test a hypothesis (in this case, the manufacturer's claim of purity), so it is an experiment.

Exercise 1.2.1

Classify each statement as a law, theory, experiment, hypothesis, qualitative observation, or quantitative observation.

1. Measured amounts of acid were added to a Rolaids tablet to see whether it really "consumes 47 times its weight in excess stomach acid."
2. Heat always flows from hot objects to cooler ones, not in the opposite direction.
3. The universe was formed by a massive explosion that propelled matter into a vacuum.
4. Michael Jordan is the greatest pure shooter ever to play professional basketball.
5. Limestone is relatively insoluble in water, but dissolves readily in dilute acid with the evolution of a gas.

Answer 1: available at the end of this section.

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.
- The scientific method is a method of investigation involving experimentation and observation to acquire new knowledge, solve problems, and answer questions. The key steps in the scientific method are:
  - Step 1: Make observations.
  - Step 2: Formulate a hypothesis.
  - Step 3: Design and perform experiments.
  - Step 4: Accept or modify the hypothesis.
  - Step 5: Development of a law and/or theory.

Contributors and Attributions

-
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Answers

Exercise 1.2.1

experiment
law
theory
hypothesis
observation

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1.3: Working with Numbers in Chemistry

Learning Objective

- Apply the concept of significant figures to limit a measurement to the proper number of digits.
- Recognize the number of significant figures in a given quantity.
- Learn how to round results to the correct number of digits after multiplication/division and addition/subtraction calculations.
- Learn to convert between numbers in standard notation and scientific notation.

As introduced in the previous section, we make qualitative and quantitative observations when studying matter in chemistry. The quantitative observations are called measurements, but when taking and working with measurements it is important to report them to the proper number of digits.

Taking Measurements

Consider using a ruler to measure the width of an object, as shown in Figure 1.3.1. The object is definitely more than 1 cm long, so we know that the first digit in our measurement is 1. We see by counting the tick marks on the ruler that the object is at least three ticks after the 1. If each tick represents 0.1 cm, then we know the object is at least 1.3 cm wide. But our ruler does not have any more ticks between the 0.3 and the 0.4 marks, so we can't know exactly how much the next decimal place is. But with a practiced eye we can estimate it. Let us estimate it as about six-tenths of the way between the third and fourth tick marks, which estimates our hundredths place as 6, so we identify a measurement of 1.36 cm for the width of the object.

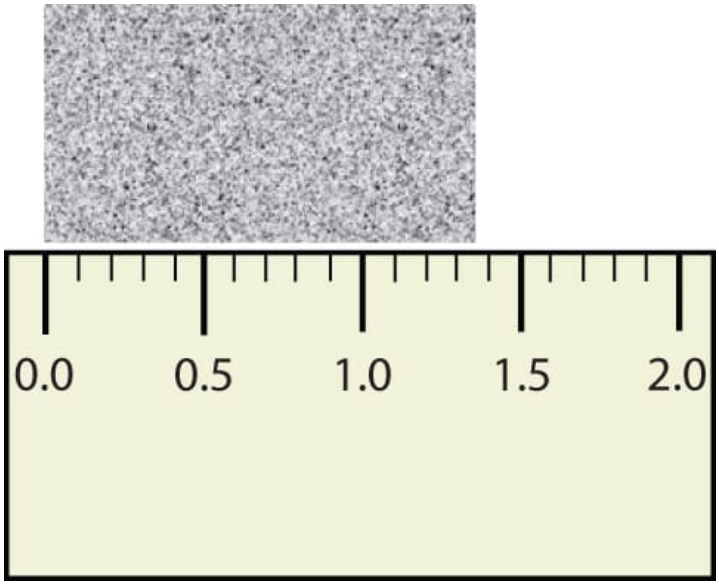


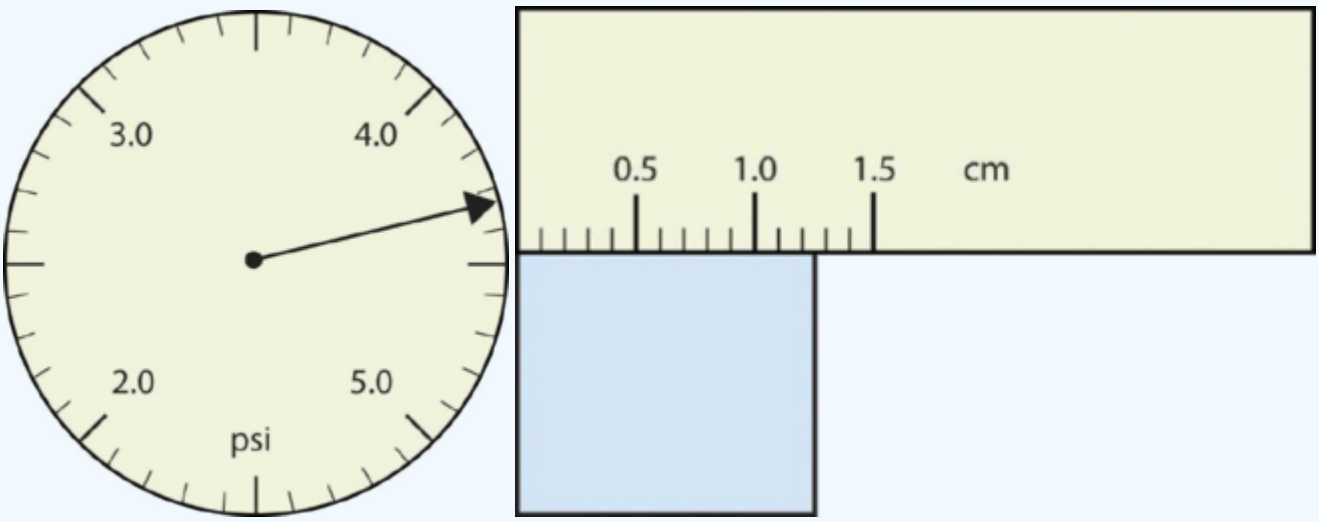
Figure 1.3.1: Expressing Width. What is the proper way to express the width of this object?

Does it make any sense to try to report a thousandths place for the measurement? No, it doesn't; we are not exactly sure of the hundredths place (after all, it was an estimate only), so it would be fruitless to estimate a thousandths place. Our best measurement, then, stops at the hundredths place, and we report 1.36 cm as proper measurement.

This concept of reporting the proper number of digits in a measurement or a calculation is called significant figures. **Significant figures** (sometimes called **significant digits**) represent the limits of what values of a measurement or a calculation we are sure of. The convention for a measurement is that the quantity reported should include all certain digits plus one estimated digit.

Example 1.3.1

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

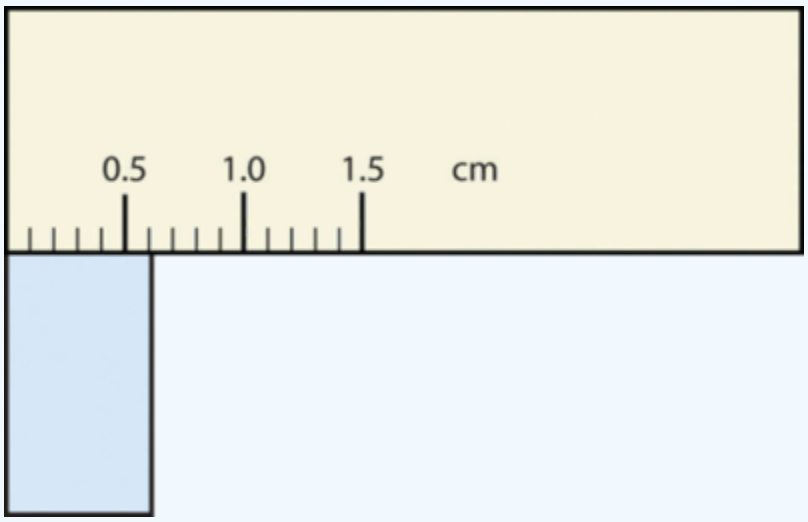


Solution

1. 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. Combining the digits, we have a measurement of 4.33 psi (psi stands for "pounds per square inch" and is a unit of pressure, like air in a tire). We say that the measurement is reported to three significant figures.
2. 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. Again, the measurement is reported to three significant figures.

Exercise 1.3.1

What would be the reported width of this rectangle?



Answer available at the end of this section.

Counting Significant Figures

In many cases, you will be given a measurement. How can you tell by looking what digits are significant? For example, the reported population of the United States is 306,000,000. Does that mean that it is *exactly* three hundred six million, or is some estimation occurring?

The following conventions dictate which numbers in a reported measurement are significant and which are not significant:

1. Any nonzero digit is significant.
2. Any zeros between nonzero digits (i.e., **captive zeros**) are significant.
3. Zeros at the beginning of a decimal number (i.e., **leading zeros**) are not significant; they serve only to put the significant digits in the correct positions.
4. Zeros at the end of a number (i.e., trailing zeros) ...
    a. ... without a decimal point are not significant; they serve only to put the significant digits in the correct positions.
    b. ... with a decimal point are significant.

So, by these rules, the population figure of the United States has only three significant figures: the 3, the 6, and the zero between them. The remaining six zeros simply put the 306 in the millions position.

Example 1.3.2

Give the number of significant figures in each measurement.

a. 36.7 m
b. 0.006606 s
c. 2,002 kg
d. 306,490,000 people

Solution

a. By rule 1, all nonzero digits are significant, so this measurement has three significant figures.
b. By rule 4, the first three zeros are not significant, but by rule 2 the zero between the sixes is; therefore, this number has four significant figures.
c. By rule 2, the two zeros between the twos are significant, so this measurement has four significant figures.
d. The four trailing zeros in the number are not significant, but the other five numbers are, so this number has five significant figures.

Exercise 1.3.2

Give the number of significant figures in each measurement.

a. 0.000601 m
b. 65.080 kg

Answer a available at the end of this section.

As you have probably realized by now, the biggest issue in determining the number of significant figures in a value is the zero. A quick shortcut that you can use to determine the number of significant figures in a value is to start with the first nonzero digit and count to the right, stopping at the last nonzero digit if there is no decimal in the number or counting all the way to the right if there is a decimal in the number.

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 calculated result $207.518\ \text{m}$. Right now, this result contains six significant figures. Table $1.3.1$ shows what the measurement would be rounded to based on the number of significant figures to be reported.

Table $1.3.1$*: Rounding examples*

Number of Significant Figures	Rounded Value	Reasoning
7	207.5180	a trailing zero must be added
6	207.518	the result already has 6 significant figures
5	207.52	the 8 rounds the 1 up to 2
4	207.5	the 1 rounds down, leaving the 5 unchanged
3	208	the 5 rounds the 7 up to 8
2	210	the 7 rounds the 0 up to 1 a 0 is put in place of the 7 to keep the 2 in the hundreds place and 1 in the tens place
1	200	th 0 rounds down, leaving the 2 unchanged a 0 is put in place of the 7 to keep the 2 in the hundreds place

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Significant Figures in Calculation Results

Calculators are great devices, but they do only what is asked of them—no more and no less. The calculator cannot determine how accurate each number is, and the answer given on the screen must be assessed by the user for reliability. As a result, sometimes you will need to round the answer that is given by the calculator whereas sometimes you will need to add trailing zeros that are not shown in order to provide the answer with the correct significant figures.

If you multiply 2.48 times 6.3, you get an answer of 15.687, a value that ignores the number of significant figures in either number. When you divide 12.2 by 1.7, the answer you obtain is 7.176470588, but neither piece of data is accurate to ten significant figures or nine decimal places. The calculator does not know how many digits to report, instead, you have to make the decision about how many digits to report based on the significant figures used in the calculation.

Uncertainty in Multiplication and Division

The density of a certain object is calculated by dividing the mass by the volume. Suppose that a mass of $37.46\ \text{g}$ is divided by a volume of $12.7\ \text{cm}^3$. The result on a calculator would be:

$$D = \frac{m}{V} = \frac{37.46\ \text{g}}{12.7\ \text{cm}^3} = 2.949606299\ \text{g/cm}^3$$

The value of the mass measurement has four significant figures, while the value of the volume measurement has only three significant figures. For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the least number of significant figures. Applying this rule results in a density of $2.95\ \text{g/cm}^3$, with three significant figures—the same as the volume measurement.

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Example 1.3.1

Perform the following calculations, rounding the answers to the appropriate number of significant figures.

A. $0.048\ \text{m} \times 32.97\ \text{m}$

B. $14,570\ \text{kg} \div 5.81\ \text{L}$

Solution

Step 1: Plan the problem.

Analyze each of the measured values to determine how many significant figures should be in the result. Perform the calculation and round appropriately. Apply the correct units to the answer. When multiplying or dividing, the units are also multiplied or divided.

Step 2: Calculate.

A. $0.048\ \text{m} \times 32.97\ \text{m} = 1.6\ \text{m}^2$ Round to two significant figures because 0.048 has two.

B. $14,570\ \text{kg} \div 5.81\ \text{L} = 2510\ \text{kg/L}$ Round to three significant figures because 5.81 has three.

Exercise 1.3.1

How many significant figures should the answer contain?

$10.61 \times 12.133 \times 3.25$

Answer available at the end of this section.

Uncertainty in Addition and Subtraction

Consider two separate mass measurements: $16.7\ \text{g}$ and $5.24\ \text{g}$. The first mass measurement, $(16.7\,\text{g})$, is known only to the tenths place, or to one digit after the decimal point. There is no information about its hundredth place and so that digit cannot be

assumed to be zero. The second measurement, $(5.24\ \text{g})$, is known to the hundredths place, or to two digits after the decimal point.

When these masses are added together, the result on a calculator is $16.7 + 5.24 = 21.94\ \text{g}$. Reporting the answer as $21.94\ \text{g}$ suggests that the sum is known all the way to the hundredths place. However, that cannot be true because the hundredths place of the first mass was completely unknown. The calculated answer needs to be rounded in such a way as to reflect the certainty of each of the measured values that contribute to it. For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the least number of decimal places. The sum of the above masses would be properly rounded to a result of $21.9\ \text{g}$.

Example 1.3.1

Determine the combined molecular mass of a glucose molecule and a maltose molecule.

Glucose molecule = $180.156\frac{g}{mol}$

Maltose molecule = $342.3\frac{g}{mol}$

Solution

$180.156 + 342.4 = 522.456$

When adding and subtracting we know to look at the least number of decimals in our starting values; in this case 342.3 has only 1 digit after the decimal, so we need to round our answer to the same place.

\(522.456\to 522.5\frac{g}{mol}

When working with whole numbers, pay attention to the last significant digit that is to the left of the decimal point, and round your answer to that same point. For example, consider the subtraction: $78,500\ \text{m} - 362\ \text{m}$. The calculated result is $78,138\ \text{m}$. However, the first measurement is known only to the hundreds place, as the 5 is the last significant digit. Rounding the result to that same point means that the correct result is $78,100\ \text{m}$.

Example 1.3.2

What is $4200 + 540$ = ?

Solution

$4200 + 540 = 4740$

To determine where to round our answer, we look at our starting numbers to see which has the fewest number of decimal places. They both have 0 so we round to the nearest whole number, 4740.

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Standard Notation

Standard notation is the straightforward expression of a number. Numbers such as 17, 101.5, and 0.00446 are expressed in

standard notation. For relatively small numbers, standard notation is fine. However, for very large numbers, such as 306,000,000, or for very small numbers, such as 0.000000419, standard notation can be cumbersome because of the number of zeros needed to place nonzero numbers in the proper position.

Scientific Notation

Scientific notation is a shorthand notation for writing very large and very small numbers in a more concise form. A number is expressed in scientific notation by shifting the decimal point immediately after the first nonzero digit followed by all other significant digits times 10 with an exponent that corresponds to the number of places the decimal point was shifted. If the decimal place was shifted to the left the exponent will be positive and if it was shifted to the right it will be negative. Instead of trying to memorize whether the exponent should be positive or negative based on the direction the decimal was shifted, you could consider whether the number in standard notation was a large number (10 or greater) or a small number (less than 1).

For example, to write 79,345 in scientific notation the decimal point must be shifted four places to the left. Thus, the number in scientific notation is 7.9345×10^4. Notice that the exponent is positive because the decimal was shifted to the left, also the starting number was a large number.

Another example, to write 0.000411 in scientific notation the decimal point must be shifted four places to the right. Thus the number in scientific notation is 4.11×10^{-4}. This time the exponent is negative because the decimal was shifted to the right, also the starting number was a small number.

Scientific notation only includes the significant digits before the times 10 portion. This makes scientific notation not only useful for writing very large and very small numbers concisely, but it is also helps to clarify which digits in a reported value are significant.

Example 1.3.1: Expressing Numbers in Scientific Notation

Express these numbers in scientific notation.

a. 306,000
b. 0.00884
c. 2,760,000
d. 0.000000559

Solution

a. The number 306,000 is 3.06 times 100,000, or 3.06 times 10^5. In scientific notation, the number is 3.06×10^5.
b. The number 0.00884 is 8.84 times 1/1,000, which is 8.84 times 10^{-3}. In scientific notation, the number is 8.84×10^{-3}.
c. The number 2,760,000 is 2.76 times 1,000,000, which is the same as 2.76 times 10^6. In scientific notation, the number is written as 2.76×10^6. Note that we omit the zeros at the end of the original number.
d. The number 0.000000559 is 5.59 times 1/10,000,000, which is 5.59 times 10^{-7}. In scientific notation, the number is written as 5.59×10^{-7}.

Exercise 1.3.1

Express these numbers in scientific notation.

a. 23,070
b. 0.0009706

Answer a available at the end of this section.

Answer b available at the end of this section.

Summary

- Significant figures in a quantity include the known digits plus one digit that is estimated.
- There are rules for which numbers in a quantity are significant and which are not significant.
- Rounding involves the adjustment of a value to account for the proper number of significant digits.
- Rules exist for rounding numbers.
- For multiplication and division problems, the answer should be rounded to the same number of significant figures as the measurement with the least number of significant figures.
- For addition and subtraction problems, the answer should be rounded to the same number of decimal places as the measurement with the least number of decimal places.
- Scientific notation expresses a number as a coefficient times a power of 10, where the exponent corresponds to the number of places the decimal is shifted and in what direction it is shifted.

Contributors and Attributions

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Answers

Exercise 1.3.1

0.63 cm

Exercise 1.3.2

three significant figures

five significant figures

Exercise 1.3.1

=418; 3 significant figures

Exercise 1.3.1

2.307×10^{4}

Exercise 1.3.1

9.706×10^{-4}

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1.4: Math Concepts Needed in Chemistry

Learning Objectives

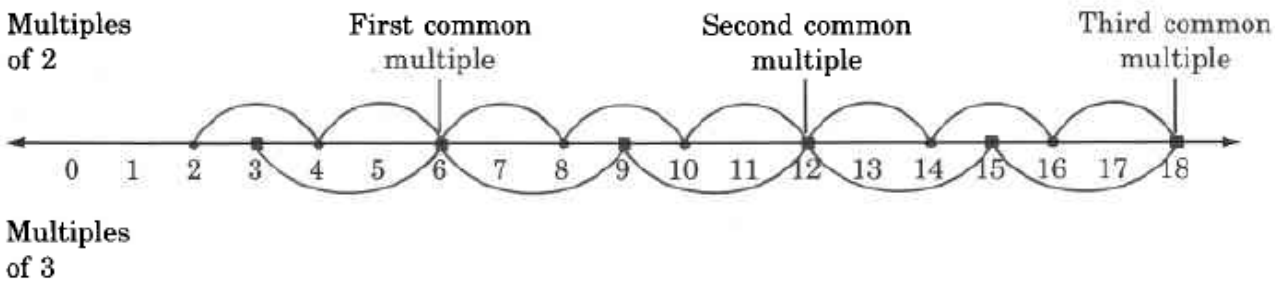
- Be able to find the least common multiple of two or more whole numbers.
- Be able to use a formula to solve for an unknown variable, including units.
- Learn the difference and similarity between units and variables.
- Be able to determine the unit of the result from a calculation using measurements.
- Learn how fractions and percents are related.

Least Common Multiple

When a whole number is multiplied by other whole numbers, with the exception of zero, the resulting products are called **multiples** of the given whole number. See the table below containing multiples of 2, 3, 4, and 5. Also, note that any whole number is a multiple of itself.

Multiples of 2	Multiples of 3	Multiples 4	Multiples of 5
$2 \times 1 = 2$	$3 \times 1 = 3$	$4 \times 1 = 4$	$5 \times 1 = 5$
$2 \times 2 = 4$	$3 \times 2 = 6$	$4 \times 2 = 8$	$5 \times 2 = 10$
$2 \times 3 = 6$	$3 \times 3 = 9$	$4 \times 3 = 12$	$5 \times 3 = 15$
$2 \times 4 = 8$	$3 \times 4 = 12$	$4 \times 4 = 16$	$5 \times 4 = 20$
$2 \times 5 = 10$	$3 \times 5 = 15$	$4 \times 5 = 20$	$5 \times 5 = 25$
$2 \times 6 = 12$	$3 \times 6 = 18$	$4 \times 6 = 24$	$5 \times 6 = 30$
...	...	...	...

Expanding on the concept of multiples, **common multiples** are multiples that are the same for a set of two or more numbers while the **least common multiple** is the smallest of the common multiples for a set of two or more numbers. For example, the least common multiple between 2 and 3 is 6. You can determine this from the multiples of 2 and 3 found in the table above, where 6 and 12 are common multiples of 2 and 3. Although you can continue listing more and more multiples of 2 and 3 to find more common multiples it isn't necessary since we are only looking for the lowest of the common multiples, which is 6 because 6 is less than 12. If you prefer a visual representation for understanding the concept the following diagram may help. The diagram below shows you the common multiples of 2 and 3 each on a number line, and where they intersect are the common multiples. Again, yielding the result that 6 is the least common multiple between 2 and 3.



Being able to determine the least common multiple will be useful when combining ions to write the formula of an ionic compound.

Using Formulas to Solve for Unknowns

In some instances in chemistry you will need to use a formula to solve for an unknown variable. A formula is simply a mathematical equation showing the relationship between multiple variables. In order to solve for an unknown variable in a formula you need to know how to isolate the unknown variable, substitute the known variables into the formula with their units, and use this to calculate the unknown variable and determine its units. In order to isolate a variable in a formula you need to get it on one side of the equation and everything else on the other side. Recall from algebra that you can do whatever you want to an equation as long as you do the same thing to both sides of the equation, which is helpful because it allows you to cancel something from one side of the equation and move it to the other side using inverse relationships which cancel one another out (addition cancels subtraction, subtraction cancels addition, multiplication cancels division, divisiion cancels multiplication). In addition, if you are solving for a variable that is on the bottom of a fraction you cannot leave it on the bottom of the fraction, you have to multiply both sides of the equation by the variable which will move it to the top on the other side of the equation in order to be able to solve for that variable.

The following examples demonstrate solving formulas for an unknown variable when the values of the other variables are given, but for simplicity they exclude all units. However, it is important for you to realize that in the future we will not ignore the units, they will need to be substituted into the equation along with the values and also included with the answer.

Example 1

Use the formula $V = rt$ to solve for t given that $V = 210$ and $r = 7$.

Solution

$V = rt$ Divide both sides of the equation by r to isolate t.

$t = \frac{V}{r}$ Substitute the given values into the formula.

$t = \frac{210}{7}$ Calculate the result.

$t = 30$

Example 2

Use the formula $r = \frac{P}{h}$ to solve for P given that $r = 16$ and $h = 5$.

Solution

$r = \frac{P}{h}$ Multiply both sides of the equation by h to isolate P.

$P = rh$ Substitute the given values into the formula.

$P = (16)(5)$ Calculate the result.

$P = 80$

Example 3

Use the formula $k = \frac{F}{t}$ to solve for t given that $k = 8$ and $F = 20$.

Solution

$k = \frac{F}{t}$ Multiply both sides of the equation by t to move t to the top.

$kt = F$ Divide both sides of the equation by k to isolate t.

$t = \frac{F}{k}$ Substitute the given values into the formula.

$t = \frac{20}{8}$ Calculate the result.

$t = 2.5$

Example 4

Use the formula $T = 2a + b$ to solve for a given that $T = 42$ and $b = 10$.

Solution

$T = 2a + b$ Subtract b from both sides of the equation.

$2a = T - b$ Divide both sides of the equation by 2.

$a = \frac{T-b}{2}$ Substitute the given values into the formula.

$a = \frac{42-10}{2}$ Calculate the result.

$a = 16$

Variables and Units

It is important to realize that variables and units are not interchangeable things. A **variable** is a symbol that stands for an unknown or changing value, whereas a **unit** is a fixed standard that indicates the scale of a value. When plugging measurements into a formula they should include their unit; this helps to determine the unit of the result when the equation is solved. Units behave like variables:

- When you add or subtract measurements that have the same unit the values are added or subtracted and the result has the same unit.
  - Example: $10.00\ \text{mL} - 5.83\ \text{mL} = 4.17\ \text{mL}$
- When you multiply measurements with the same unit the values are multiplied and the units also multiply, changing the exponent of the unit if the multiplied units are the same.
  - Example:$1.50\ \text{L} \times 1.04\ \text{atm} = 1.56\ \text{L} \cdot \text{atm}$
  - Example: $2.00\ \text{in} \times 2.00\ \text{in} = 4.00\ \text{in}^2$
  - Example: $2.00\ \text{in} \times 2.00\ \text{in} \times 2.00\ \text{in} = 8.00\ \text{in}^2$
- When you divide measurements with the same unit the values are divided and the unit are also divided, canceling units if they are the same.
  - Example: $0.5\ \text{g} \div 2.0\ \text{g} = 0.25$
  - Example: $180\ \text{mi} \div 3\text{hr} = 60\frac{\text{mi}}{\text{hr}}$

Fractions and Percents

Fractions and percent essentially tell you the same thing, they describe the ratio of a part to the whole. The difference is that a fraction gives you the information scaled such that the whole is equal to 1 and a percentage is scaled so the whole is 100.

$$Fraction = \frac{Part}{Whole} = \frac{Part}{SumofParts} \tag{1.4.1}$$

Exercise 1.4.1

Consider Mixing 23.5 g of table salt (NaCl) with 91 g of water

a. What is the fraction salt?
b. What is the fraction water?

Answer a available at the end of this section.

Answer b available at the end of this section.

Percentages are 100 times the fraction. While a fraction is unitless, a percentage is reported with a percent symbol, %, as the unit.

$$Percent = Fraction(100) = \frac{Part}{Whole}(100) \tag{1.4.2}$$

Exercise 1.4.2

Consider Mixing 23.5 g of table salt (NaCl) with 91 g of water

a. What is the percent salt?
b. What is the percent water?

Answer a available at the end of this section.

Answer b available at the end of this section.

Answers

Exercise 1.4.1

$$\text{fraction salt} = \frac{m_{salt}}{m_{salt} + m_{water}} = \frac{23.5g}{23.5g + 91g} = \frac{23.5g}{114.5g} = 0.20524 = 0.205$$

Exercise 1.4.1

$$\text{fraction water} = \frac{m_{water}}{m_{salt} + m_{water}} = \frac{91g}{23.5g + 91g} = \frac{91g}{114.5g} = 0.79476 = 0.79 \tag{1.4.3}$$

Exercise 1.4.2

$$\%\ salt = \frac{m_{salt}}{m_{salt} + m_{water}}(100) = \frac{23.5g}{23.5g + 91g}(100) = \frac{23.5g}{114.5g}(100) = 20.524\% = 20.5\% \tag{1.4.4}$$

Exercise 1.4.2

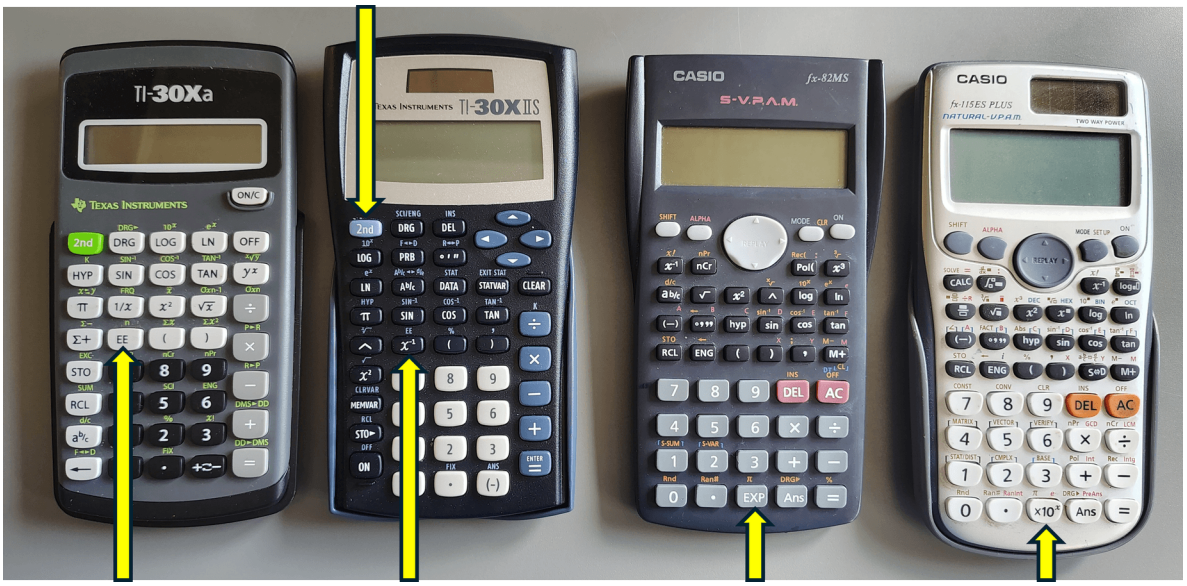
$$\%\ water = \frac{m_{water}}{m_{salt} + m_{water}}(100) = \frac{91g}{23.5g + 91g}(100) = \frac{91g}{114.5g}(100) = 79.476\% = 79\% \tag{1.4.5}$$

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1.5: Using a Scientific Calculator

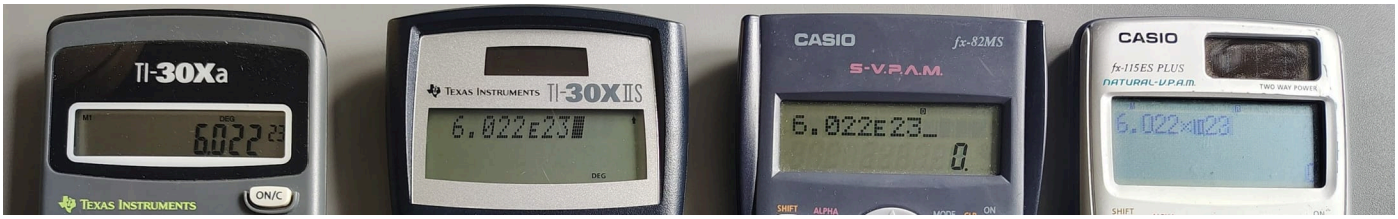
When doing calculations in chemistry you will need to know how to use a scientific calculator for calculations involving scientific notation. Although you can type the times 10 portion using the multiplication button and its exponent using the exponent button (could be ^, y^x, or $x^{\blacksquare}$), that also requires the use of extra parentheses around the entire scientific notation if you are trying to divide by it otherwise the times 10 portion will get multiplied instead of divided. Using the scientific notation button on a scientific calculator is more efficient and eliminates possible errors when divided by scientific notation. Below is an image of a variety of scientific calculators that should help you more easily identify the scientific notation button on most scientific calculators.



Each of the pictured calculators have slightly different scientific notation buttons:

- TI-30Xa - the EE button
- TI-30XIIS - still EE but a 2nd function instead of a regular button, you push the 2nd button and then the x^{-1} button to use the EE function
- fx-82MS - the EXP button
- fx-115ES PLUS - the $\times 10^x$ button

To use these buttons for scientific notation, you first enter the number before the times 10, then push the scientific notation button, and then enter the exponent on the 10. What you see on the screen when you enter your value using the scientific notation can vary slightly depending on the calculator. For example, when entering 6.022×10^{23} into each of the calculators above you will see the following on the screens:



In addition, the output of scientific notation from the calculator may also look slightly different from others and also could look different than how it looks when you enter scientific notation. The picture below shows the same number as entered above as an output from each calculator. The first calculator only has one line so you cannot still see what was entered, but scientific notation looks the same upon entry and as an output on that calculator. The other three calculators are two line calculators where the first line shows what you typed in and the second shows what the output would look like if the value was an answer to a calculation you input.



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1.6: Problem Solving in Chemistry

Converting Units with Dimensional Analysis

Learning Objective

- Be able to write conversion factors from unit relationships.
- Be able to use conversion factors to perform dimensional analysis for converting units.

Consider a simple example: *How many feet are there in 4 yards?* Most people will almost automatically answer that there are 12 feet in 4 yards. But how did you make this determination? Well, if there are 3 feet in 1 yard and there are 4 yards, then there are 4 × 3 = 12 feet in 4 yards.

This is correct, of course, but if you did not already have a good grasp on the units you were trying to convert between it may not be so easy to do in your head. In chemistry we primarily work in metric units which you may not be as familiar with, and thus may not always be able to intuitively convert with little effort.

Instead of relying on memorizing how to do every type of conversion, we use a formal method called **dimensional analysis** to convert measurements from one unit to another. In this method you allow the units to lead the way, showing you what you will need to multiply by or divide by to convert to the unit desired. This is done using conversion factors that are built from known unit relationships. A **conversion factor** is a ratio relating two units that is equal to 1.

To solve the example: *How many feet are there in 4 yards?* we need to know the relationship between yards and feet. We know that 1 yard (yd) equals 3 feet (ft), which can be written as an equality:

$$1\ yd = 3\ ft$$

The rules of algebra say that you can change (i.e., multiply or divide or add or subtract) the equality (as long as you do not divide by zero) and the new expression will still be an equality. For example, suppose we divide both sides of the equation by 1 yard (number **and** unit):

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

The expression is still an equality, by the rules of algebra. The left fraction equals 1, because it has the same quantity in the numerator and the denominator.

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

Similarly, if we divide both sides of the equation by 3 feet:

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

The right fraction equals 1, because it has the same quantity in the numerator and the denominator.

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

Notice that we were able to create two different ratios that equal 1 from one unit relationship. These ratios are conversion factors that can be used to convert between the contained units. The same is true for any unit relationship, you can write two conversion factors from any unit relationship. The table below provides several unit relationships and corresponding conversion factors for units used in everyday life.

Unit Relationship	Conversion Factors
$1\ mi = 1760\ yd$	$\frac{1\,mi}{1760\,yd}\ \frac{1760\,yd}{1\,mi}$
$1\ yd = 3\ ft$	$\frac{1\,yd}{3\,ft}\ \frac{3\,ft}{1\,yd}$
$1\ ft = 12\ in$	$\frac{1\,ft}{12\,in}\ \frac{12\,in}{1\,ft}$
$1\ gal = 4\ qt$	$\frac{1\,gal}{4\,qt}\ \frac{4\,qt}{1\,gal}$
$1\ qt = 2\ pt$	$\frac{1\,qt}{2\,pt}\ \frac{2\,pt}{1\,qt}$
$1\ pt = 2\ cup$	$\frac{1\,pt}{2\,cup}\ \frac{2\,cup}{1\,pt}$
$1\ day = 24\ hr$	$\frac{1\,day}{24\,hr}\ \frac{24\,hr}{1\,day}$
$1\ hr = 60\ min$	$\frac{1\,hr}{60\,min}\ \frac{60\,min}{1\,hr}$
$1\ min = 60\ s$	$\frac{1\,min}{60\,s}\ \frac{60\,s}{1\,min}$

To perform dimensional analysis we start with the measurement we are trying to convert and then we multiply by one or more conversion factors to cancel the unit we have and get to the new unit. However, it is important to realize that multiplying by conversion factors does not change the value represented, this is because conversion factors are equal to 1. Although the number and the unit will change when we perform dimensional analysis to do unit conversions, the actual value represented is still the same. When doing dimensional analysis it is best to start by focusing on the placement of the units in the unit factors and filling in the numbers after the units have been placed in the correct positions, this is because each conversion factor can be written in two ways and in some cases you may need more than one conversion factor.

Now to solve the example: *How many feet are there in 4 yards?* using dimensional analysis. Focusing first on the units of the conversion factor to identify the correct placement:

$$4\ yd \times \frac{yd}{ft}$$

$$4\ yd \times \frac{ft}{yd}$$

The 4 yd term can be thought of as 4yd/1; that is, the yd of the 4 yd is on the top of the fraction. Therefore, the yd in the conversion factor will need to be on the bottom of the fraction in order to cancel, meaning the second option above is the proper placement of units to cancel yd and get to ft. Using the known relationship between yards and feet we then place the 1 with yd and 3 with feet, thus the 3 on top and 1 on the bottom of the conversion factor:

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

When we perform this calculation the yd units cancel, the 4 is multiplied by 3 since it is on the top of the conversion factor, and the result has the remaining unit of ft:

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

Again, we get an answer of 12 ft, just as we did originally. But in this case, we used a more formal procedure that is useful for converting units whether or not we have a good understanding of them. Note: In the above example the answer 12 ft is assuming the question is asking the number of feet in exactly 4 yd, however, if 4 yd is one significant figures then the answer should be reported to one significant figure (10 ft).

Dimensional analysis is not limited to simple, one-step conversions. It can also be applied to multi-step conversions, including

conversions of measurements with fractional units. Below are some examples of multi-step conversions using dimensional analysis. Note: The solutions are provided in steps, first listing unit relationships that link the given units to the units you are looking for, then setting up the calculation with only units in the unit factors, and then filling in the values and solving for the answer.

Example 1.6.1

Use dimensional analysis to convert 12.0 cups to gallons.

Solution

$2\ cup = 1\ pt,\ 2\ pt = 1\ qt,\ 4\ qt = 1\ gal$

$$12.0\ cup \times \frac{pt}{cup} \times \frac{qt}{pt} \times \frac{gal}{qt}$$

$$12.0\ \cancel{cup} \times \frac{1\ \cancel{pt}}{2\ \cancel{cup}} \times \frac{1\ \cancel{qt}}{2\ \cancel{pt}} \times \frac{1\,gal}{4\ \cancel{qt}} = 0.750\ gal$$

Example 1.6.2

Use dimensional analysis to convert 3.00 feet squared to inches squared.

Solution

$1\ ft = 12\ in,\ ft^2 = ft \times ft,\ in^2 = in \times in$

$$3.00\ ft^2 \times \frac{in}{ft} \times \frac{in}{ft}$$

$$3.00\ \cancel{ft^2} \times \frac{12\,in}{1\ \cancel{ft}} \times \frac{12\,in}{1\ \cancel{ft}} = 432\ in^2$$

Example 1.6.3

Use dimensional analysis to convert 75 miles per hour to inches per second.

Solution

$1\ mi = 1760\ yd,\ 1\ yd = 3\ ft,\ 1\ ft = 12\ in,\ 1\ hr = 60\ min,\ 1\ min = 60\ s$

$$\frac{75\,mi}{hr} \times \frac{yd}{mi} \times \frac{ft}{yd} \times \frac{in}{ft} \times \frac{hr}{min} \times \frac{min}{s}$$

$$\frac{75\ \cancel{mi}}{hr} \times \frac{1760\ \cancel{yd}}{1\ \cancel{mi}} \times \frac{3\ \cancel{ft}}{1\ \cancel{yd}} \times \frac{12\,in}{1\ \cancel{ft}} \times \frac{1\ \cancel{hr}}{60\ \cancel{min}} \times \frac{1\ \cancel{min}}{60\,s} = 1300\ \frac{in}{s}$$

Problem Solving Strategy

When initially reading a problem, it can be common to get overwhelmed and not know where to start. This is especially true as more and more concepts build on one another in chemistry. So it is good to have a strategy to fall back on when you are feeling like you don't have any idea how to solve the problem or even where to start. In these cases it is best to focus first on extracting key information from the problem and thinking about things you know that relate the given information to what you are looking for instead of worrying about knowing exactly how to get to the answer right away.

First, list given information that could be helpful and what you are being asked to determine; this includes measurements, unit relationships, or other information specific to substances involved. Next, think about the relationship between what you are given and what you are looking for; this step may require you to look up unit relationships or equations not given in the problem or other information specific to the substances involved. Use the given information and other information you've looked up to map out how to get to the answer requested; focusing on units can often be helpful because many problems in chemistry can be solved by using dimensional analysis without the need for equations.

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2: MATTER AND THE PERIODIC TABLE

2.1: Classification of Matter

2.2: States of Matter

2.3: Classifying of Properties of Matter

2.4: The Periodic Table

2.5: Atoms, Isotopes, and Ions

2.6: Periodic Trends

2.7: Compounds

2.8: Solutions

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2.1: Classification of Matter

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 2.1.1). 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 it's the same everywhere) and properties that are constant throughout the sample (meaning there is only one set of properties such as melting point, color, boiling point, etc. throughout the matter). A material composed of two or more substances is a **mixture**. A mixture is a physical blend of two or more components, each of which retains its own identity and properties in the mixture.

Pure Substances

Elements and compounds are both examples of pure substances. A pure substance that cannot be broken down into chemically simpler components is an **element**. Aluminum, which is used in soda cans, is an element. A pure 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 that can be broken down into 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.

Two additional terms often arise when referring to the makeup of pure substances. An **atom** is the smallest unit of an element that retains its properties, whereas a **molecule** consists of two or more atoms bound together and the atoms may or may not be the same element. For now we will say that all compounds are made up of molecules (although it is a bit more complicated for compounds made up of ions which will be discussed later), but the atoms that make the molecules of a compound cannot all be the same element otherwise it would not be a compound. An element may be made up of atoms or molecules, but unlike molecules of a compound, molecules of an element must be made up of atoms of the same element.

Mixtures

Table salt easily dissolves in water forming a mixture referred to as salt water. Ordinary table salt is called sodium chloride. It is considered a pure substance because it has a uniform and definite composition (all samples of sodium chloride are chemically identical), and it is a compound because it contains more than one element, specifically sodium and chlorine. Salt water cannot be classified as a pure substance because its composition varies; you may dissolve a small amount or a large amount of table salt into a given amount of water, resulting in mixtures that differ in composition. However, it is important to note that only the form of the salt is changed when it is dissolved into water, it does not react with water forming something new, it simply dissolves but retains its composition and properties.

A mixture can be homogeneous or heterogeneous. 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 a pure substance is always the same, whereas the composition of a homogeneous mixture can vary from one sample to another (the proportion of salt and water can be different in different samples of salt water). 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.

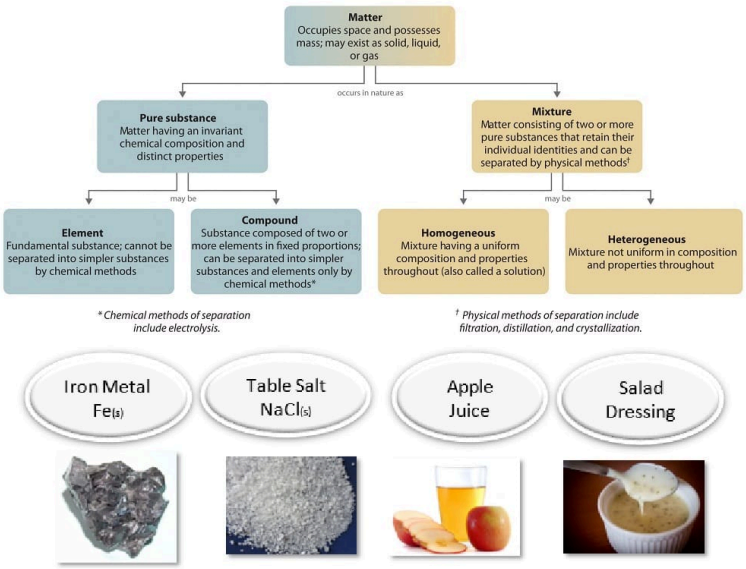


Figure 2.1.3: Relationships between the types of matter and the methods used to separate mixtures.

Example 2.1.2

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

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 2.1.2

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

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

Answer a: available at the end of this section.

Example 2.1.3

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.
b. Soil is composed of small pieces of a variety of materials, so it is a heterogeneous mixture.
c. Water is a pure substance; more specifically, because water is composed of hydrogen and oxygen, it is a compound.
d. Oxygen, a pure substance, is an element.

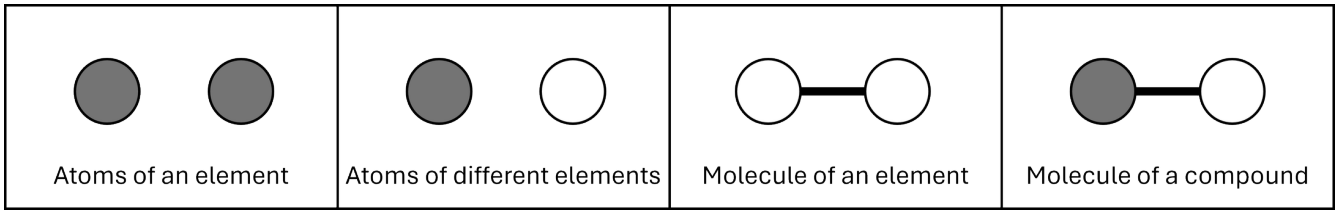
Exercise 2.1.3

How would a chemist categorize each example of matter?

a. filtered coffee
b. hydrogen
c. an egg

Answer a: available at the end of this section.

Often when trying to better understand the matter we are talking about we use pictorial representations of the microscopic particles. In these representations, circles (or spheres) are used to represent atoms, and circles (or spheres) that are linked or merged together represent molecules. Colors or shading are used to represent atoms of different elements. Thus a circle represents an atom of an element, multiple circles with the same shading represent multiple atoms of the same element, and multiple circles with different shading represent atoms of different elements. Circles with the same shading that are attached together represent molecules of an element, but circles with different shading that are attached together represent molecules of a compound.



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.
- Elements and compounds are both examples of pure substances.
- A pure substance can be made up of atoms or molecules.
- Compounds are pure substances that are made up of more than one type of atom, whereas elements are the simplest pure substances made up of only one type of atom.
- Mixtures are physical combinations of two or more elements and/or compounds.
- Mixtures can be classified as homogeneous or heterogeneous.
- Microscopic representations of the particles that make up matter use circles to represent atoms, connected circles to represent molecules, and different shading for different elements.

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Answers

Exercise 2.1.2

homogeneous mixture
element
heterogeneous mixture
compound

Exercise 2.1.3

a homogeneous mixture
element
heterogeneous mixture

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2.2: States of Matter

The States of Matter

Matter typically exists in one of three states: **solid**, **liquid**, or **gas**. Some substances exist as gases at room temperature (such as oxygen and carbon dioxide); others, like water and ethanol, exist as liquids. Most metals exist as solids at room temperature. All substances can exist in any of these three states.





Figure 2.2.1: Matter is usually classified into one of three states, with plasma sometimes added as a fourth state. (From left to right:) Quartz (solid), water (liquid), nitrogen dioxide (gas).

Note

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.

Solid

Solids are defined by the following characteristics:

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

If we heat a solid to its melting point and continue heating it, then it will be converted to a liquid.

Liquid

Liquids have the following characteristics:

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

If we cool a liquid to its freezing point and continue to cool it, then it will convert to a solid. Note: The freezing point of a liquid is the same temperature as its melting point in the solid phase. Furthermore, if we heat a liquid to its boiling point and continue to heat it, then it will be converted to a gas.

Gas

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

If we cool a gas to its condensation point and continue to cool it, then it will be converted to a liquid. Note: The condensation point of a gas is the same temperature as its boiling point in the liquid phase.

The characteristics of the three states of matter are listed in Table 2.2.1 and an animation of the movement and position of the individual particles is shown in Figure 2.2.2

Table 2.2.1 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

Table 2.2.2 images of the states of matter.

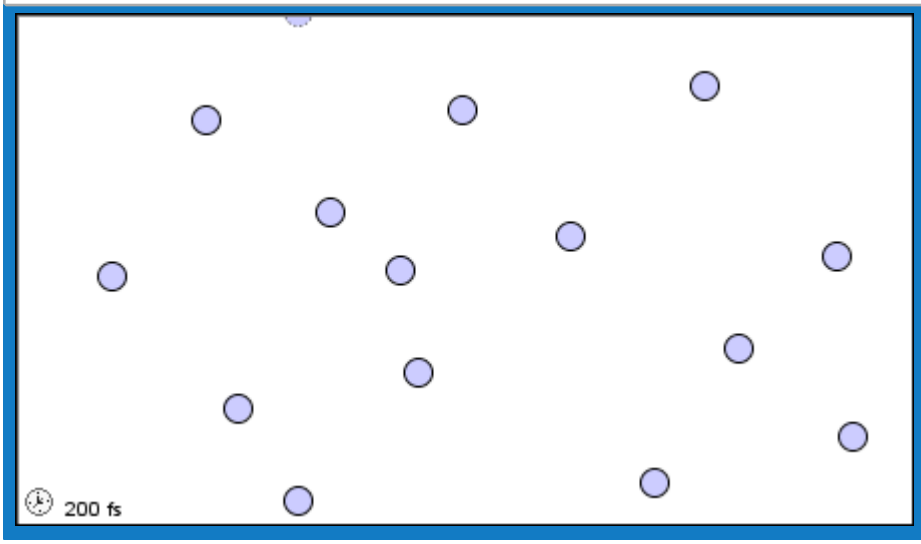
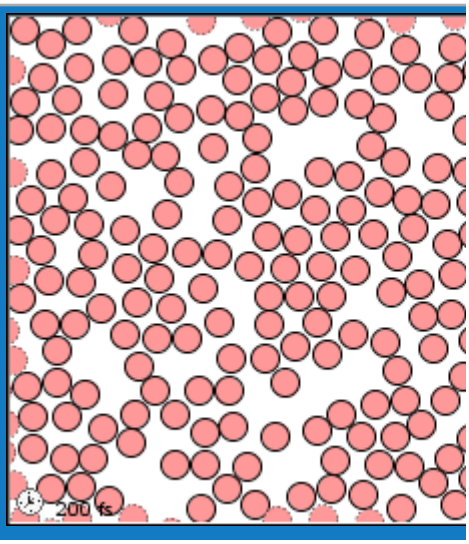
Imag	
The gaseous state	*The liqui*
	

Figure *2.2.2 A microscopic model showing particles (atoms or molecules) in the gaseous, liquid, and solid states.*

Example 2.2.1

What state(s) of matter does each statement describe?

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

Solution

a. This statement describes the liquid state.
b. This statement describes the gas state.
c. This statement describes the liquid state.

Exercise 2.2.1

What state or states of matter does each statement describe?

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

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

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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- Henry Agnew (UC Davis)
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Answers

Exercise 2.2.1

solid

Exercise 2.2.1

gas

Exercise 2.2.1

solid

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2.3: Classifying of Properties of Matter

As we study matter we observe its properties, and these properties can be categorized in a variety of ways. This includes categorizations related to whether the property is described by words or numbers, whether it depends on sample amount, and whether observation of the property affects the identity of the sample.

Qualitative vs. Quantitative Properties

Qualitative properties describe qualities of a sample that do not rely on numbers, such as color, physical state, texture, and odor. For example, water is a clear, colorless liquid and mercury is a shiny, silver liquid, each example providing information about the the qualitative properties color and physical state.

Quantitative properties are describe a quantity using numbers, such as mass, volume, temperature, and density. For example, a particular sample of water at 4°C has a mass of 5.00 grams and a volume of 5.00 mL. This example provides information about the quantitative properties temperature, mass, and volume for a sample of water.

Intensive vs. Extensive Properties

An **intensive property** is one that does not depend on the amount of a substance, it is the same no matter how much of the sample is considered. For instance, physical state and color of a sample do not depend on how much of the sample you observe. If you have a gallon of water which is a colorless liquid, if you pour a portion of that water into a glass the portion will also by a colorless liquid. Temperature is another example of an intensive property; if you were to put a thermometer in a gallon of water to measure its temperature, pour a portion of the water into a glass and measure the temperature, the temperature will be the same not a fraction of the temperature since it is only a fraction of the water.

An **extensive property** is one that depends on the amount of a substance; if you take a larger portion of a sample an extensive property will be larger, or smaller if you take a smaller portion of the sample. For example, if you measure the mass of a block of cheese and then cut it in half and measure the mass of half of the block of cheese, the masses will be different, specifically, the mass of half the block of cheese will be half the mass of the full block of cheese (of course, this assumes you cut it perfectly in half). Another common extensive property is volume; the more you have of a sample, the larger the space it will occupy.

Physical vs. Chemical Properties

A **physical property** is a characteristic of a substance that can be observed or measured without changing the identity of the substance. Physical properties of matter include mass, volume, color, hardness, malleability, solubility, electrical conductivity, density, melting point, and boiling point. For example, if you examine a piece of silver you could observe that it is a shiny metal that conducts electricity very well, and it can be molded into thin sheets—a property called malleability. Or if you examine a sample of salt you could observe that it is dull and brittle and conducts electricity when it has been dissolved into water, which it does quite easily. Each of these observations describes a physical property because the observation of the property does not change the identity of the substance.

A chemical property of matter describes its "potential" to undergo some chemical change or reaction by virtue of its composition (what elements, electrons, and bonding are present to give the potential for chemical change). 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 acids—this is a chemical property. Zinc reacts with hydrochloric acid to produce hydrogen gas—this is a chemical property. A chemical property of iron is that it is capable of combining with oxygen to form iron oxide, the chemical name of rust (Figure 2.3.2). 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.



Figure 2.3.2 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;).

Example 2.3.1

Which of the following is a chemical property of iron? Which are quantitative properties?

1. Iron corrodes in moist air.
2. Density = 7.874 g/cm^3
3. Iron is soft when pure.
4. Iron melts at 1808 K.

Solution

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

Density = 7.874 g/cm^3 and iron melts at 1808 K are quantitative properties.

Exercise 2.3.1A

Which of the following is a physical property of matter? Which are extensive properties?

1. corrosiveness
2. density
3. flammability

Answer available at the end of this section.

Exercise 2.3.1B

Which of the following is a chemical property? Which are qualitative properties?

1. flammability
2. melting point
3. boiling point
4. density

Answer available at the end of this section.

Summary

- A qualitative property is described by words, whereas a quantitative property is described by numbers.
- Intensive properties do not change with the amount of sample, but extensive properties increase or decrease with the amount of sample.
- 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.

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Answers

Exercise $2.3.1A$

Density is a physical property.

None of the given are extensive properties; if a substance that is corrosive or flammable it will be no matter how much you have, and the density of a substance is the same no matter how much of the sample you have.

Exercise $2.3.1B$

Flammability is a chemical property and a qualitative property. All of the other properties are physical and quantitative properties.

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2.4: The Periodic Table

Learning Objectives

- Explain how elements are organized into the periodic table.
- Describe how some characteristics of elements relate to their positions on the periodic table.
- Know the classification of the different elements.
- Define valence electrons.
- Describe how the number of valence electrons are related the arrangement of the main group elements and its chemical behavior.

The modern periodic table has evolved through a long history of attempts by chemists to arrange the elements according to their properties, as an aid in predicting chemical behavior. The periodic table achieved its modern form through the work of the German chemist Julius Lothar Meyer (1830–1895) and the Russian chemist Dimitri Mendeleev (1834–1907), both of whom focused on the relationships between atomic mass and various physical and chemical properties. In 1869, they independently proposed essentially identical arrangements of the elements. Mendeleev, who first published his periodic table in 1869 (Figure 2.4.1), is usually credited with the origin of the modern periodic table. The key difference between his arrangement of the elements, and that of Meyer and others, is that Mendeleev did not assume that all the elements had been discovered (in fact, only about two-thirds of the naturally occurring elements were known at the time). Instead, he deliberately left blanks in his table at atomic masses 44, 68, 72, and 100—in the expectation that elements with those atomic masses would be discovered. Those blanks correspond to the elements we now know as scandium, gallium, germanium, and technetium. Until roughly 1960, a rectangular table based on reactivity—and developed from Mendeleev's table—was standard at the front of chemistry lecture halls.

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

Figure 2.4.1: **Mendeleev's Periodic Table, as Published in the German Journal *Annalen der Chemie und Pharmacie* in 1872.** The column headings "Reihen" and "Gruppe" are German for "row" and "group." Formulas indicate the type of compounds formed by each group, with "R" standing for "any element" and superscripts used where we now use subscripts. Atomic masses are shown after equal signs and increase across each row from left to right.

The most convincing evidence in support of Mendeleev's arrangement of the elements was the discovery of two previously unknown elements whose properties closely corresponded with his predictions (Table 2.4.1). Two of the blanks Mendeleev had left in his original table were below aluminum and silicon, awaiting the discovery of two as-yet-unknown elements, *eka*-aluminum and *eka*-silicon (from the Sanskrit *eka*, meaning "one," as in "one beyond aluminum"). The observed properties of gallium and germanium matched those of *eka*-aluminum and *eka*-silicon so well that once they were discovered, Mendeleev's periodic table rapidly gained acceptance.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 2.4.1: The genius of Mendeleev's periodic table.

One group of elements that was absent from Mendeleev's table is the noble gases, all of which were discovered more than 20 years later—between 1894 and 1898—by Sir William Ramsay (1852–1916; Nobel Prize in Chemistry 1904). Initially, Ramsay did not know where to place these elements in the periodic table. Argon, the first to be discovered, had an atomic mass of 40. This was greater than chlorine's and comparable to that of potassium; so Ramsay, using the same kind of reasoning as Mendeleev, decided to place the noble gases between the halogens and the alkali metals. In 1913, however, young British physicist H. G. J. Moseley (1887–1915) analyzed the frequencies of x-rays emitted by the elements, and discovered that the underlying foundation of the order of the elements was by the *atomic number*—not the atomic mass. Moseley hypothesized that the placement of each element in his series corresponded to its atomic number *Z*, which is the number of positive charges (protons) in its nucleus. Finally, in 1945, the Manhattan Project yielded the discovery of many new radioactive elements. Glenn T. Seaborg suggested an addition of the actinide and lanthanide series at the bottom of the table. This idea came with the discovery of Americium and Curium, and their unique properties. The change was not accepted at first, but is now included in all periodic tables. Figure 2.4.2 shows the time period for the discovery of the different elements.

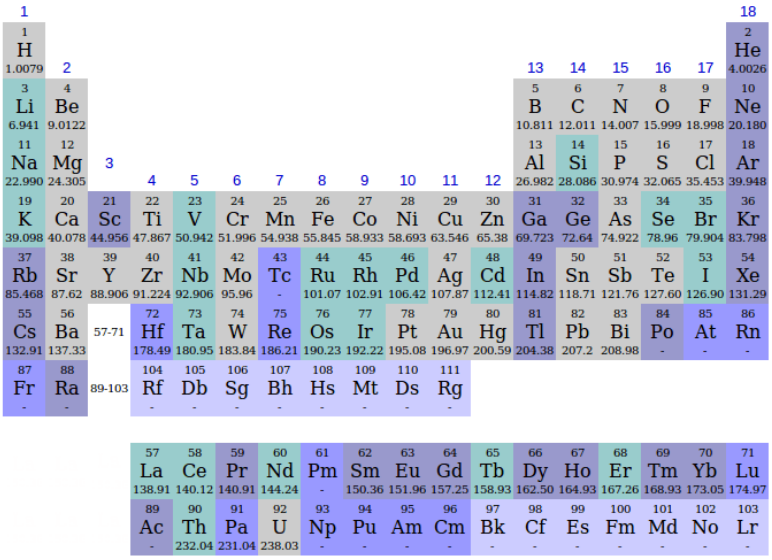


Figure 2.4.2 *Periodic table showing when each element was discovered*

- Before 1800 (36 elements): Discoveries during and before the Age of Enlightenment.
- 1800-1849 (+22 elements): Impulse from Scientific Revolution and Atomic theory and Industrial Revolution.
- 1850-1899 (+23 elements): The age of Classifying Elements received an impulse from the Spectrum analysis.
- 1900-1949 (+13 elements): Impulse from the old quantum theory, the Refinements to the periodic table, and quantum mechanics.
- 1950-1999 (+15 elements): Manhattan Project and Particle physics issues, for atomic numbers 97 and above.

Each element on the modern periodic table is represented by a one or two letter symbol, and often are accompanied by their **atomic number** and **atomic mass** (these will be discussed in further detail in later sections of the text). In the example shown below (Figure 2.4.3), the atomic mass is found under the element symbol and the atomic number (Z) is found above the atomic symbol.

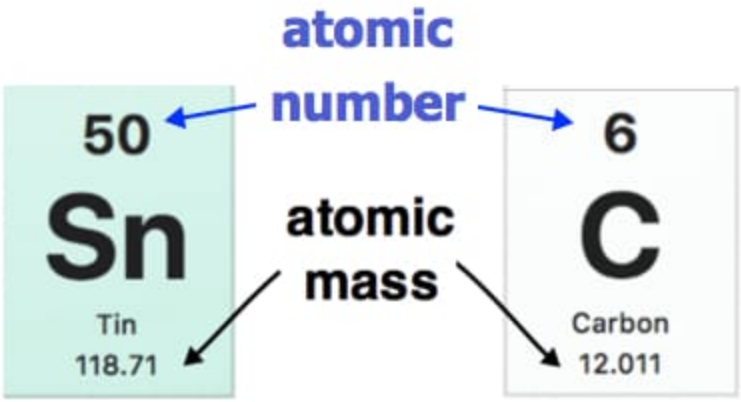


Figure 2.4.3: Key information of an element indicated on entries of the Periodic Table. (public Domain; Pubchem)

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, the elements in the first column are called **alkali metals** , the elements in the second column are called **alkaline earth metals**, the elements in the next-to-last column are called **halogens**, and the elements in the last column are called **noble gases**. The **main group elements** are the elements in groups 1 and 2, and groups 13 to 18; they are sometimes also called the **representative elements**.

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.

One way to categorize the elements of the periodic table is shown in Figure 2.4.1. The first two columns on the left and the last six columns on the right as mentioned earlier are 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**.

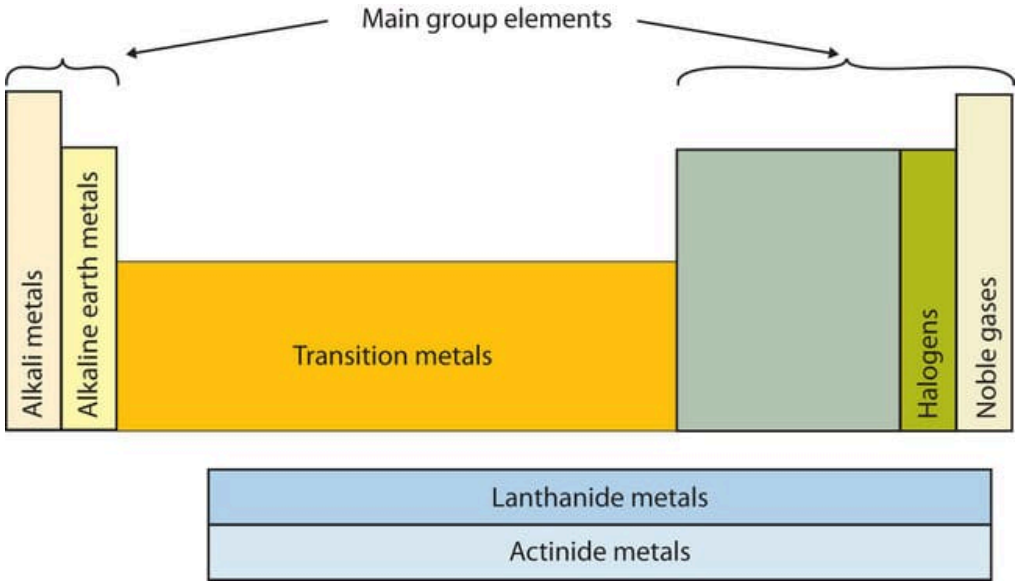


Figure 2.4.1: Special names for sections of the periodic table.

Example 2.4.1: 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 2.4.1: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a: available at the end of this section.

Metals and Nonmetals

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 semimetal), as shown in Figure 2.4.2. A **metal** is a substance that is shiny, typically silvery in color (but not always), 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 2.4.2, 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.

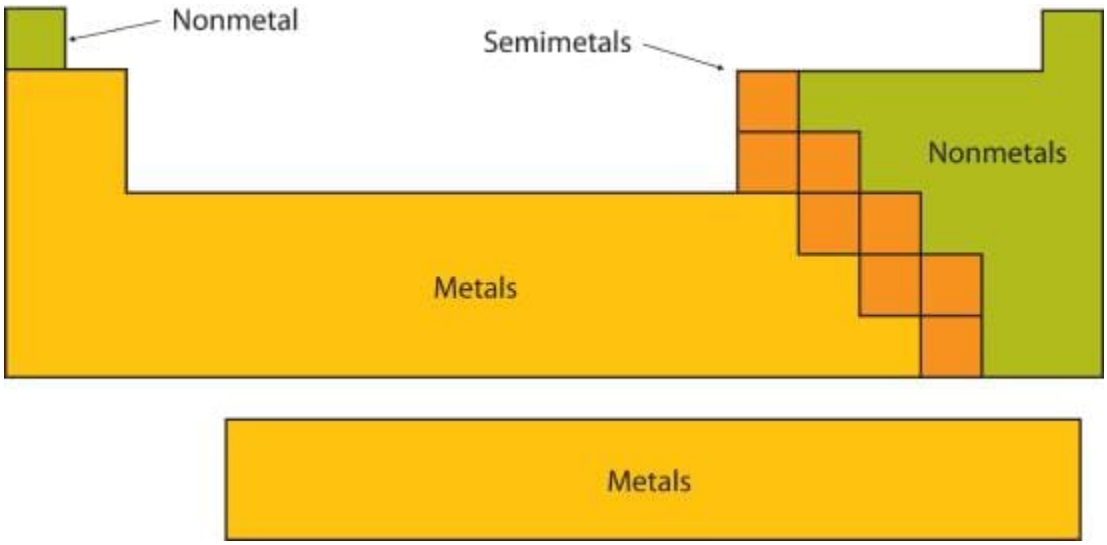


Figure 2.4.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table.

Example 2.4.2

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

a. Se
b. Mg
c. Ge

Solution

a. In Figure 2.4.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

Exercise 2.4.2

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

Answer available at the end of this section.

Example 2.4.3: 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 2.4.3: Classification of Elements

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

a. F
b. U
c. Cu

Answer a: available at the end of this section.

Summary

- Elements in the periodic table are arranged by 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.
- 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).
- The elements can be broadly divided into metals, nonmetals, and semi metals. 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 semimetals that exhibit properties intermediate between those of metals and nonmetals.
- Metals are lustrous, good conductors of electricity, and readily shaped (they are ductile and malleable). Solid nonmetals are generally brittle and poor electrical conductors.

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
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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- Wikipedia

Answers

Exercise 2.4.1**: Groups**

Fluorine is a halogen (Group 17)
Calcium is a alkaline earth metal (Group 2)
Krypton is a noble gas (Group 18)

Exercise 2.4.2

Indium is a metal.

Exercise 2.4.3**: Classification of Elements**

Fluorine is a nonmetal.
Uranium is a metal (and a inner transition metal too)
Copper is a metal (and a transition metal too)

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2.5: Atoms, Isotopes, and Ions

Learning Objectives

- State the three parts of modern atomic theory.
- Learn how atoms are constructed.
- Know how ions form.
- Learn the characteristic charges that ions have

The smallest piece of an element that maintains the identity of that element is called an atom. Individual atoms are extremely small. It would take about fifty million atoms in a row to make a line that is 1 cm long. The period at the end of a printed sentence has several million atoms in it. Atoms are so small that it is difficult to believe that all matter is made from atoms—but it is.

The concept that atoms play a fundamental role in chemistry is formalized by the modern atomic theory. This theory is based on the atomic theory first proposed by John Dalton, an English scientist, in 1808. According to the modern atomic theory:

1. All matter is composed of atoms.
2. An atom contains a dense, positive nucleus with protons and neutrons, surrounded by negative electrons (Figure 2.5.1).
3. The identity of an atom is defined by its atomic number (the number of protons in its nucleus).
4. Atoms of different elements combine in whole-number ratios to form compounds.

These concepts form the basis of chemistry. Although the word *atom* comes from a Greek word that means "indivisible," we understand now that atoms themselves are composed of smaller parts called *subatomic particles*. The first part to be discovered was the electron, a tiny subatomic particle with a negative charge. It is often represented as e^-, with the right superscript showing the negative charge. Later, two larger particles were discovered. The proton is a more massive subatomic particle (but still tiny) with a positive charge, represented as p^+. The neutron is a subatomic particle with about the same mass as a proton, but no charge. It is represented as either n or n^0. We now know that all atoms of all elements are composed of electrons, protons, and neutrons (with one exception). Table 2.5.1 summarizes the properties of these three subatomic particles.

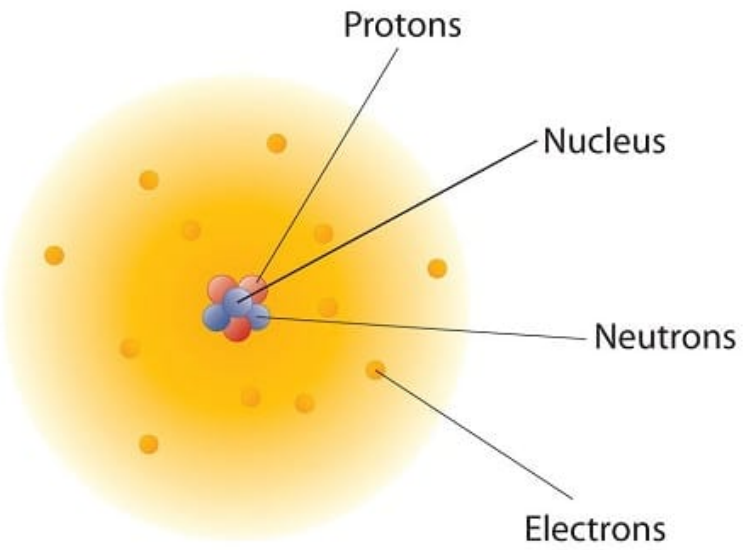


Figure 2.5.1: The Structure of the Atom. Atoms have protons and neutrons in the center, making the nucleus, while the electrons orbit the nucleus.

Table 2.5.1: Properties of the Three Subatomic Particles

Name	Symbol	Approximate Mass (kg)	Relative Charge
Proton	p^+	1.6×10^{-27}	1+
Neutron	n, n^0	1.6×10^{-27}	none
Electron	e^-	9.1×10^{-31}	1–

The fundamental characteristic that all atoms of the same element share is the *number of protons*. All atoms of hydrogen have one and only one proton in the nucleus; all atoms of iron have 26 protons in the nucleus. This number of protons is so important to the identity of an atom that it is called the **atomic number**. The number of protons in an atom is the atomic number of the element. Thus, hydrogen has an atomic number of 1, while iron has an atomic number of 26. Each element has its own characteristic atomic number. However, atoms of the same element can have different numbers of neutrons. Atoms of the same element (i.e., atoms with the same number of protons) with different numbers of neutrons are called **isotopes**. Most naturally occurring elements exist as isotopes. For example, most hydrogen atoms have a single proton in their nucleus, but a small number (about one in a million) of hydrogen atoms have a proton and a neutron in their nucleus. This particular isotope of hydrogen is called deuterium. A very rare form of hydrogen has one proton and two neutrons in the nucleus; this isotope of hydrogen is called tritium. The sum of the number of protons and neutrons in the nucleus is called the **mass number** of the isotope.

Neutral atoms have the same number of electrons as they have protons, so their overall charge is zero. However, as we shall see later, this will not always be the case.

Example 2.5.1

a. The most common carbon atoms have six protons and six neutrons in their nuclei. What are the atomic number and the mass number of these carbon atoms?
b. An isotope of uranium has an atomic number of 92 and a mass number of 235. What are the number of protons and neutrons in the nucleus of this atom?

Solution

a. If a carbon atom has six protons in its nucleus, its atomic number is 6. If it also has six neutrons in the nucleus, then the mass number is 6 + 6, or 12.
b. If the atomic number of uranium is 92, then that is the number of protons in the nucleus. Because the mass number is 235, then the number of neutrons in the nucleus is 235 – 92, or 143.

Exercise 2.5.1

The number of protons in the nucleus of a tin atom is 50, while the number of neutrons in the nucleus is 68. What are the atomic number and the mass number of this isotope?

Answer available at the end of this section.

When referring to an atom, we simply use the element's name: the term *sodium* refers to the element as well as an atom of sodium. But it can be unwieldy to use the name of elements all the time. Instead, chemistry defines a symbol for each element. The atomic symbol is a one- or two-letter representation of the name of an element. By convention, the first letter of an element's symbol is always capitalized, while the second letter (if present) is lowercase. Thus, the symbol for hydrogen is H, the symbol for sodium is Na, and the symbol for nickel is Ni. Most symbols come from the English name of the element, although some symbols come from an element's Latin name. For example, the symbol for sodium is Na, which comes from its Latin name, *natrium*. Table 2.5.2 lists some elements names and symbols that you will need to memorize.

Table 2.5.2: Names and Symbols of Common Elements

Element Name	Symbol	Element Name	Symbol
Aluminum	Al	Lithium	Li
Antimony	Sb	Magnesium	Mg
Argon	Ar	Manganese	Mn
Arsenic	As	Mercury	Hg
Barium	Ba	Neon	Ne
Beryllium	Be	Nickel	Ni
Bismuth	Bi	Nitrogen	N
Boron	B	Oxygen	O
Bromine	Br	Phosphorus	P
Cadmium	Cd	Potassium	K
Calcium	Ca	Rubidium	Rb
Carbon	C	Scandium	Sc
Cesium	Cs	Selenium	Se
Chlorine	Cl	Silicon	Si
Chromium	Cr	Silver	Ag
Cobalt	Co	Sodium	Na
Copper	Cu	Strontium	Sr
Fluorine	F	Sulfur	S
Gold	Au	Tellurium	Te
Helium	He	Vanadium	V
Hydrogen	H	Tin	Sn
Iodine	I	Titanium	Ti
Iron	Fe	Xenon	Xe
Krypton	Kr	Zinc	Zn
Lead	Pb		

Often when we are dealing with elements or compounds the atoms involved are a mixture of various isotopes formed by each element, but in some cases we will need to focus on a specific isotope of an element. To do this we must not only indicate the element using either the name or symbol, but we must also indicate the number of neutrons which is different for different isotopes of an element. This is typically done indirectly by reporting the mass number of the isotope. Recall, the mass number is the number of protons and neutrons in an atom, and if we know the element then we know how many protons are in the atom which is equivalent to the atomic number. By taking the mass number and subtracting the atomic number we can determine the number of neutrons.

Isotopes can be referred to by there element name and mass number, such as hydrogen-1, carbon-14, chlorine-37, etc. Alternatively, isotopes can be represented using atomic notation:

$${}^{A}_{Z}\mathrm{X}$$

where X is the symbol of the element, A is the mass number, and Z is the atomic number. In some cases the atomic number is left out since it can be determined by looking up the element on the periodic table.

For the isotope of carbon that has 6 protons and 6 neutrons, the symbol is:

$^{12}_{6}\mathrm{C}$ or $^{12}_{6}\mathrm{C}$

where C is the symbol for the element, 6 represents the atomic number, and 12 represents the mass number.

Example 2.5.2

a. What is the symbol for an isotope of uranium that has an atomic number of 92 and a mass number of 235?
b. How many protons and neutrons are in $^{56}_{26}\mathrm{Fe}$

Solution

a. The symbol for this isotope is $^{235}_{92}\mathrm{U}$
b. This iron atom has 26 protons and 56 – 26 = 30 neutrons.

Exercise 2.5.2

How many protons are in $^{23}_{11}\mathrm{Na}$

Answer available at the end of this section.

So far, we have discussed atoms that are electrically neutral. They have the same number of electrons as protons, so the negative charges of the electrons are balanced by the positive charges of the protons. However, this is not always the case. Electrons can move from one atom to another; when they do, species with overall electric charges are formed. Such species are called **ions**. Species with overall positive charges are termed **cations**, while species with overall negative charges are called **anions**. Remember that ions are formed only when *electrons* move from one atom to another; a proton never moves from one atom to another.

Individual atoms can gain or lose electrons. When they do, they become **monatomic ions**. Each element that forms a cation is a metal, except for one (hydrogen), while each element that forms an anion is a nonmetal. This is actually one of the chemical properties of metals and nonmetals: metals tend to form cations, while nonmetals tend to form anions. Many atoms form ions of a single characteristic charge which can be determined from their location on the periodic table. For example, group I metals always form cations with a 1+ charge due to the loss of 1 electron, whereas group II metals always form cations with a 2+ charge due to the loss of 2 electrons. Nonmetals in group 17 always form a charge of 1- due to gaining 1 electron, while those in group 16 form a 2- and those in group 15 form a 3-. Each of these charges formed corresponds to the atom losing or gaining electrons to become like the nearest noble gas (having the same number of electrons as the nearest noble gas). There are some exceptions to the previous point. Some elements, all of which are metals, can form more than one possible charge. For example, iron atoms can form 2+ cations or 3+ cations. Cobalt is another element that can form more than one possible charged ion (2+ and 3+), while lead can form 2+ or 4+ cations. The charges formed by these elements can sometimes be determined based on patterns related to the periodic table, but the reasoning behind the charges formed by them is associated with details of electronic structure that are beyond the scope of this course.

The convention for indicating an ion is to write the charge as a superscript to the right of the element symbol. If the charge is a single positive or negative one, the number 1 is not written; if the magnitude of the charge is greater than 1, then the number is written *before* the + or – sign. An element symbol without a charge written next to it is assumed to be the uncharged atom.

Key Takeaways

- Chemistry is based on the modern atomic theory, which states that all matter is composed of atoms.
- Atoms themselves are composed of protons, neutrons, and electrons.
- Each element has its own atomic number, which is equal to the number of protons in its nucleus.
- Isotopes of an element contain different mass numbers due to a differing number of neutrons.
- In order to specify an isotope the mass number is provided with the element name or as a part of atomic notation.
- Ions form when atoms lose or gain electrons.

Answers

Exercise 2.5.1

Atomic number = 50, mass number = 118

Exercise 2.5.2

11 protons

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2.6: Periodic Trends

Learning Objective

- Be able to state how certain properties of atoms vary based on their relative position on the periodic table.

One reason the periodic table is so useful is that its structure allows us to qualitatively determine how some properties of the elements vary according to their position on the periodic table. The variations of properties according to positions on the periodic table are called **periodic trends**. There is no other tool in science that allows us to judge relative properties of a class of objects like this, which makes the periodic table a very useful tool. Many periodic trends are general. There may be a few points where an opposite trend is seen, but there is an overall trend when considered across a whole row, or down an entire column of the periodic table.

The first periodic trend we will consider is atomic radius. The **atomic radius** is an indication of the size of an atom. Although the concept of a definite radius of an atom is a bit fuzzy, atoms behave as if they have a certain radius. Such radii can be estimated from various experimental techniques, such as the x-ray crystallography of crystals.

As you go down a column of the periodic table, the atomic radii increase. This is because the valence electron shell is getting larger and there is a larger principal quantum number, so the valence shell lies physically farther away from the nucleus. This trend can be summarized as follows:

\[as\downarrow PT,atomic\; radius \uparrow \]

where PT stands for periodic table. Going across a row on the periodic table, left to right, the trend is different. This is because although the valence shell maintains the same principal quantum number, the number of protons—and hence the nuclear charge—is increasing as you go across the row. The increasing positive charge casts a tighter grip on the valence electrons, so as you go across the periodic table, the atomic radii decrease. Again, we can summarize this trend as follows:

$$as \rightarrow PT, atomic\ radius \downarrow \tag{2.6.1}$$

Figure 2.6.1 shows spheres representing the atoms of the *s* and *p* blocks from the periodic table to scale, showing the two trends for the atomic radius.

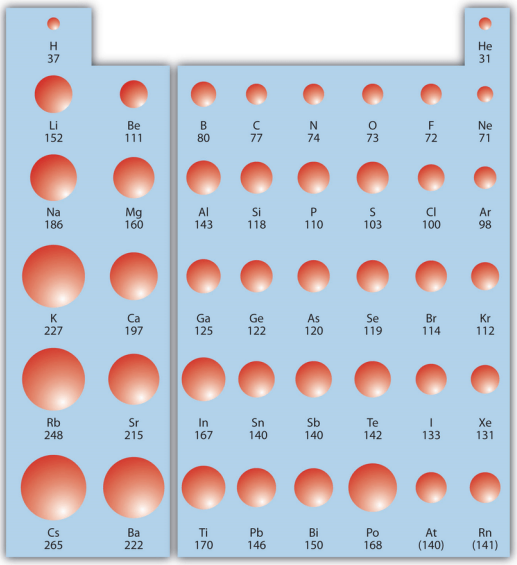


Figure 2.6.1: Atomic Radii Trends on the Periodic Table. Although there are some reversals in the trend (e.g., see Po in the bottom row), atoms generally get smaller as you go across the periodic table and larger as you go down any one column. Numbers are the radii in pm.

Example 2.6.1: Atomic Radii

Referring only to a periodic table and not to Figure 2.6.1, which atom is larger in each pair?

a. Si or S
b. S or Te

Solution

a. Si is to the left of S on the periodic table, so it is larger because as you go across the row, the atoms get smaller.
b. S is above Te on the periodic table, so Te is larger because as you go down the column, the atoms get larger.

Exercise 2.6.1: Atomic Radii

Referring only to a periodic table and not to Figure 2.6.1, which atom is smaller, Ca or Br?

Answer

Br

Ionization energy (IE) is the amount of energy required to remove an electron from an atom in the gas phase:

$$A(g) \rightarrow A^{+}(g) + e^{-} \qquad \Delta H \equiv IE$$

IE is usually expressed in kJ/mol of atoms. It is always positive because the removal of an electron always requires that energy be put in (i.e., it is endothermic). IE also shows periodic trends. As you go down the periodic table, it becomes easier to remove an electron from an atom (i.e., IE decreases) because the valence electron is farther away from the nucleus. Thus,

$$as \downarrow PT,\ IE \downarrow \tag{2.6.2}$$

However, as you go across the periodic table and the electrons get drawn closer in, it takes more energy to remove an electron. As a result, IE increases:

$$as \rightarrow PT,\ IE \uparrow \tag{2.6.3}$$

Figure 2.6.2 shows values of IE versus position on the periodic table. Again, the trend isn't absolute, but the general trends going across and down the periodic table should be obvious.

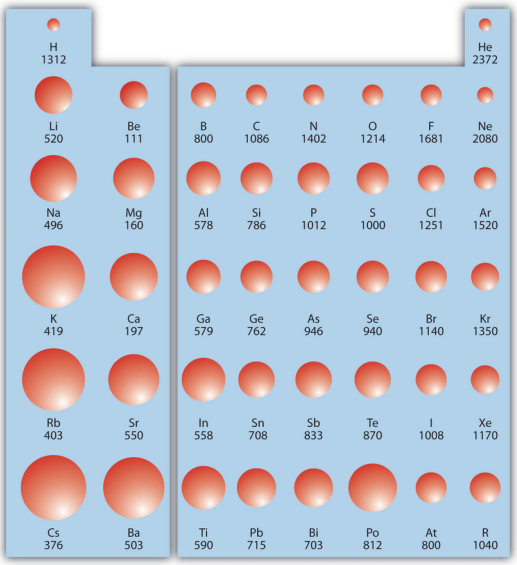


Figure 2.6.2: Ionization Energy on the Periodic Table. Values are in kJ/mol.

IE also shows an interesting trend within a given atom. This is because more than one IE can be defined by removing successive electrons (if the atom has them to begin with):

- First Ionization Energy (IE_1):

$$A(g) \rightarrow A^+(g) + e^- \tag{2.6.4}$$

- Second Ionization Energy (IE_2):

$$A^+(g) \rightarrow A^{2+}(g) + e^- \tag{2.6.5}$$

- Third Ionization Energy (IE_3):

$$A^{2+}(g) \rightarrow A^{3+}(g) + e^- \tag{2.6.6}$$

and so forth.

Each successive IE is larger than the previous because an electron is being removed from an atom with a progressively larger positive charge. However, IE takes a large jump when a successive ionization goes down into a new shell. For example, the following are the first three IEs for Mg, whose electron configuration is $1s^2 2s^2 2p^6 3s^2$:

First Ionization Energy (IE_1) = 738 kJ/mol:

$$Mg(g) \rightarrow Mg^+(g) + e^- \tag{2.6.7}$$

Second Ionization Energy (IE_2) = 1,450 kJ/mol:

$$Mg^+(g) \rightarrow Mg^{2+}(g) + e^- \tag{2.6.8}$$

Third Ionization Energy (IE_3) = 7,734 kJ/mol:

$$Mg^{2+}(g) \rightarrow Mg^{3+}(g) + e^- \tag{2.6.9}$$

The second IE is twice the first, which is not a surprise: the first IE involves removing an electron from a neutral atom, while the second one involves removing an electron from a positive ion. The third IE, however, is over *five times* the previous one. Why is it so much larger? Because the first two electrons are removed from the 3*s* subshell, but the third electron has to be removed from the $n = 2$ shell (specifically, the 2*p* subshell, which is lower in energy than the $n = 3$ shell). Thus, it takes much more energy than just overcoming a larger ionic charge would suggest. Trends like this one demonstrate that electrons within atoms are organized in groups.

Example 2.6.2: Ionization Energies

Which atom in each pair has the larger first ionization energy?

a. Ca or Sr
b. K or K^+

Solution

a. Because Sr is below Ca on the periodic table, it is easier to remove an electron from it; thus, Ca has the higher IE.
b. Because K^+ has a positive charge, it will be harder to remove another electron from it, so its IE is larger than that of K. Indeed, it will be significantly larger because the next electron in K^+ to be removed comes from another shell.

Exercise 2.6.2: Ionization Energies

Which atom has the lower ionization energy, C or F?

Answer

C

The opposite of IE is described by **electron affinity (EA),** which is the energy change when a gas-phase atom accepts an electron:

$$A(g) + e^- \rightarrow A^-(g) \qquad \Delta H \equiv EA \tag{2.6.10}$$

EA is also usually expressed in kJ/mol. EA also demonstrates some periodic trends, although they are less obvious than the other periodic trends discussed previously. Generally, as you go across the periodic table, EA increases its magnitude:

$$as \rightarrow PT,\ EA \uparrow \tag{2.6.11}$$

There is not a definitive trend as you go down the periodic table; sometimes EA increases, sometimes it decreases. Figure 2.6.3 shows EA values versus position on the periodic table for the *s*- and *p*-block elements. The trend isn't absolute, especially considering the large positive EA values for the second column. However, the general trend going across the periodic table should be obvious.

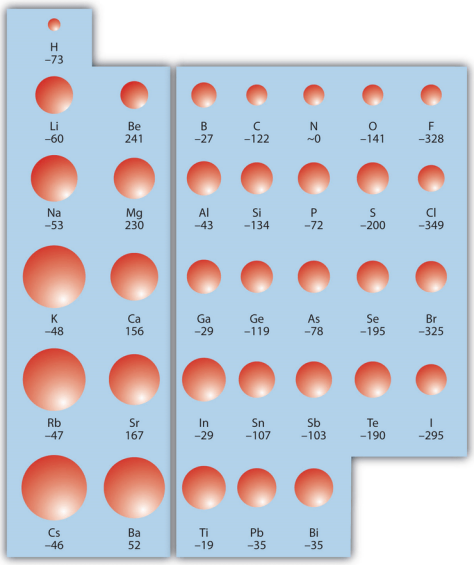


Figure 2.6.3: Electron Affinity on the Periodic Table. Values are in kJ/mol.

Example 2.6.3: Electron Affinities

Predict which atom in each pair will have the highest magnitude of Electron Affinity.

a. C or F
b. Na or S

Solution

a. C and F are in the same row on the periodic table, but F is farther to the right. Therefore, F should have the larger magnitude of EA.
b. Na and S are in the same row on the periodic table, but S is farther to the right. Therefore, S should have the larger magnitude of EA.

Exercise 2.6.3: Electron Affinities

Predict which atom will have the highest magnitude of Electron Affinity: As or Br.

Answer

Br

Metallic character is used to define the chemical properties that metallic elements present. Generally, metals tend to lose electrons to form cations. Nonmetals tend to gain electrons to form anions. As you move across the table from left to right, the metallic character decreases, because the elements easily accept electrons. Therefore, these elements take on the nonmetallic character of forming anions. As you move up the table, the metallic character decreases, due to the greater pull that the nucleus has on the outer electrons. This greater pull makes it harder for the atoms to lose electrons and form cations.

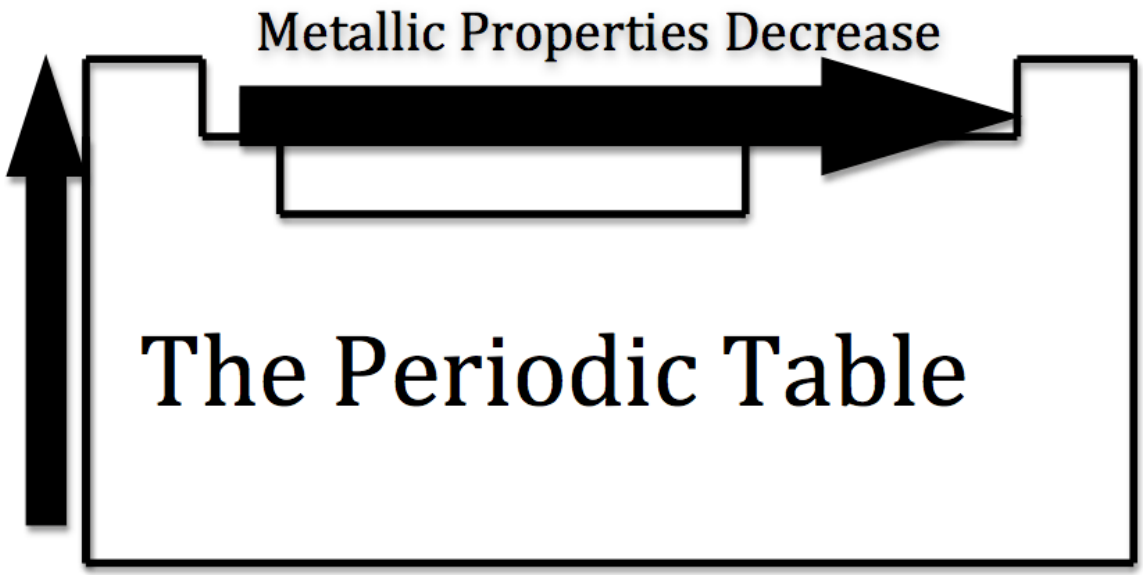


Figure 2.6.4: Courtesy of Jessica Thornton (UCD)

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur) and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

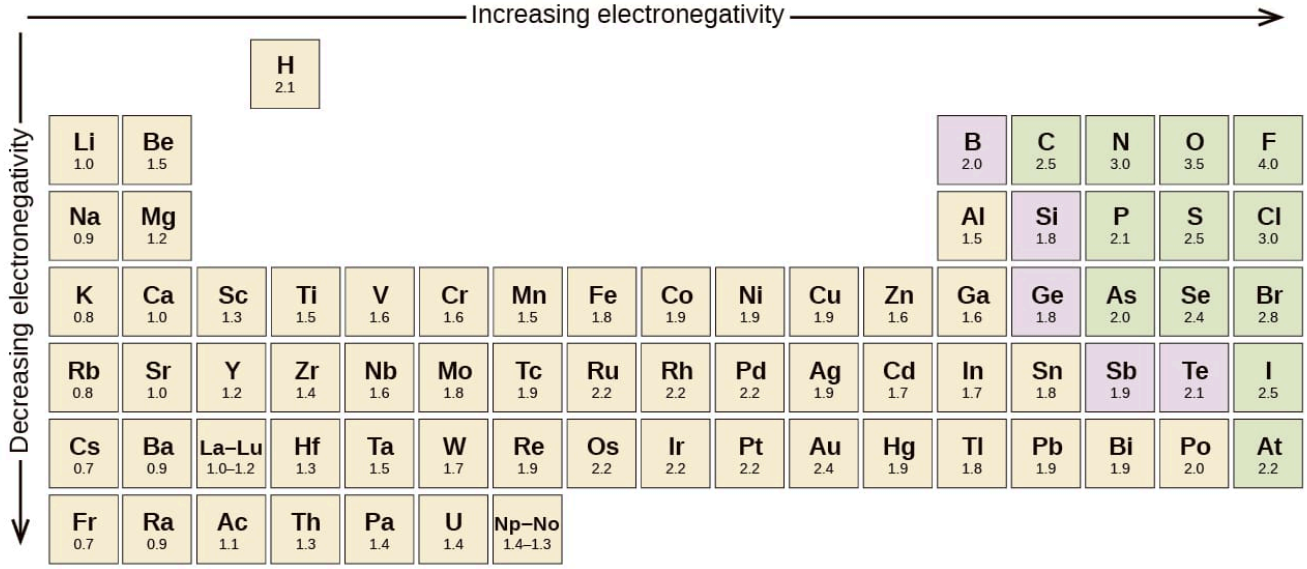


Figure 2.6.1 Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

Contributors and Attributions

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Summary

- Certain properties—notably atomic radius, ionization energies, and electron affinities—can be qualitatively understood by the positions of the elements on the periodic table.

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2.7: Compounds

Compounds are molecules made up of atoms of two or more elements. There are two important laws that tell us more about compounds. One is the **Law of Definite Proportions** which was formulated by Joseph Proust (1754-1826), also called the **Law of Constant Composition** or **Proust's Law**. This law states that if a compound is broken down into its constituent elements, the masses of the constituents will always have the same proportions, regardless of the quantity or source of the original substance. The illustration below depicts this law in action.

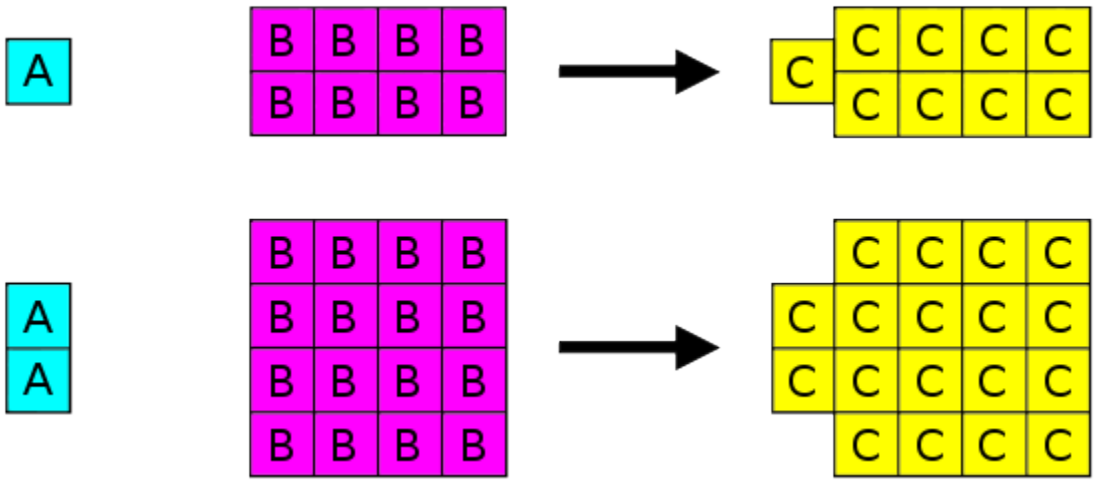


Figure 2.7.4: If 1 gram of A reacts with 8 grams of B, then by the Law of Definite Proportions, 2 grams of A must react with 16 grams of B. If 1 gram of A reacts with 8 grams of B, then by the Law of Conservation of Mass, they must produce 9 grams of C. Similarly, when 2 grams of A react with 16 grams of B, they must produce 18 grams of C.

> Law of Definite Proportions states that in a given type of chemical substance, the elements are always combined in the same proportions by mass.

Example 2.7.2: water

Oxygen makes up 88.8% of the mass of any sample of pure water, while hydrogen makes up the remaining 11.2% of the mass. You can get water by melting ice or snow; or by condensing steam from a river, sea, pond, etc. It can be from different places: USA, UK, Australia, or anywhere. Water can be made by chemical reactions, such as burning hydrogen in oxygen.

However, if the water is **pure**, it will **always** consist of 88.8 % oxygen by mass and 11.2 % hydrogen by mass, irrespective of its source or method of preparation.

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

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Video 2.7.2 Law of definite proportions.

The Law of Definite Proportions applies when elements are reacted together to form *the same* product, however, many combinations of elements can react to form more than one compound. In such cases, the masses of a particular element that combine with a fixed mass of a second element form in ratios of small whole numbers. This is known as the **Law of Multiple Proportions**. For example, nitrogen forms a very large number of oxides, including NO, NO_2, N_2O, N_2O_4, and N_2O_5. When comparing masses of nitrogen and oxygen in NO and NO_2 we find that for every gram of nitrogen there is 1.14 grams of oxygen in

NO whereas NO_2 contains 2.28 grams of oxygen per gram of nitrogen, taking a ratio of the grams of oxygen per gram of nitrogen in these compounds shows the 1:2 ratio of oxygen in the two compounds.

> The law of multiple proportions states that if two elements form more than one compound between them, the masses of one element combined with a fixed mass of the second element form in ratios of small integers.

Example 2.7.1: Law Multiple Proportions

A sample of compound A (a clear, colorless gas) is analyzed and found to contain 4.27 g carbon and 5.69 g oxygen. A sample of compound B (also a clear, colorless gas) is analyzed and found to contain 5.19 g carbon and 13.84 g oxygen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances A and B?

Solution

In compound A, the mass ratio of carbon to oxygen is:

$$\frac{1.33\text{ g O}}{1\text{ g C}}$$

In compound B, the mass ratio of carbon to oxygen is:

$$\frac{2.67\text{ g O}}{1\text{ g C}}$$

The ratio of these ratios is:

$$\frac{\frac{1.33\text{ g O}}{1\text{ g C}}}{\frac{2.67\text{ g O}}{1\text{ g C}}} = \frac{1}{2}$$

This supports the law of multiple proportions. This means that A and B are different compounds, with A having one-half as much carbon per amount of oxygen (or twice as much oxygen per amount of carbon) as B. A possible pair of compounds that would fit this relationship would be A = CO_2 and B = CO.

Exercise 2.7.1

A sample of compound X (a clear, colorless, combustible liquid with a noticeable odor) is analyzed and found to contain 14.13 g carbon and 2.96 g hydrogen. A sample of compound Y (a clear, colorless, combustible liquid with a noticeable odor that is slightly different from X's odor) is analyzed and found to contain 19.91 g carbon and 3.34 g hydrogen. Are these data an example of the law of definite proportions, the law of multiple proportions, or neither? What do these data tell you about substances X and Y?

Answer available at the end of this section.

When two atoms approach each other, they have the potential to bond (or connect). The bonds formed involve the outermost electrons (**valence electrons**) of the atoms and can result in the formation of ionic compounds or covalent compounds.

Ionic Compounds

If a metal and a nonmetal interact, then an ionic bond will result. These types of bonds involve the metal donating its valence electrons to a nonmetal; resulting in a metal cation and nonmetal anion that are then attracted to one another. Both atoms become more stable and a less reactive compound is formed. In addition, they are crystalline and solid in nature (Figure 2.7.1). A few examples of real-world ionic compounds would be NaCl (table salt) and NaF (active ingredient in toothpaste).

Table salt is the result of Na^+ ions and Cl^- ions bonding together (Figure 2.7.2) in a 1:1 ratio which is represented using the formula

NaCl. Formulas of ionic compounds are always reported using the smallest whole number ratio of the cation and anion which is referred to as a **formula unit**, but in reality ionic compounds exist as large clusters of the cations and anions and not discrete units of the ions.



Figure $2.7.1$ *NaCl crystals.*

(Public Domain; NASA).

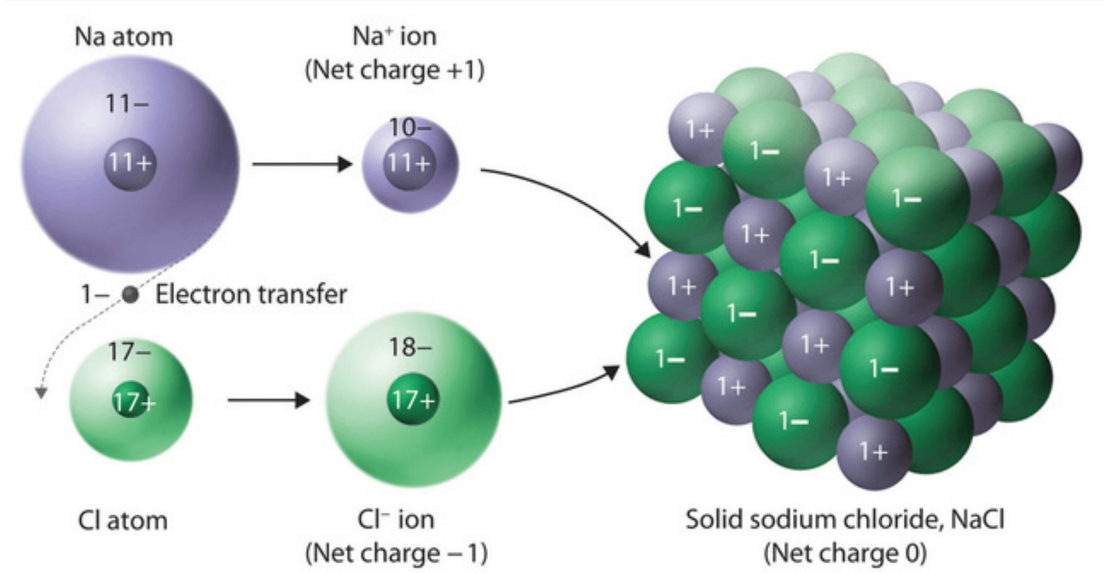


Figure $2.7.2$ The formation of sodium chloride from the attraction of sodium and chloride ions.

Covalent Compounds (Molecular Compounds)

If nonmetals interact, then a covalent bond will result. These types of bonds involve the nonmetals sharing their valence electrons. Molecular compounds 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 $(NaCl)$. Instead of a transfer of electrons forming ions that are electrostatically attracted to each other, 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 $2.7.3$). 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.

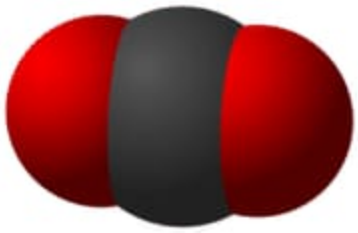


Figure 2.7.3: Carbon dioxide molecules consist of a central carbon atom bonded to 2 oxygen atoms.

Ionic vs. Molecular Compounds

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays. The table below summarizes some of the differences between ionic and molecular compounds.

Table 2.7.1 Comparison of Ionic and Molecular Compounds

Property	Ionic Compounds	Molecular Compounds
Type of elements	Metal and nonmetal	Nonmetals only
Bonding	Ionic - transfer of electron(s) between atoms	Covalent - sharing of pair(s) of electrons between atoms
Representative unit	Formula unit	Molecule
Physical state at room temperature	Solid	Gas, liquid, or solid
Water solubility	Usually high	Variable
Melting and boiling temperatures	Generally high	Generally low
Electrical conductivity	Good when molten or in solution	Poor

Summary

- The law of definite proportions states that a given chemical compound always contains the same elements in the exact same proportions by mass.
- The law of multiple proportions states that whenever the same two elements form more than one compound, the different masses of one element that combine with the same mass of the other element are in the ratio of small whole numbers.
- An ionic bond is formed when a metal donates it(s) valence electron(s) to a nonmetal.
- The resulting ionic compound is more stable and less reactive
- A molecular compound is usually composed of two or more nonmetal elements.
- Ionic compounds usually form hard crystalline solids with high melting points. Covalent molecular compounds, in contrast, consist of discrete molecules held together by weak intermolecular forces and can be gases, liquids, or solids at room temperature and pressure.
- Ionic compounds in molten form or in solution can conduct electricity while molecular compounds do not.

References

1. Petrucci, Ralph, William Harwood, Geoffrey Herring, and Jeffry Madura. General Chemistry. 9th ed. Upper Saddle River, New Jersey: Pearson Prentince Hall, 2007.
2. Moore, John. Chemistry for Dummies. John Wiley & Sons Inc, 2002.
3. Asimov, Isaac. A Short History of Chemistry. , CT.: Greenwood Press, 1965.
4. Patterson, Elizabeth C. John Dalton and the Atomic Theory. Garden City, NY: Doubleday, 1970.
5. Myers, Richard. The Basics of Chemistry. Greenwood, 2003.
6. Demtröder, Wolfgang. Atoms, Molecules and Photons: An Introduction to Atomic- Molecular- and Quantum Physics. 1st ed.

Springer. 2002.
7. Vollhardt, K. Peter C., and Neil E. Schore. Organic Chemistry Structure and Function. New York: W. H. Freeman, 2007.
8. Petrucci, Ralph H. General Chemistry: Principles and Modern Applications. Upper Saddle River, NJ: Pearson Education, 2007.
9. Brown, Theodore L., Eugene H. Lemay, and Bruce E. Bursten. Chemistry: The Central Science. 6th ed. Englewood Cliffs, NJ: Prentice Hall, 1994.

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Answers

Exercise 2.7.1

In compound X, the mass ratio of carbon to hydrogen is $\frac{14.13\text{ g C}}{2.96\text{ g H}}$.

In compound Y, the mass ratio of carbon to hydrogen is $\frac{19.91\text{ g C}}{3.34\text{ g H}}$.

The ratio of these ratios is

$$\frac{\frac{14.13\text{ g C}}{2.96\text{ g H}}}{\frac{19.91\text{ g C}}{3.34\text{ g H}}} = \frac{4.77\text{ g C/g H}}{5.96\text{ g C/g H}} = 0.800 = \frac{4}{5}.$$

This small, whole-number ratio supports the law of multiple proportions. This means that X and Y are different compounds.

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2.8: Solutions

Learning Objective

- Learn some terminology involving solutions.
- Learn how temperature and pressure can affect solubility for different types of solutes.
- Be able to describe the differences between electrolytes, nonelectrolytes, strong electrolytes, and weak electrolytes.
- Be able to use solubility rules to determine the solubility of common ionic compounds.
- Identify whether given substances are nonelectrolytes, strong electrolytes, or weak electrolytes based on the type of compound.

A solution is a *homogeneous mixture*: a mixture of two or more substances that are so intimately mixed that the mixture behaves in many ways like a single substance. Many chemical reactions occur when the reactants are dissolved in solution. In this section, we will introduce concepts that are applicable to solutions and the chemical reactions that occur in them.

More than 70% of the earth's surface is covered by a very important solution—seawater. It is likely that without seawater, no life would exist on Earth. At its simplest, seawater is mostly H_2O. However, about 3.5% of seawater is dissolved solids—mostly NaCl—but other ions as well. Table 2.8.1 lists the percentage by mass of the various ions in seawater. Because it is highly likely that life on Earth originated in the oceans, it should not be surprising that many bodily fluids resemble seawater, especially blood. Table 2.8.1 also lists the percentage by mass of ions in a typical sample of blood.

Table 2.8.1: Percentage by Mass of Ions in Seawater and Blood

Ion	Percentage in Seawater	Percentage in Blood
Na^+	2.36	0.322
Cl^-	1.94	0.366
Mg^{2+}	0.13	0.002
SO_4^{2-}	0.09	-
K^+	0.04	0.016
Ca^{2+}	0.04	0.0096
HCO_3^-	0.002	0.165
HPO_4^{2-}, $H_2PO_4^-$	-	0.01

Most ions are more abundant in seawater than they are in blood, with some notable exceptions. There is far more hydrogen carbonate ion (HCO_3^-) in blood than in seawater; indeed, it is the third most common ion in blood. This difference is significant because the HCO_3^- ion and some related species [CO_3^{2-}, $CO_2(aq)$] have an important role in controlling the acid-base properties of blood. Although there is a negligible amount of the two hydrogen phosphate ions (HPO_4^{2-} and $H_2PO_4^-$) in seawater, there is a small amount in blood, where these ions affect acid-base properties. Another notable difference is that blood has a negligible amount of the sulfate ion (SO_4^{2-}), but this ion is present in seawater. Gold is present in seawater—but only a tiny amount. A current estimate of the amount of gold is about 1 part per every 1×10^{13} parts of seawater, which makes the extraction of gold from seawater unfeasible. However, it does mean that there are about 1.4×10^{14} g of gold in the world's oceans!

Important Terminology Related to Solutions

The major component of a solution is called the **solvent.** The minor component of a solution is called the **solute**. "Major" and "minor" indicate which component has the greater or lesser presence by mass or moles, respectively. Sometimes this becomes confusing, especially when considering substances with very different molar masses. We will confine the discussion here to solutions for which the major component and the minor component are obvious.

Solutions exist for every possible phase of the solute and the solvent. Salt water, for example, is a solution of solid NaCl in liquid water; soda water is a solution of gaseous CO_2 in liquid water, while air is a solution of a gaseous solute (O_2) in a gaseous solvent (N_2). In all cases, however, the overall phase of the solution is the same phase as the solvent.

Example 2.8.1: Sugar Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

The obvious minor component is sucrose, so it is the solute. Water—the majority component—is the solvent. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 2.8.1

A solution is made by dissolving 3.33 g of HCl(g) in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.

Answer

solute: HCl(g); solvent: CH_3OH

One important concept of solutions is in defining how much solute is dissolved in a given amount of solvent. This concept is called **concentration**, which we will learn more about later. Various words are used to describe the relative amounts of solute. **Dilute** describes a solution that has very little solute, while **concentrated** describes a solution that has a lot of solute. One problem is that these terms are qualitative; they describe more or less but not exactly how much.

A solid solute that dissolves when it is added to a solvent is referred to as **soluble**, while a solid solute that does not dissolve when added to a solvent is referred to as **insoluble**. However, this terminology is somewhat of an oversimplification, because often solutes we describe as insoluble are actually partially soluble, just much less than ones called soluble.

In most cases, only a certain maximum amount of solute can be dissolved in a given amount of solvent. This maximum amount is called the **solubility** of the solute. It is usually expressed in terms of the amount of solute that can dissolve in 100 g of the solvent at a given temperature. Table 2.8.1 lists the solubilities of some simple ionic compounds. These solubilities vary widely: NaCl can dissolve up to 31.6 g per 100 g of H_2O, while AgCl can dissolve only 0.00019 g per 100 g of H_2O.

Table 2.8.1: Solubilities of Some Ionic Compounds

Solute	Solubility (g per 100 g of H_2O at 25°C)
AgCl	0.00019
$CaCO_3$	0.0006
KBr	70.7
NaCl	36.1
$NaNO_3$	94.6

When the maximum amount of solute has been dissolved in a given amount of solvent, we say that the solution is **saturated** with solute. When less than the maximum amount of solute is dissolved in a given amount of solute, the solution is **unsaturated**. These terms are also qualitative terms because each solute has its own solubility. A solution of 0.00019 g of AgCl per 100 g of H_2O may be saturated, but with so little solute dissolved, it is also rather dilute. A solution of 36.1 g of NaCl in 100 g of H_2O is also saturated but rather concentrated. When we need to know the actual amount of solute in a solution we will quantify the concentration of the solution.

In some circumstances, it is possible to dissolve more than the maximum amount of a solute in a solution. Usually, this happens by heating the solvent, dissolving more solute than would normally dissolve at regular temperatures, and letting the solution cool down slowly and carefully. Such solutions are called **supersaturated** solutions and are not stable; given an opportunity (such as dropping a crystal of solute in the solution), the excess solute will precipitate from the solution.

Influence of Temperature on Solubility

The **solubility** of a substance is the amount of that substance that is required to form a saturated solution in a given amount of solvent at a specified temperature. Solubility is often measured as the grams of solute per $100\ \text{g}$ of solvent. The solubility of sodium chloride in water is $36.0\ \text{g}$ per $100\ \text{g}$ water at 20°C. The temperature must be specified because solubility varies with temperature. For gases, the pressure must also be specified. Solubility is specific for a particular solvent. In this section, we will consider solubility of material in water as solvent.

The solubility of the majority of solid substances increases as the temperature increases. However, the effect is difficult to predict and varies widely from one solute to another. The temperature dependence of solubility can be visualized with the help of a **solubility curve**, a graph of the solubility vs. temperature (see figure below).

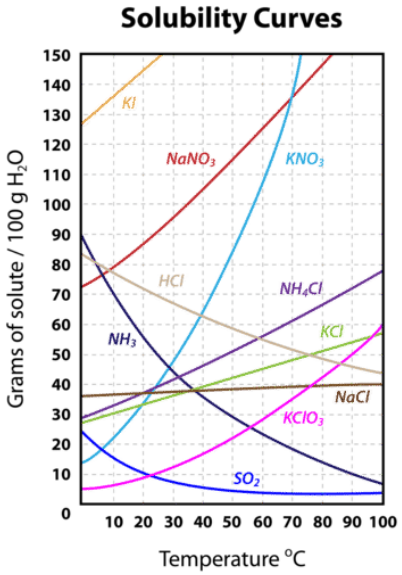


Figure 2.8.1: Solubility curves for several compounds.

Notice how the temperature dependence of NaCl is fairly flat, meaning that an increase in temperature has relatively little effect on the solubility of NaCl. The curve for KNO_3, on the other hand, is very steep, and so an increase in temperature dramatically increases the solubility of KNO_3.

Several substances—HCl, NH_3, and SO_2—have solubility that decreases as temperature increases. They are all gases at standard pressure. When a solvent with a gas dissolved in it is heated, the kinetic energy of both the solvent and solute increase. As the kinetic energy of the gaseous solute increases, its molecules have a greater tendency to escape the attraction of the solvent molecules and return to the gas phase. Therefore, the solubility of a gas decreases as the temperature increases. A real life example of this involves nuclear power plants and their impact on organisms living in nearby water. Nuclear power plants require large amounts of water to generate steam for turbines and to cool equipment. They are usually situated near bodies of water to use that water as a coolant, returning the warmer water back to the lake or river. This increases the overall temperature of the water, which lowers the quantity of dissolved oxygen, affecting the survival of fish and other organisms.

Solubility curves can be used to determine if a given solution is saturated or unsaturated. Suppose that $80\ \text{g}$ of KNO_3 is added to $100\ \text{g}$ of water at 30°C. According to the solubility curve, approximately $48\ \text{g}$ of KNO_3 will dissolve at 30°C. This means that the solution will be saturated since $48\ \text{g}$ is less than $80\ \text{g}$. We can also determine that there will be $80 - 48 = 32\ \text{g}$ of undissolved KNO_3 remaining at the bottom of the container. In a second scenario, suppose that this saturated solution is heated to 60°C. According to the curve, the solubility of KNO_3 at 60°C is about $107\ \text{g}$. The solution, in this case, is unsaturated since it contains only the original $80\ \text{g}$ of dissolved solute. Suppose in a third case, that the solution is cooled all the way down to 0°C. The solubility at 0°C is about $14\ \text{g}$, meaning that $80 - 14 = 66\ \text{g}$ of the KNO_3 will recrystallize.

Influence of Pressure on Solubility

Pressure has very little effect on the solubility of solids or liquids, but has a significant effect on the solubility of gases. Gas solubility increases as the partial pressure of a gas above the liquid increases. Suppose a certain volume of water is in a closed container with the space above it occupied by carbon dioxide gas at standard pressure. Some of the CO_2 molecules come into contact with the surface of the water and dissolve into the liquid. Now suppose that more CO_2 is added to the space above the container, causing a pressure increase. In this case, more CO_2 molecules are in contact with the water and so more of them dissolve. Thus, the solubility increases as the pressure increases. As with a solid, the CO_2 that is undissolved reaches an equilibrium with the dissolved CO_2, represented by the equation:

$$\text{CO}_2\,(g) \rightleftharpoons \text{CO}_2\,(aq)$$

At equilibrium, the rate of gaseous CO_2 dissolution is equal to the rate of dissolved CO_2 coming out of the solution.

When carbonated beverages are packaged, they are done so under high CO_2 pressure so that a large amount of carbon dioxide dissolves in the liquid. When the bottle is open, the equilibrium is disrupted because the CO_2 pressure above the liquid decreases. Immediately, bubbles of CO_2 rapidly exit the solution and escape out of the top of the open bottle. The amount of dissolved CO_2 decreases. If the bottle is left open for an extended period of time, the beverage becomes "flat" as more and more CO_2 comes out of the liquid.

The relationship of gas solubility to pressure is described by Henry's law, named after English chemist William Henry (1774-1836). **Henry's Law** states that the solubility of a gas in a liquid is directly proportional to the partial pressure of the gas above the liquid. Henry's law can be written as follows:

$$\frac{S_1}{P_1} = \frac{S_2}{P_2}$$

S_1 and P_1 are the solubility and the pressure at an initial set of conditions; S_2 and P_2 are the solubility and pressure at another changed set of conditions. The solubility of a gas is typically reported in g/L.

Example 2.8.1

The solubility of a certain gas in water is $0.745\ \text{g/L}$ at standard pressure (1atm). What is its solubility when the pressure above the solution is raised to $4.50\ \text{atm}$? The temperature is constant at 20°C.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $S_1 = 0.745\ \text{g/L}$
- $P_1 = 1.00\ \text{atm}$
- $P_2 = 4.50\ \text{atm}$

Unknown

Substitute into Henry's law and solve for S_2.

Step 2: Solve.

$$S_2 = \frac{S_1 \times P_2}{P_1} = \frac{0.745\ \text{g/L} \times 4.50\ \text{atm}}{1.00\ \text{atm}} = 3.35\ \text{g/L}$$

Step 3: Think about your result.

The solubility is increased to 4.5 times its original value, according to the direct relationship.

Electrolytes and Nonelectrolytes

An **electrolyte** is a compound that conducts an electric current when it is in an aqueous solution. In order to conduct a current, a substance must contain mobile ions that can move from one electrode to the other. All ionic compounds are electrolytes. When ionic compounds dissolve, they break apart into ions which are then able to conduct a current. Even insoluble ionic compounds such as CaCO_3 are electrolytes because although the majority of the compound does not dissolve, there are still a small number of ions that do break free in solution which can conduct a current. All acids and bases are also electrolytes, producing hydrogen ions (H^+) or hydroxide ions (OH^-) in solution along with counterions (ions that balance the charge to zero).

A **nonelectrolyte** is a compound that does not conduct an electric current in either aqueous solution. Many molecular compounds, such as sugar or ethanol, are nonelectrolytes. When these compounds dissolve in water, they do not produce ions. The figure below illustrates the difference between an electrolyte and a nonelectrolyte.

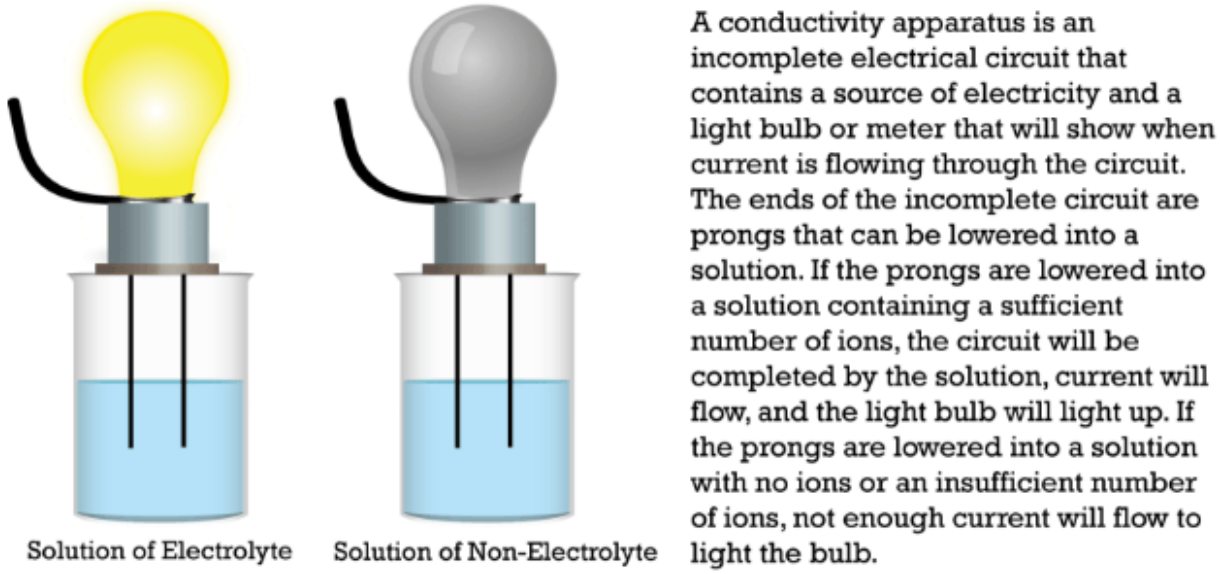


Figure 2.8.1: Conductivity apparatus.

Strong and Weak Electrolytes

An electrolyte is a compound that conducts electricity when dissolved in aqueous solution, but some electrolytes conduct electricity very well while others do not conduct electricity very well. What allows an electrolyte to conduct electricity in solution is the fact that they split into or form ions when dissolved in water. The more ions in solution, the better the conduction of electricity. Often electrolytes are split into two categories: **strong electrolytes** which conduct electricity very well due to an abundance of ions (a large fraction of dissolved solute exists as ions) and **weak electrolytes** which only weakly conduct electricity due to the presence of a small amount of ions (only a small fraction of the dissolved solute exists as ions).

All ionic compounds are electrolytes, but some are strong and some are weak. You can distinguish between whether an ionic compound is strong or weak based on its solubility. Soluble ionic compounds are strong electrolytes, whereas insoluble ionic compounds are weak electrolytes. The following tables provide a general set of guidelines for identifying whether an ionic compound is soluble or insoluble.

Ionic compounds containing the following ions are generally soluble

Li^+, Na^+, K^+, Rb^+, Cs^+

NH_4^+

NO_3^-

$C_2H_3O_2^-$

Cl^-, Br^-, I^-

- Exceptions: Insoluble when these ions are paired with Ag^+, Hg_2^{2+}, or Pb^{2+}.

SO_4^{2-}

- Exceptions: Insoluble when paired with Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}

Ionic compounds containing the following ions are generally insoluble

OH^-

- Exceptions: Soluble when paired with Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+
- Exceptions: Slightly soluble when paired with Ca^{2+}, Sr^{2+}, Ba^{2+}.

S^{2-}

- Exceptions: Soluble when paired with Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+, Ca^{2+}, Sr^{2+}, Ba^{2+}.

CO_3^{2-}, PO_4^{3-}

- Exceptions: Soluble when paired with Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+

Some molecular compounds are also electrolytes, specifically molecular compounds that act as acids or bases in water. A molecular compound that produces hydrogen ions (H^+) when dissolved in water is an acid, whereas one that produces hydroxide ion (OH^-) when dissolved in water is a base. For now, the bases we will deal with will all be metal hydroxides, meaning they an ionic compound made up of a metal ion combined with hydroxide ion. These bases are ionic compounds, so solubility rules can be used to determine whether they are soluble or insoluble and whether they are a strong or weak electrolyte. However, the acids we will deal with for now are all molecular compounds, thus the solubility rules will not help us to identify whether they are a strong or weak electrolyte. Instead, an acid's classification as a strong acid or weak acid tells us whether the acid is a strong or weak electrolyte. There are only seven acids we consider to be strong acids (see table below), and all others are considered to be weak acids. Strong acids fully ionize in water forming a large number of hydrogen ions and counterions, thus are strong electrolytes. Weak acids only partially ionize in water forming a small amount of hydrogen ions and counterions, thus are weak electrolytes.

Strong Acids	Some Common Weak Acids
HCl	HF
HBr	$HC_2H_3O_2$
HI	HNO_2
HNO_3	H_2CO_3
$HClO_3$	H_3PO_4
$HClO_4$	
H_2SO_4	

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Concentration is the expression of the amount of solute in a given amount of solvent and can be described by several qualitative terms.
- Solubility is the maximum amount of solute that can dissolve in a given amount of solvent at a specified temperature.
- The solubility of a solid in water increases with an increase in temperature.
- Gas solubility decreases as the temperature increases.
- Electrolytes conduct electric current when dissolved in water.
- Nonelectrolytes do not conduct electric current when dissolved in water.
- A strong electrolyte exists mainly as ions in solution.
- A weak electrolyte has only a small amount of ionization in solution.

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3: THE LANGUAGE OF CHEMISTRY (UNDER CONSTRUCTION)

3.1: Element and Ion Nomenclature
3.2: Writing Formulas of Ionic Compounds
3.3: Naming Ionic Compounds
3.4: Acid Nomenclature
3.5: Binary Covalent Compound Nomenclature

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3.1: Element and Ion Nomenclature

Learning Objectives

- Use the rules for naming ions.
- Write the correct formula for an ion from its name.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 4.4.1) . Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system** (Table 3.1.1), an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word ion. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table 3.1.1: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name (Stock System)	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 3.1.1

lists the elements that use the common system, along with their respective cation names.

Table 3.1.2: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 3.1.2 lists the names of some common monatomic ions.

Example 3.1.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the copper(I) ion or the cuprous ion

Exercise 3.1.1

Name each ion.

a. Fe^{2+}
b. Fe^{3+}
c. Ba^{2+}

Answer a: available at the end of this section.

Example 3.1.2

Write the formula for each ion.

a. the bromide ion
b. the cupric ion
c. the magnesium ion

Solution

a. Br^-
b. Cu^{2+}
c. Mg^{2+}

Exercise 3.1.2

Write the formula for each ion.

a. the fluoride ion
b. the stannous ion
c. the potassium ion

Answer a: available at the end of this section.

Summary

- Some cations have multiple forms and are named accordingly with the use of roman numerals in parenthesis.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
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Learning Objectives

- Know how ions form.
- Learn the characteristic charges that ions have.
- Construct a proper formula for an ionic compound.
- Generate a proper name for an ionic compound.

There also exists a group of ions that contain more than one atom. These are called **polyatomic ions**. Table 3.1.2 lists the formulas, charges, and names of some common polyatomic ions. Only one of them, the ammonium ion, is a cation; the rest are anions. Most of them also contain oxygen atoms, so sometimes they are referred to as ***oxyanions***. Some of them, such as nitrate and nitrite, and sulfate and sulfite, have very similar formulas and names, so care must be taken to get the formulas and names correct. Note that the -ite polyatomic ion has one less oxygen atom in its formula than the -ate ion but with the same ionic charge.

Table 3.1.2: Common Polyatomic Ions

Name	Formula and Charge	Name	Formula and Charge
ammonium	NH_4^+	hydroxide	OH^-
acetate	$C_2H_3O_2^-$, or CH_3COO^-	nitrate	NO_3^-

Name	Formula and Charge		Name	Formula and Charge
bicarbonate (hydrogen carbonate)	HCO_3^-		nitrite	NO_2^-
bisulfate (hydrogen sulfate)	HSO_4^-		peroxide	O_2^{2-}
carbonate	CO_3^{2-}		perchlorate	ClO_4^-
chlorate	ClO_3^-		phosphate	PO_4^{3-}
chromate	CrO_4^{2-}		sulfate	SO_4^{2-}
cyanide	CN^-		sulfite	SO_3^{2-}
dichromate	$Cr_2O_7^{2-}$		triiodide	I_3^-

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions and negative ions.
- Ionic formulas balance the total positive and negative charges.
- Ionic compounds have a simple system of naming.
- Groups of atoms can have an overall charge and make ionic compounds.

Answers

Exercise 3.1.1

iron(II) ion

iron(III) ion

barium ion

Exercise 3.1.2

F^-

Sn^{2+}

K^+

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3.2: Writing Formulas of Ionic Compounds

Learning Objectives

- Know how ions form.
- Learn the characteristic charges that ions have.
- Construct a proper formula for an ionic compound.
- Generate a proper name for an ionic compound.

Ionic Formulas

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula has a cation and an anion in it; an ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is simple: the total positive charge must balance the total negative charge. Because the charges on the ions are characteristic, sometimes we have to have more than one of a cation or an anion to balance the overall positive and negative charges. It is conventional to use the lowest ratio of ions that are needed to balance the charges.

For example, consider the ionic compound between Na^+ and Cl^-. Each ion has a single charge, one positive and one negative, so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions: (1) write the formula for the cation first and the formula for the anion second, but (2) do not write the charges on the ions. Thus, for the compound between Na^+ and Cl^-, we have the ionic formula NaCl (Figure 3.2.1). The formula Na_2Cl_2 also has balanced charges, but the convention is to use the lowest ratio of ions, which would be one of each. (Remember from our conventions for writing formulas that we do not write a 1 subscript if there is only one atom of a particular element present.) For the ionic compound between magnesium cations (Mg^{2+}) and oxide anions (O^{2-}), again we need only one of each ion to balance the charges. By convention, the formula is MgO.



Figure 3.2.1: The ionic compound NaCl is table salt and is very common. (Public Domain; Pexels.)

For the ionic compound between Mg^{2+} ions and Cl^- ions, we now consider the fact that the charges have different magnitudes: 2+ on the magnesium ion and 1– on the chloride ion. To balance the charges with the lowest number of ions possible, we need to have two chloride ions to balance the charge on the one magnesium ion. Rather than write the formula MgClCl, we combine the two chloride ions and write it with a 2 subscript: $MgCl_2$.

What is the formula $MgCl_2$ telling us? There are two chloride ions in the formula. Although chlorine as an element is a diatomic molecule, Cl_2, elemental chlorine is not part of this ionic compound. The chlorine is in the form of a negatively charged *ion*, not the neutral *element*. The 2 subscript is in the ionic formula because we need two Cl^- ions to balance the charge on one Mg^{2+} ion.

Example 3.2.2

Write the proper ionic formula for each of the two given ions.

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

Solution

a. We need two Cl^- ions to balance the charge on one Ca^{2+} ion, so the proper ionic formula is $CaCl_2$.
b. We need three F^- ions to balance the charge on the Al^{3+} ion, so the proper ionic formula is AlF_3.
c. With Al^{3+} and O^{2-}, note that neither charge is a perfect multiple of the other. This means we have to go to a least common multiple, which in this case will be six. To get a total of 6+, we need two Al^{3+} ions; to get 6–, we need three O^{2-} ions. Hence the proper ionic formula is Al_2O_3.

Exercise 3.2.2

Write the proper ionic formulas for each of the two given ions.

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

Answers

a. FeS
b. Fe_2S_3

Writing the formulas of ionic compounds has one important difference. If more than one polyatomic ion is needed to balance the overall charge in the formula, enclose the formula of the polyatomic ion in parentheses and write the proper numerical subscript to the right and *outside* of the parentheses. Thus, the formula between calcium ions, Ca^{2+}, and nitrate ions, NO_3^-, is properly written $Ca(NO_3)_2$, not $CaNO_{32}$ or CaN_2O_6. Use parentheses where required. The name of this ionic compound is simply calcium nitrate.

Example 3.2.4

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

a. NH_4^+ and S^{2-}
b. Al^{3+} and PO_4^{3-}
c. Fe^{2+} and PO_4^{3-}

Solution

a. Because the ammonium ion has a 1+ charge and the sulfide ion has a 2– charge, we need two ammonium ions to balance the charge on a single sulfide ion. Enclosing the formula for the ammonium ion in parentheses, we have $(NH_4)_2S$. The compound's name is ammonium sulfide.
b. Because the ions have the same magnitude of charge, we need only one of each to balance the charges. The formula is $AlPO_4$, and the name of the compound is aluminum phosphate.
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, we need three iron(II) ions, and to get 6–, we need two phosphate ions. The proper formula is $Fe_3(PO_4)_2$, and the compound's name is iron(II) phosphate.

Exercise 3.2.4

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

a. NH_4^+ and PO_4^{3-}
b. Co^{3+} and NO_2^-

Answers

a. $(NH_4)_3PO_4$, ammonium phosphate
b. $Co(NO_2)_3$, cobalt(III) nitrite

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions and negative ions.
- Ionic formulas balance the total positive and negative charges.
- Ionic compounds have a simple system of naming.
- Groups of atoms can have an overall charge and make ionic compounds.

Learning Objectives

- Use the rules for naming ionic compounds.
- Write the correct formula for an ionic compound.

Writing Formulas of Ionic Compounds

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

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charge to cancel each other out.

Example 3.2.5: Aluminum Nitride and Lithium Oxide

Write the formula for aluminum nitride and lithium oxide.

Solution

Example 3.2.5: Steps for Problem Solving Writing Formulas

Steps for Problem Solving	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] [redacted]	Li^+ [redacted]
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

An alternative way to writing a correct formula for an ionic compound is to use the **crisscross method**. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped.

Example 3.2.6: The Crisscross Method for lead (IV) oxide

Write the formula for lead (IV) oxide.

Solution

Example 3.2.6: Steps For Problem Solving the Crisscross Method for writing formulas.

Crisscross Method	Example: Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] [redacted]
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all subscripts that are 1.	PbO_2

Exercise 3.2.6

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion

b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a: available at the end of this section.

Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.

Example 3.2.7: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Example 3.2.7: Steps For Problem Solving the Crisscross Method for writing formulas.

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Na^+ [redacted]
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all all subscripts that are 1.	Na_2S

Exercise 3.2.7

Write the formula for each ionic compound.

a. sodium bromide
b. lithium chloride
c. magnesium oxide

Answer a: available at the end of this section.

Answer available at the end of this section.

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parenthesis.
- Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

Contributors and Attributions

-

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- Henry Agnew (UC Davis)

Learning Objectives

- Recognize polyatomic ions in chemical formulas.
- Write the correct formula for ionic compounds with polyatomic ions.
- Use the rules for naming ionic compounds with polyatomic ions.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

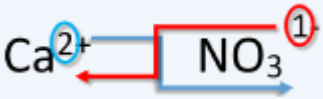
Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 3.2.1: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Example 3.2.1: Steps For Problem Solving, Write the formula for calcium nitrate

Criss Cross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] NO_3^-
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 3.2.2

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion

Solution

Example 3.2.2: Explanation for Writing the Chemical Formula for an Ionic Compound Composed of the Potassium Ion and the Sulfate Ion

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2–. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 3.2.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a: available at the end of this section.

Summary

- Formulas for ionic compounds contain the symbols and number of each atom and/or polyatomic ion present in a compound in the lowest whole number ratio.
- The process of naming ionic compounds with polyatomic ions are the same as binary ionic compounds.

Contributors and Attributions

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Answers

Exercise 3.2.6

CaO

CuS

Cu_2S

Exercise 3.2.7

NaBr

LiCl

MgO

Exercise 3.2.1

$MgCO_3$

$Al(CH_3COO)_3$

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3.3: Naming Ionic Compounds

Learning Objectives

- Use the rules for naming ionic compounds.
- Write the correct formula for an ionic compound.

Names of Binary ionic Compounds

After learning a few more details about the names of individual ions, you will be a step away from knowing how to name ionic compounds. This section begins the formal study of nomenclature, the systematic naming of chemical compounds.

Naming Ions

The name of a monatomic cation is simply the name of the element followed by the word *ion*. Thus, Na^+ is the sodium ion, Al^{3+} is the aluminum ion, Ca^{2+} is the calcium ion, and so forth.

We have seen that some elements lose different numbers of electrons, producing ions of different charges (Figure 4.4.1) . Iron, for example, can form two cations, each of which, when combined with the same anion, makes a different compound with unique physical and chemical properties. Thus, we need a different name for each iron ion to distinguish Fe^{2+} from Fe^{3+}. The same issue arises for other ions with more than one possible charge.

There are two ways to make this distinction. In the simpler, more modern approach, called the **Stock system** (Table 3.3.1), an ion's positive charge is indicated by a roman numeral in parentheses after the element name, followed by the word ion. Thus, Fe^{2+} is called the iron(II) ion, while Fe^{3+} is called the iron(III) ion. This system is used only for elements that form more than one common positive ion. We do not call the Na^+ ion the sodium(I) ion because (I) is unnecessary. Sodium forms only a 1+ ion, so there is no ambiguity about the name *sodium ion*.

Table 3.3.1: The Modern and Common System of Cation Names

Element	Stem	Charge	Modern Name (Stock System)	Common Name
iron	ferr-	2+	iron(II) ion	ferrous ion
		3+	iron(III) ion	ferric ion
copper	cupr-	1+	copper(I) ion	cuprous ion
		2+	copper(II) ion	cupric ion
tin	stann-	2+	tin(II) ion	stannous ion
		4+	tin(IV) ion	stannic ion
lead	plumb-	2+	lead(II) ion	plumbous ion
		4+	lead(IV) ion	plumbic ion
chromium	chrom-	2+	chromium(II) ion	chromous ion
		3+	chromium(III) ion	chromic ion
gold	aur-	1+	gold(I) ion	aurous ion

Element	Stem	Charge	Modern Name (Stock System)	Common Name
		3+	gold(III) ion	auric ion

The second system, called the **common system**, is not conventional but is still prevalent and used in the health sciences. This system recognizes that many metals have two common cations. The common system uses two suffixes (*-ic* and *-ous*) that are appended to the stem of the element name. The *-ic* suffix represents the greater of the two cation charges, and the *-ous* suffix represents the lower one. In many cases, the stem of the element name comes from the Latin name of the element. Table 3.3.1 lists the elements that use the common system, along with their respective cation names.

Table 3.3.2: Some Monatomic Anions

Ion	Name
F^-	fluoride ion
Cl^-	chloride ion
Br^-	bromide ion
I^-	iodide ion
O^{2-}	oxide ion
S^{2-}	sulfide ion
P^{3-}	phosphide ion
N^{3-}	nitride ion

The name of a monatomic anion consists of the stem of the element name, the suffix *-ide*, and then the word *ion*. Thus, as we have already seen, Cl^- is "chlor-" + "-ide ion," or the chloride ion. Similarly, O^{2-} is the oxide ion, Se^{2-} is the selenide ion, and so forth. Table 3.3.2 lists the names of some common monatomic ions.

Example 3.3.1

Name each ion.

a. Ca^{2+}
b. S^{2-}
c. Cu^+

Solution

a. the calcium ion
b. the sulfide ion
c. the copper(I) ion or the cuprous ion

Exercise 3.3.1

Name each ion.

a. Fe^{2+}

b. Fe^{3+}
c. Ba^{2+}

Answer a: available at the end of this section.

Example 3.3.2

Write the formula for each ion.

a. the bromide ion
b. the cupric ion
c. the magnesium ion

Solution

a. Br^{-}
b. Cu^{2+}
c. Mg^{2+}

Exercise 3.3.2

Write the formula for each ion.

a. the fluoride ion
b. the stannous ion
c. the potassium ion

Answer a: available at the end of this section.

Naming Binary Ionic Compounds with a Metal that Forms Only One Type of Cation

Now that we know how to name ions, we are ready to name ionic compounds. A **binary** ionic compound is a compound composed of a **monatomic** metal **cation** and a monatomic nonmetal **anion**. The metal cation is named first, followed by the nonmetal anion as illustrated in Figure 3.3.1 for the compound $BaCl_2$. The word *ion* is dropped from both parts.

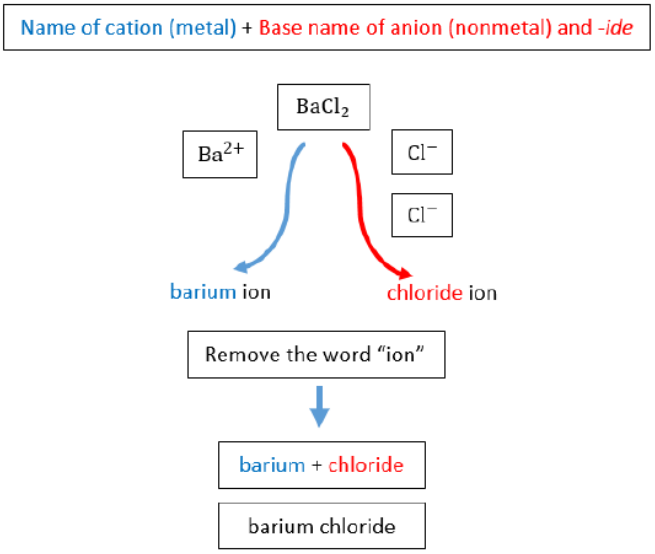


Figure 3.3.1 *Naming* $BaCl_2$

Naming formula: Name of metal cation + base name of nonmetal anion + suffix -ide. BaCl2 is named as barium chloride.

Subscripts in the formula do not affect the name.

Example 3.3.3: Naming Ionic Compounds

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. KCl

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride.
b. The name of this ionic compound is aluminum fluoride.
c. The name of this ionic compound is potassium chloride

Exercise 3.3.3

Name each ionic compound.

a. AgI
b. MgO
c. Ca_3P_2

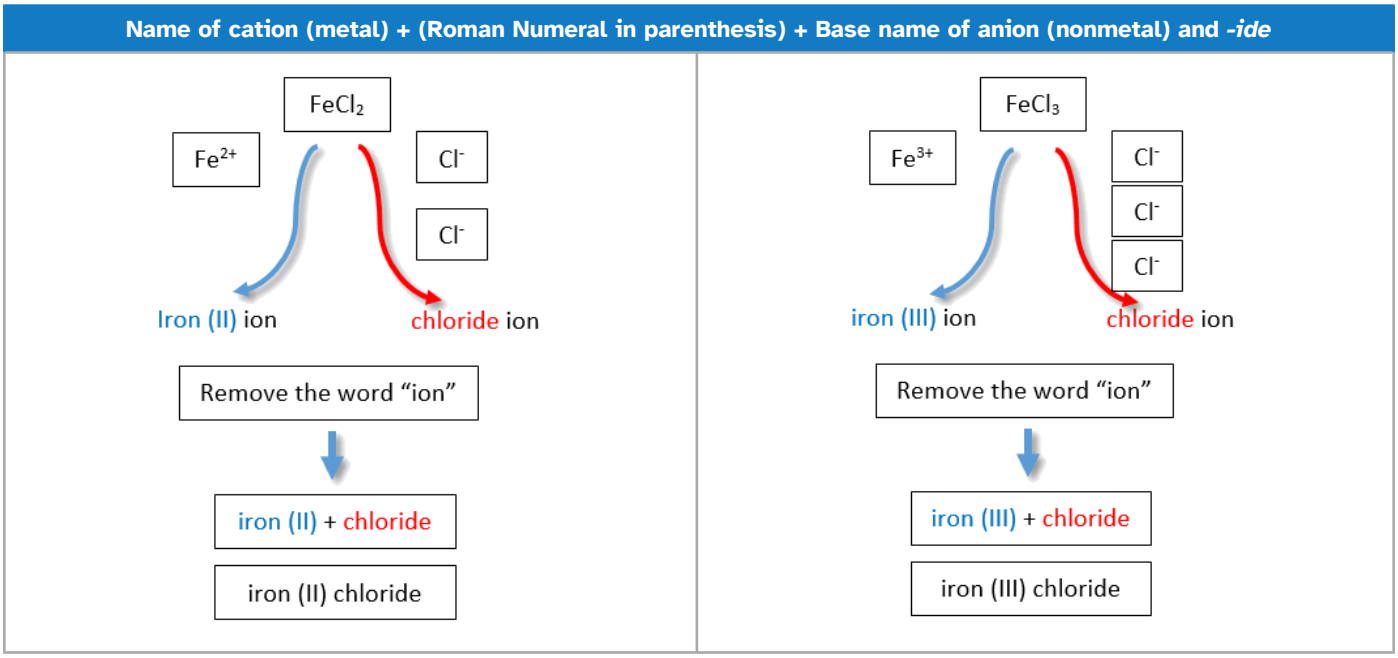
Answer a: available at the end of this section.

Naming Binary Ionic Compounds with a Metal That Forms More Than One Type of Cation

If you are given a formula for an ionic compound whose cation can have more than one possible charge, you must first determine the charge on the cation before identifying its correct name. For example, consider $FeCl_2$ and $FeCl_3$. In the first compound, the iron ion has a 2+ charge because there are two Cl^- ions in the formula (1– charge on each chloride ion). In the second compound,

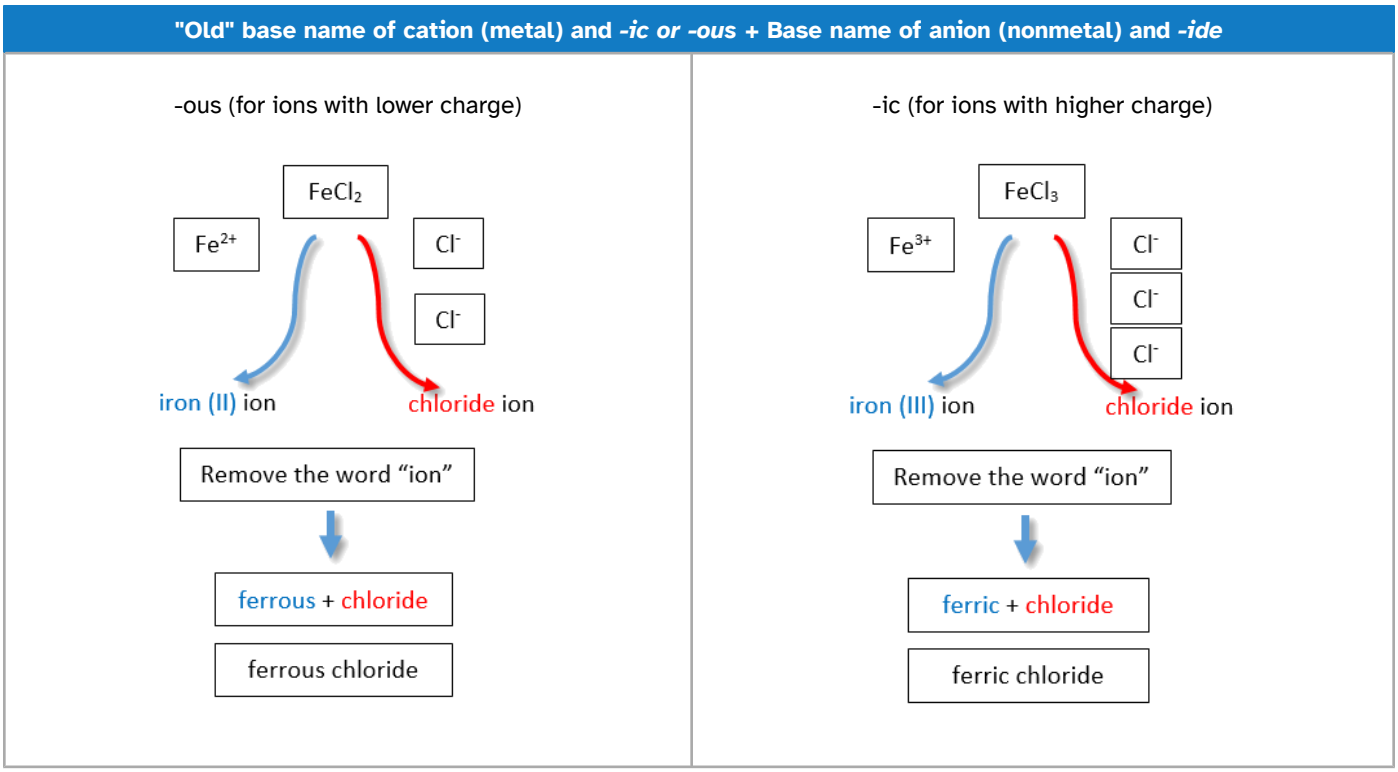
the iron ion has a 3+ charge, as indicated by the three Cl^- ions in the formula. These are two different compounds that need two different names. By the Stock system, the names are iron(II) chloride and iron(III) chloride (Figure 3.3.2).

Figure 3.3.2 Naming the $FeCl_2$ and $FeCl_3$ compounds in the modern/stock system.



If we were to use the stems and suffixes of the common system, the names would be ferrous chloride and ferric chloride, respectively (Figure 3.3.3) .

Figure 3.3.3 Naming the $FeCl_2$ and $FeCl_3$ compounds in the old/common System.



Example 3.3.4:

Name each ionic compound.

a. Co_2O_3
b. $FeCl_2$

Solution

Example 3.3.4: Explanation of the names of ionic compounds

Ionic Compound	Explanation	Answer
Co_2O_3	We know that cobalt can have more than one possible charge; we just need to determine what it is. • Oxide always has a 2– charge, so with three oxide ions, we have a total negative charge of 6–. • This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. • Therefore, the proper name for this ionic compound is cobalt(III) oxide.	cobalt(III) oxide
$FeCl_2$	Iron can also have more than one possible charge. • Chloride always has a 1– charge, so with two chloride ions, we have a total negative charge of 2–. • This means that the one iron ion must have a 2+charge. • Therefore, the proper name for this ionic compound is iron(II) chloride.	iron(II) chloride

Exercise 3.3.4

Name each ionic compound.

a. $AuCl_3$
b. PbO_2
c. CuO

Answer a: available at the end of this section.

Figure 3.3.4 is a synopsis of how to name simple ionic compounds.

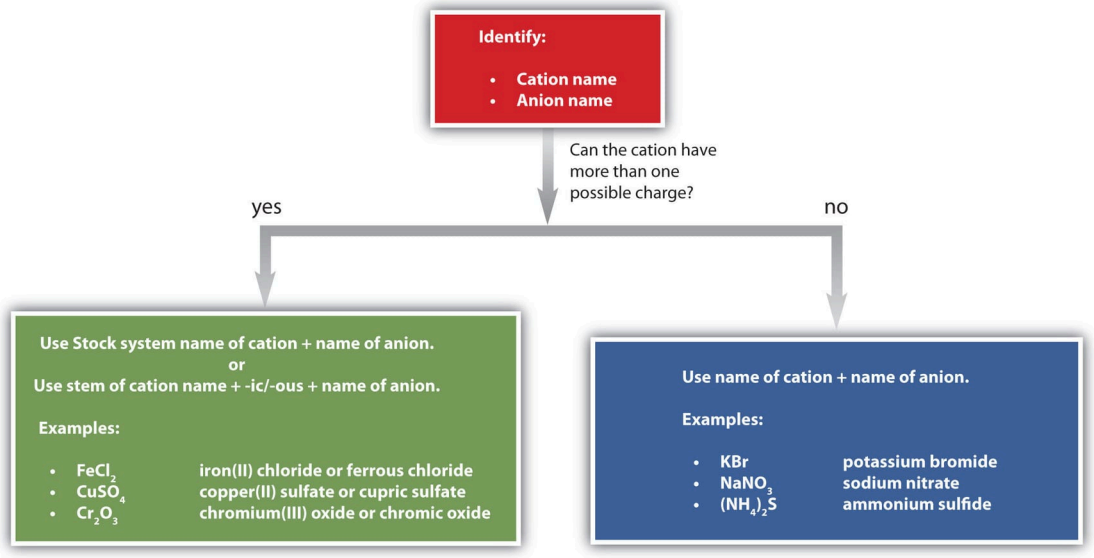


Figure 3.3.4: A Guide to Naming Simple Ionic Compounds.

Exercise 3.3.5

Name each ionic compound.

a. $ZnBr_2$
b. Al_2O_3
c. AuF_3
d. AgF

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

Writing Formulas of Ionic Compounds

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

If you know the name of a binary ionic compound, you can write its **chemical formula**. Start by writing the metal ion with its charge, followed by the nonmetal ion with its charge. Because the overall compound must be electrically neutral, decide how many of each ion is needed in order for the positive and negative charge to cancel each other out.

Example 3.3.5: Aluminum Nitride and Lithium Oxide

Write the formula for aluminum nitride and lithium oxide.

Solution

Example 3.3.5: Steps for Problem Solving Writing Formulas

Steps for Problem Solving	Write the formula for aluminum nitride	Write the formula for lithium oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] [redacted]	Li^+ [redacted]
2. Use a multiplier to make the total charge of the cations and anions equal to each other.	total charge of cations = total charge of anions 1(3+) = 1(3-) +3 = -3	total charge of cations = total charge of anions 2(1+) = 1(2-) +2 = -2
3. Use the multipliers as subscript for each ion.	Al_1N_1	Li_2O_1
4. Write the final formula. Leave out all charges and all subscripts that are 1.	AlN	Li_2O

An alternative way to writing a correct formula for an ionic compound is to use the **crisscross method**. In this method, the numerical value of each of the ion charges is crossed over to become the subscript of the other ion. Signs of the charges are dropped.

Example 3.3.6: The Crisscross Method for lead (IV) oxide

Write the formula for lead (IV) oxide.

Solution

Example 3.3.6: Steps For Problem Solving the Crisscross Method for writing formulas.

Crisscross Method	Example: Write the formula for lead (IV) oxide
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] [redacted]
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	Pb_2O_4
4. Write the final formula. Leave out all subscripts that are 1.	PbO_2

Exercise 3.3.6

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the calcium ion and the oxygen ion
b. the 2+ copper ion and the sulfur ion
c. the 1+ copper ion and the sulfur ion

Answer a: available at the end of this section.

Be aware that ionic compounds are empirical formulas and so must be written as the lowest ratio of the ions.

Example 3.3.7: Sulfur Compound

Write the formula for sodium combined with sulfur.

Solution

Example 3.3.7: Steps For Problem Solving the Crisscross Method for writing formulas.

Crisscross Method	Write the formula for sodium combined with sulfur
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	Na^+ [redacted]
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	
3. Reduce to the lowest ratio.	This step is not necessary.
4. Write the final formula. Leave out all all subscripts that are 1.	Na_2S

Exercise 3.3.7

Write the formula for each ionic compound.

a. sodium bromide
b. lithium chloride
c. magnesium oxide

Answer a: available at the end of this section.

Recognizing Ionic Compounds

There are two ways to recognize ionic compounds. First, compounds between metal and nonmetal elements are usually ionic. For example, $CaBr_2$ contains a metallic element (calcium, a group 2 (or 2A) metal) and a nonmetallic element (bromine, a group 17 (or 7A) nonmetal). Therefore, it is most likely an ionic compound. (In fact, it *is* ionic.) In contrast, the compound NO_2 contains two elements that are both nonmetals (nitrogen, from group 15 (or 5A), and oxygen, from group 16 (or 6A). It is not an ionic compound; it belongs to the category of covalent compounds discussed elsewhere. Also note that this combination of nitrogen and oxygen has no electric charge specified, so it is *not* the nitrite ion.

Second, if you recognize the formula of a polyatomic ion in a compound, the compound is ionic. For example, if you see the formula $Ba(NO_3)_2$, you may recognize the "NO_3" part as the nitrate ion, NO_3^-. (Remember that the convention for writing formulas for ionic compounds is not to include the ionic charge.) This is a clue that the other part of the formula, Ba, is actually the Ba^{2+} ion, with the 2+ charge balancing the overall 2– charge from the two nitrate ions. Thus, this compound is also ionic.

Example 3.3.8

Identify each compound as ionic or not ionic.

a. Na_2O
b. PCl_3
c. OF_2

Solution

Example 3.3.8: Steps For Problem Solving Identifying Ionic Compounds

Explanation	Answer
a. Sodium is a metal, and oxygen is a nonmetal. Therefore, Na_2O is expected to be ionic.	Na_2O, ionic
b. Both phosphorus and chlorine are nonmetals. Therefore, PCl_3 is not ionic.	PCl_3, not ionic
c. Both oxygen and fluorine are nonmetals. Therefore, OF_2 is not ionic.	OF_2, ionic

Exercise 3.3.8

Identify each compound as ionic or not ionic.

a. N_2O
b. $FeCl_3$

Answer a: available at the end of this section.

Summary

- Ionic compounds are named by stating the cation first, followed by the anion.
- Positive and negative charges must balance.
- Some anions have multiple forms and are named accordingly with the use of roman numerals in parenthesis.
- Formulas for ionic compounds contain the symbols and number of each atom present in a compound in the lowest whole number ratio.

Contributors and Attributions

-

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Learning Objectives

- Know how ions form.
- Learn the characteristic charges that ions have.
- Construct a proper formula for an ionic compound.
- Generate a proper name for an ionic compound.

So far, we have discussed elements and compounds that are electrically neutral. They have the same number of electrons as protons, so the negative charges of the electrons are balanced by the positive charges of the protons. However, this is not always the case. Electrons can move from one atom to another; when they do, species with overall electric charges are formed. Such species are called ions. Species with overall positive charges are termed **cations**, while species with overall negative charges are called **anions**. Remember that ions are formed only when *electrons* move from one atom to another; a proton never moves from one atom to another. Compounds formed from positive and negative ions are **ionic compounds**.

Individual atoms can gain or lose electrons. When they do, they become *monatomic* ions. When atoms gain or lose electrons, they usually gain or lose a characteristic number of electrons and so take on a characteristic overall charge. Table 3.3.1 lists some common ions in terms of how many electrons they lose (making cations) or gain (making anions). There are several things to notice about the ions in Table 3.3.1. First, each element that forms a cation is a metal, except for one (hydrogen), while each element that forms an anion is a nonmetal. This is actually one of the chemical properties of metals and nonmetals: metals tend to form cations, while nonmetals tend to form anions. Second, most atoms form ions of a single characteristic charge. When

sodium atoms form ions, they always form a 1+ charge, never a 2+ or 3+ or even 1– charge. Thus, if you commit the information in Table 3.3.1 to memory, you will always know what charges most atoms form.

Table 3.3.1: Monatomic Ions of Various Charges

Ions formed by losing a single electron	H^+
Na^+	
K^+	
Rb^+	
Ag^+	
Au^+	
Ions formed by losing two electrons	Mg^{2+}
	Ca^{2+}
	Sr^{2+}
	Fe^{2+}
	Co^{2+}
	Ni^{2+}
	Cu^{2+}
	Zn^{2+}
	Sn^{2+}
	Hg^{2+}
	Pb^{2+}
Ions formed by losing three electrons	Sc^{3+}
	Fe^{3+}
	Co^{3+}
	Ni^{3+}
	Au^{3+}
	Al^{3+}
	Cr^{3+}
Ions formed by losing four electrons	Ti^{4+}
	Sn^{4+}
	Pb^{4+}

Ions formed by losing a single electron	H^+
Ions formed by gaining a single electron	F^-
	Cl^-
	Br^-
	I^-
Ions formed by gaining two electrons	O^{2-}
	S^{2-}
	Se^{2-}
Ions formed by gaining three electrons	N^{3-}
	P^{3-}

Third, there are some exceptions to the previous point. A few elements, all of which are metals, can form more than one possible charge. For example, iron atoms can form 2+ cations or 3+ cations. Cobalt is another element that can form more than one possible charged ion (2+ and 3+), while lead can form 2+ or 4+ cations. Unfortunately, there is little understanding which two charges a metal atom may take, so it is best to just memorize the possible charges a particular element can have.

Note the convention for indicating an ion. The magnitude of the charge is listed as a right superscript next to the symbol of the element. If the charge is a single positive or negative one, the number 1 is not written; if the magnitude of the charge is greater than 1, then the number is written *before* the + or – sign. An element symbol without a charge written next to it is assumed to be the uncharged atom.

Naming an ion is straightforward. For a cation, simply use the name of the element and add the word *ion* (or if you want to be more specific, add *cation*) after the element's name. So Na^+ is the sodium ion; Ca^{2+} is the calcium ion. If the element has more than one possible charge, the value of the charge comes after the element name and before the word *ion*. Thus, Fe^{2+} is the iron two ion, while Fe^{3+} is the iron three ion. In print, we use roman numerals in parentheses to represent the charge on the ion; so these two iron ions would be represented as the iron(II) cation and the iron(III) cation, respectively.

For a monatomic anion, use the stem of the element name and append the suffix *-ide* to it, and then add *ion*. This is similar to how we named molecular compounds. Thus, Cl^- is the chloride ion, and N^{3-} is the nitride ion.

Example 3.3.1

Name each species.

a. O^{2-}
b. Co
c. Co^{2+}

Solution

a. This species has a 2– charge on it, so it is an anion. Anions are named using the stem of the element name with the suffix *-ide* added. This is the oxide anion.
b. Because this species has no charge, it is an atom in its elemental form. This is cobalt.
c. In this case, there is a 2+ charge on the atom, so it is a cation. We note from Table 3.3.1 that cobalt cations can have two possible charges, so the name of the ion must specify which charge the ion has. This is the cobalt(II) cation.

Exercise 3.3.1

Name each species.

a. P^{3-}
b. Sr^{2+}

Answers

a. the phosphide anion
b. the strontium cation

Ionic Formulas

Chemical formulas for ionic compounds are called ionic formulas. A proper ionic formula has a cation and an anion in it; an ionic compound is never formed between two cations only or two anions only. The key to writing proper ionic formulas is simple: the total positive charge must balance the total negative charge. Because the charges on the ions are characteristic, sometimes we have to have more than one of a cation or an anion to balance the overall positive and negative charges. It is conventional to use the lowest ratio of ions that are needed to balance the charges.

For example, consider the ionic compound between Na^+ and Cl^-. Each ion has a single charge, one positive and one negative, so we need only one ion of each to balance the overall charge. When writing the ionic formula, we follow two additional conventions: (1) write the formula for the cation first and the formula for the anion second, but (2) do not write the charges on the ions. Thus, for the compound between Na^+ and Cl^-, we have the ionic formula NaCl (Figure 3.3.1). The formula Na_2Cl_2 also has balanced charges, but the convention is to use the lowest ratio of ions, which would be one of each. (Remember from our conventions for writing formulas that we do not write a 1 subscript if there is only one atom of a particular element present.) For the ionic compound between magnesium cations (Mg^{2+}) and oxide anions (O^{2-}), again we need only one of each ion to balance the charges. By convention, the formula is MgO.



Figure 3.3.1: The ionic compound NaCl is table salt and is very common. (Public Domain; Pexels.)

For the ionic compound between Mg^{2+} ions and Cl^- ions, we now consider the fact that the charges have different magnitudes: 2+ on the magnesium ion and 1– on the chloride ion. To balance the charges with the lowest number of ions possible, we need to have two chloride ions to balance the charge on the one magnesium ion. Rather than write the formula MgClCl, we combine the

two chloride ions and write it with a 2 subscript: $MgCl_2$.

What is the formula $MgCl_2$ telling us? There are two chloride ions in the formula. Although chlorine as an element is a diatomic molecule, Cl_2, elemental chlorine is not part of this ionic compound. The chlorine is in the form of a negatively charged *ion*, not the neutral *element*. The 2 subscript is in the ionic formula because we need two Cl^- ions to balance the charge on one Mg^{2+} ion.

Example 3.3.2

Write the proper ionic formula for each of the two given ions.

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

Solution

a. We need two Cl^- ions to balance the charge on one Ca^{2+} ion, so the proper ionic formula is $CaCl_2$.
b. We need three F^- ions to balance the charge on the Al^{3+} ion, so the proper ionic formula is AlF_3.
c. With Al^{3+} and O^{2-}, note that neither charge is a perfect multiple of the other. This means we have to go to a least common multiple, which in this case will be six. To get a total of 6+, we need two Al^{3+} ions; to get 6–, we need three O^{2-} ions. Hence the proper ionic formula is Al_2O_3.

Exercise 3.3.2

Write the proper ionic formulas for each of the two given ions.

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

Answers

a. FeS
b. Fe_2S_3

Naming ionic compounds is simple: combine the name of the cation and the name of the anion, in both cases omitting the word *ion*. *Do not use numerical prefixes if there is more than one ion necessary to balance the charges.* NaCl is sodium chloride, a combination of the name of the cation (sodium) and the anion (chloride). MgO is magnesium oxide. $MgCl_2$ is magnesium chloride—*not* magnesium dichloride.

When naming ionic compounds whose cations can have more than one possible charge, we must also include the charge, in parentheses and in roman numerals, as part of the name. Hence FeS is iron(II) sulfide, while Fe_2S_3 is iron(III) sulfide. Again, no numerical prefixes appear in the name. The number of ions in the formula is dictated by the need to balance the positive and negative charges.

Example 3.3.3

Name each ionic compound.

a. $CaCl_2$
b. AlF_3
c. Co_2O_3

Solution

a. Using the names of the ions, this ionic compound is named calcium chloride. It is *not* calcium(II) chloride, because calcium forms only one cation when it forms an ion, and it has a characteristic charge of 2+.
b. The name of this ionic compound is aluminum fluoride.
c. We know that cobalt can have more than one possible charge; we just need to determine what it is. Oxide always has a 2– charge, so with three oxide ions, we have a total negative charge of 6–. This means that the two cobalt ions have to contribute 6+, which for two cobalt ions means that each one is 3+. Therefore, the proper name for this ionic compound is cobalt(III) oxide.

Exercise 3.3.3

Name each ionic compound.

a. Sc_2O_3
b. $AgCl$

Answers

a. scandium oxide
b. silver chloride

How do you know whether a formula—and by extension, a name—is for a molecular compound or for an ionic compound? Molecular compounds form between nonmetals and nonmetals, while ionic compounds form between metals and nonmetals. The periodic table can be used to determine which elements are metals and nonmetals.

There also exists a group of ions that contain more than one atom. These are called **polyatomic ions**. Table 3.3.2 lists the formulas, charges, and names of some common polyatomic ions. Only one of them, the ammonium ion, is a cation; the rest are anions. Most of them also contain oxygen atoms, so sometimes they are referred to as ***oxyanions***. Some of them, such as nitrate and nitrite, and sulfate and sulfite, have very similar formulas and names, so care must be taken to get the formulas and names correct. Note that the -ite polyatomic ion has one less oxygen atom in its formula than the -ate ion but with the same ionic charge.

Table 3.3.2: Common Polyatomic Ions

Name	Formula and Charge	Name	Formula and Charge
ammonium	NH_4^+	hydroxide	OH^-
acetate	$C_2H_3O_2^-$, or CH_3COO^-	nitrate	NO_3^-
bicarbonate (hydrogen carbonate)	HCO_3^-	nitrite	NO_2^-
bisulfate (hydrogen sulfate)	HSO_4^-	peroxide	O_2^{2-}
carbonate	CO_3^{2-}	perchlorate	ClO_4^-
chlorate	ClO_3^-	phosphate	PO_4^{3-}
chromate	CrO_4^{2-}	sulfate	SO_4^{2-}
cyanide	CN^-	sulfite	SO_3^{2-}

Name	Formula and Charge		Name	Formula and Charge
dichromate	$Cr_2O_7^{2-}$		triiodide	I_3^-

The naming of ionic compounds that contain polyatomic ions follows the same rules as the naming for other ionic compounds: simply combine the name of the cation and the name of the anion. Do not use numerical prefixes in the name if there is more than one polyatomic ion; the only exception to this is if the name of the ion itself contains a numerical prefix, such as dichromate or triiodide.

Writing the formulas of ionic compounds has one important difference. If more than one polyatomic ion is needed to balance the overall charge in the formula, enclose the formula of the polyatomic ion in parentheses and write the proper numerical subscript to the right and *outside* of the parentheses. Thus, the formula between calcium ions, Ca^{2+}, and nitrate ions, NO_3^-, is properly written $Ca(NO_3)_2$, not $CaNO_{32}$ or CaN_2O_6. Use parentheses where required. The name of this ionic compound is simply calcium nitrate.

Example 3.3.4

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

a. NH_4^+ and S^{2-}
b. Al^{3+} and PO_4^{3-}
c. Fe^{2+} and PO_4^{3-}

Solution

a. Because the ammonium ion has a 1+ charge and the sulfide ion has a 2– charge, we need two ammonium ions to balance the charge on a single sulfide ion. Enclosing the formula for the ammonium ion in parentheses, we have $(NH_4)_2S$. The compound's name is ammonium sulfide.
b. Because the ions have the same magnitude of charge, we need only one of each to balance the charges. The formula is $AlPO_4$, and the name of the compound is aluminum phosphate.
c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, we need three iron(II) ions, and to get 6–, we need two phosphate ions. The proper formula is $Fe_3(PO_4)_2$, and the compound's name is iron(II) phosphate.

Exercise 3.3.4

Write the proper formula and give the proper name for each ionic compound formed between the two listed ions.

a. NH_4^+ and PO_4^{3-}
b. Co^{3+} and NO_2^-

Answers

a. $(NH_4)_3PO_4$, ammonium phosphate
b. $Co(NO_2)_3$, cobalt(III) nitrite

Food and Drink Application: Sodium in Your Food

The element sodium, at least in its ionic form as Na^+, is a necessary nutrient for humans to live. In fact, the human body is approximately 0.15% sodium, with the average person having one-twentieth to one-tenth of a kilogram in their body at any given time, mostly in fluids outside cells and in other bodily fluids.

Sodium is also present in our diet. The common table salt we use on our foods is an ionic sodium compound. Many processed foods also contain significant amounts of sodium added to them as a variety of ionic compounds. Why are sodium compounds used so much? Usually sodium compounds are inexpensive, but, more importantly, most ionic sodium compounds dissolve easily. This allows processed food manufacturers to add sodium-containing substances to food mixtures and know that the compound will dissolve and distribute evenly throughout the food. Simple ionic compounds such as sodium nitrite ($NaNO_2$) are added to cured meats, such as bacon and deli-style meats, while a compound called sodium benzoate is added to many packaged foods as a preservative. Table 3.3.3 is a partial list of some sodium additives used in food. Some of them you may recognize after reading this chapter. Others you may not recognize, but they are all ionic sodium compounds with some negatively charged ion also present.

Table 3.3.3: Some Sodium Compounds Added to Food

Sodium Compound	Use in Food
Sodium acetate	preservative, acidity regulator
Sodium adipate	food acid
Sodium alginate	thickener, vegetable gum, stabilizer, gelling agent, emulsifier
Sodium aluminum phosphate	acidity regulator, emulsifier
Sodium aluminosilicate	anti-caking agent
Sodium ascorbate	antioxidant
Sodium benzoate	preservative
Sodium bicarbonate	mineral salt
Sodium bisulfite	preservative, antioxidant
Sodium carbonate	mineral salt
Sodium carboxymethylcellulose	emulsifier
Sodium citrates	food acid
Sodium dehydroacetate	preservative
Sodium erythorbate	antioxidant
Sodium erythorbin	antioxidant
Sodium ethyl para-hydroxybenzoate	preservative
Sodium ferrocyanide	anti-caking agent
Sodium formate	preservative
Sodium fumarate	food acid
Sodium gluconate	stabilizer
Sodium hydrogen acetate	preservative, acidity regulator
Sodium hydroxide	mineral salt
Sodium lactate	food acid
Sodium malate	food acid

Sodium Compound	Use in Food
Sodium metabisulfite	preservative, antioxidant, bleaching agent
Sodium methyl para-hydroxybenzoate	preservative
Sodium nitrate	preservative, color fixative
Sodium nitrite	preservative, color fixative
Sodium orthophenyl phenol	preservative
Sodium propionate	preservative
Sodium propyl para-hydroxybenzoate	preservative
Sodium sorbate	preservative
Sodium stearoyl lactylate	emulsifier
Sodium succinates	acidity regulator, flavor enhancer
Sodium salts of fatty acids	emulsifier, stabilizer, anti-caking agent
Sodium sulfite	mineral salt, preservative, antioxidant
Sodium sulfite	preservative, antioxidant
Sodium tartrate	food acid
Sodium tetraborate	preservative

The use of so many sodium compounds in prepared and processed foods has alarmed some physicians and nutritionists. They argue that the average person consumes too much sodium from his or her diet. The average person needs only about 500 mg of sodium every day; most people consume more than this—up to 10 times as much. Some studies have implicated increased sodium intake with high blood pressure; newer studies suggest that the link is questionable. However, there has been a push to reduce the amount of sodium most people ingest every day: avoid processed and manufactured foods, read labels on packaged foods (which include an indication of the sodium content), avoid oversalting foods, and use other herbs and spices besides salt in cooking.



Nutrition Facts

Serving Size 8 oz (227 g/8 oz)
Servings Per Container About 3

Amount Per Serving

Calories 180 Calories from Fat 60

	% Daily Value*
Total Fat 6g	**10%**
Saturated Fat 1g	**5%**
Trans Fat 0g	
Cholesterol 5mg	**2%**
Sodium 75mg	**3%**
Total Carbohydrate 26g	**9%**
Dietary Fiber 5g	**19%**
Sugars 11g	
Protein 8g	

Vitamin A 60% • Vitamin C 70%
Calcium 8% • Iron 10%

* Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs.

	Calories	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	25g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

Calories per gram:
Fat 9 • Carbohydrate 4 • Protein 4

Figure 3.3.2: Food labels include the amount of sodium per serving. This particular label shows that there are 75 mg of sodium in one serving of this particular food item.

Key Takeaways

- Ions form when atoms lose or gain electrons.
- Ionic compounds have positive ions and negative ions.
- Ionic formulas balance the total positive and negative charges.
- Ionic compounds have a simple system of naming.
- Groups of atoms can have an overall charge and make ionic compounds.

Learning Objectives

- Recognize polyatomic ions in chemical formulas.
- Write the correct formula for ionic compounds with polyatomic ions.
- Use the rules for naming ionic compounds with polyatomic ions.

Polyatomic Ions

Some ions consist of groups of atoms bonded together and have an overall electric charge. Because these ions contain more than one atom, they are called polyatomic ions. Polyatomic ions have characteristic formulas, names, and charges that should be memorized. For example, NO_3^- is the nitrate ion; it has one nitrogen atom and three oxygen atoms and an overall 1– charge. Table 3.3.1 lists the most common polyatomic ions.

Table 3.3.1 Some Polyatomic Ions

Name	Formula
ammonium ion	NH_4^+
acetate ion	$C_2H_3O_2^-$ (also written $CH_3CO_2^-$)
carbonate ion	CO_3^{2-}
chromate ion	CrO_4^{2-}
dichromate ion	$Cr_2O_7^{2-}$
hydrogen carbonate ion (bicarbonate ion)	HCO_3^-
cyanide ion	CN^-
hydroxide ion	OH^-
nitrate ion	NO_3^-
nitrite ion	NO_2^-
permanganate ion	MnO_4^-
phosphate ion	PO_4^{3-}
hydrogen phosphate ion	HPO_4^{2-}
dihydrogen phosphate ion	$H_2PO_4^-$
sulfate ion	SO_4^{2-}
hydrogen sulfate ion (bisulfate ion)	HSO_4^-
sulfite ion	SO_3^{2-}

The rule for constructing formulas for ionic compounds containing polyatomic ions is the same as for formulas containing monatomic (single-atom) ions: the positive and negative charges must balance. If more than one of a particular polyatomic ion is needed to balance the charge, the *entire formula* for the polyatomic ion must be enclosed in parentheses, and the numerical subscript is placed *outside* the parentheses. This is to show that the subscript applies to the entire polyatomic ion. An example is $Ba(NO_3)_2$.

Writing Formulas for Ionic Compounds Containing Polyatomic Ions

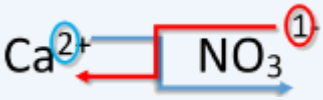
Writing a formula for ionic compounds containing polyatomic ions also involves the same steps as for a binary ionic compound. Write the symbol and charge of the cation followed by the symbol and charge of the anion.

Example 3.3.1: Calcium Nitrate

Write the formula for calcium nitrate.

Solution

Example 3.3.1: Steps For Problem Solving, Write the formula for calcium nitrate

Criss Cross Method	Write the formula for calcium nitrate
1. Write the symbol and charge of the cation (metal) first and the anion (nonmetal) second.	[redacted] NO_3^-
2. Transpose only the number of the positive charge to become the subscript of the anion and the number only of the negative charge to become the subscript of the cation.	 The 2+ charge on Ca becomes the subscript of NO3 and the 1- charge on NO3 becomes the subscript of Ca.
3. Reduce to the lowest ratio.	$Ca_1(NO_3)_2$
4. Write the final formula. Leave out all subscripts that are 1. If there is only 1 of the polyatomic ion, leave off parentheses.	$Ca(NO_3)_2$

Example 3.3.2

Write the chemical formula for an ionic compound composed of the potassium ion and the sulfate ion

Solution

Example 3.3.2: Explanation for Writing the Chemical Formula for an Ionic Compound Composed of the Potassium Ion and the Sulfate Ion

Explanation	Answer
Potassium ions have a charge of 1+, while sulfate ions have a charge of 2–. We will need two potassium ions to balance the charge on the sulfate ion, so the proper chemical formula is K_2SO_4.	K_2SO_4

Exercise 3.3.1

Write the chemical formula for an ionic compound composed of each pair of ions.

a. the magnesium ion and the carbonate ion
b. the aluminum ion and the acetate ion

Answer a: available at the end of this section.

Naming Ionic Compounds with Polyatomic Ions

The process of naming ionic compounds with polyatomic ions is the same as naming binary ionic compounds. The cation is named first, followed by the anion. One example is the ammonium Sulfate compound in Figure 3.3.6.

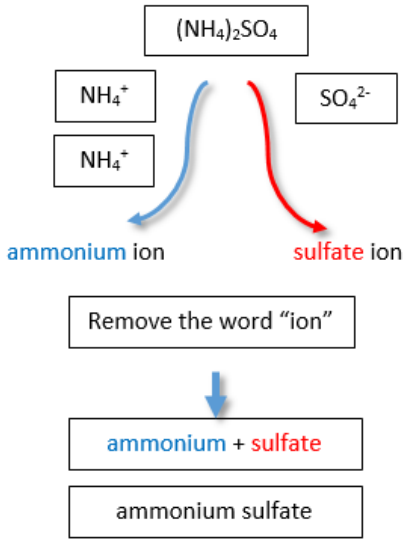


Figure 3.3.2 *Naming Ionic Compounds with Polyatomic Ions*

Example 3.3.3: Naming Ionic Compounds

Write the proper name for each ionic compound.

a. (NH4)2S
b. AlPO4,
c. Fe3(PO4)2

Solution

Example 3.3.3: Explanation for Naming Ionic Compounds

Ionic Compound	Explanation	Answer
a. **(NH4)2S**	a. The ammonium ion has a 1+ charge and the sulfide ion has a 2– charge. Two ammonium ions need to balance the charge on a single sulfide ion. The compound's name is ammonium sulfide.	ammonium sulfide
b. AlPO4,	b. The ions have the same magnitude of charge, one of each (ion) is needed to balance the charges. The name of the compound is aluminum phosphate.	aluminum phosphate

Ionic Compound	Explanation	Answer
c. $Fe_3(PO_4)_2$	c. Neither charge is an exact multiple of the other, so we have to go to the least common multiple of 6. To get 6+, three iron(II) ions are needed, and to get 6–, two phosphate ions are needed . The compound's name is iron(II) phosphate.	iron(II) phosphate

Exercise 3.3.2

Write the proper name for each ionic compound.

a. $(NH_4)_3PO_4$
b. $Co(NO_2)_3$

Answer a: available at the end of this section.

Summary

- Formulas for ionic compounds contain the symbols and number of each atom and/or polyatomic ion present in a compound in the lowest whole number ratio.
- The process of naming ionic compounds with polyatomic ions are the same as binary ionic compounds.

Contributors and Attributions

-
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Answers

Exercise 3.3.1
iron(II) ion
iron(III) ion
barium ion

Exercise 3.3.2
F^-
Sn^{2+}
K^+

Exercise 3.3.3
silver iodide
magnesium oxide
calcium phosphide

Exercise 3.3.4
gold(III) chloride
lead(IV) oxide
copper(II) oxide

Exercise 3.3.5

zinc bromide

Exercise 3.3.5

aluminum oxide

Exercise 3.3.5

gold(III) fluoride or auric fluoride

Exercise 3.3.5

silver fluoride

Exercise 3.3.6

CaO

CuS

Cu_2S

Exercise 3.3.7

NaBr

LiCl

MgO

Exercise 3.3.8

not ionic

ionic

Exercise 3.3.1

$MgCO_3$

$Al(CH_3COO)_3$

Exercise 3.3.2

ammonium phosphate

cobalt(III) nitrite

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3.4: Acid Nomenclature

Learning Objectives

- Define *acid*.
- Name a simple acid.

There is one other group of compounds that is important to us—acids—and these compounds have interesting chemical properties. Initially, we will define an acid as an ionic compound of the H^+ cation dissolved in water. To indicate that something is dissolved in water, we will use the phase label (aq) next to a chemical formula (where aq stands for "aqueous," a word that describes something dissolved in water). If the formula does not have this label, then the compound is treated as a molecular compound rather than an acid.

Acids have their own nomenclature system. If an acid is composed of only hydrogen and one other element, the name is *hydro-* + the stem of the other element + *-ic acid*. For example, the compound HCl(aq) is hydrochloric acid, while H_2S(aq) is hydrosulfuric acid. If these acids were not dissolved in water, the compounds would be called hydrogen chloride and hydrogen sulfide, respectively. Both of these substances are well known as molecular compounds; when dissolved in water, however, they are treated as acids.

If a compound is composed of hydrogen ions and a polyatomic anion, then the name of the acid is derived from the stem of the polyatomic ion's name. Typically, if the anion name ends in -ate, the name of the acid is the stem of the anion name plus *-ic acid*; if the related anion's name ends in -ite, the name of the corresponding acid is the stem of the anion name plus *-ous acid*. Table 3.4.1 lists the formulas and names of a variety of acids that you should be familiar with. You should recognize most of the anions in the formulas of the acids.

Table 3.4.1 Names and Formulas of Acids

Formula	Name
$HC_2H_3O_2$	acetic acid
$HClO_3$	chloric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydriodic acid
HF	hydrofluoric acid
HNO_3	nitric acid
$H_2C_2O_4$	oxalic acid
$HClO_4$	perchloric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid
H_2SO_3	sulfurous acid
Note: The "aq" label is omitted for clarity.	

Example 3.4.1

Name each acid without consulting Table 3.9.

1. HBr
2. H_2SO_4

Solution

1. As a binary acid, the acid's name is *hydro-* + stem name + *-ic acid*. Because this acid contains a bromine atom, the name is hydrobromic acid.
2. Because this acid is derived from the sulfate ion, the name of the acid is the stem of the anion name + *-ic acid*. The name of this acid is sulfuric acid.

Exercise 3.4.1

Name each acid.

a. HF
b. HNO_2

Answer a available at the end of this section.

All acids have some similar properties. For example, acids have a sour taste; in fact, the sour taste of some of our foods, such as citrus fruits and vinegar, is caused by the presence of acids in food. Many acids react with some metallic elements to form metal ions and elemental hydrogen. Acids make certain plant pigments change colors; indeed, the ripening of some fruits and vegetables is caused by the formation or destruction of excess acid in the plant. In Chapter 12, we will explore the chemical behavior of acids.

Acids are very prevalent in the world around us. We have already mentioned that citrus fruits contain acid; among other compounds, they contain citric acid, $H_3C_6H_5O_7(aq)$. Oxalic acid, $H_2C_2O_4(aq)$, is found in spinach and other green leafy vegetables. Hydrochloric acid not only is found in the stomach (stomach acid), but also can be bought in hardware stores as a cleaner for concrete and masonry. Phosphoric acid is an ingredient in some soft drinks.

Key Takeaways

- An acid is a compound of the H^+ ion dissolved in water.
- Acids have their own naming system.
- Acids have certain chemical properties that distinguish them from other compounds.

Answers

Exercise 3.4.1

hydrofluoric acid
nitrous acid

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3.5: Binary Covalent Compound Nomenclature

Learning Objectives

- Define *molecule*.
- Name simple molecules based on their formulas.
- Determine a formula of a molecule based on its name.

There are many substances that exist as two or more atoms connected together so strongly that they behave as a single particle. These multiatom combinations are called molecules. A molecule is the smallest part of a substance that has the physical and chemical properties of that substance. In some respects, a molecule is similar to an atom. A molecule, however, is composed of more than one atom.

Table 3.5.1: Elements That Exist as Diatomic Molecules

Hydrogen (H_2)	Oxygen (O_2)	Nitrogen (N_2)	Fluorine (F_2)
Chlorine (Cl_2)	Bromine (Br_2)	Iodine (I_2)	

Some elements exist naturally as molecules. For example, hydrogen and oxygen exist as two-atom molecules. Other elements also exist naturally as diatomic molecules (Table 3.5.1). As with any molecule, these elements are labeled with a molecular formula, a formal listing of what and how many atoms are in a molecule. (Sometimes only the word *formula* is used, and its meaning is inferred from the context.) For example, the molecular formula for elemental hydrogen is H2, with H being the symbol for hydrogen and the subscript 2 implying that there are two atoms of this element in the molecule. Other diatomic elements have similar formulas: O2, N2, and so forth. Other elements exist as molecules—for example, sulfur normally exists as an eight-atom molecule, S8, while phosphorus exists as a four-atom molecule, P4 (Figure 3.5.1). Otherwise, we will assume that elements exist as individual atoms, rather than molecules. It is assumed that there is only one atom in a formula if there is no numerical subscript on the right side of an element's symbol.

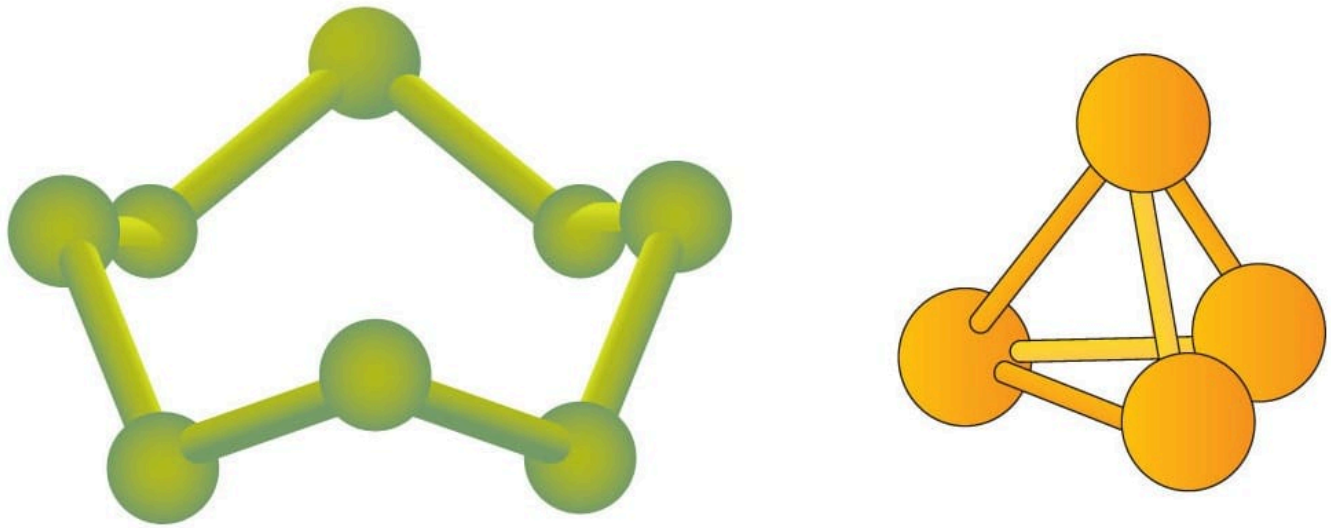


Figure 3.5.1: Molecular Art of S8 and P4 Molecules. If each green ball represents a sulfur atom, then the diagram on the left represents an S8 molecule. The molecule on the right shows that one form of elemental phosphorus exists, as a four-atom molecule.

Figure 3.5.1 shows two examples of how we will be representing molecules in this text. An atom is represented by a small ball or sphere, which generally indicates where the nucleus is in the molecule. A cylindrical line connecting the balls represents the connection between the atoms that make this collection of atoms a molecule. This connection is called a chemical bond and is the connection between two atoms in a molecule.

Many compounds exist as molecules. In particular, when nonmetals connect with other nonmetals, the compounds typically exist as molecules. (Compounds between a metal and a nonmetal are different and will be considered in Section 3.4.) In some cases, there are many different kinds of molecules that can be formed between any given elements, with all the different molecules having different chemical and physical properties. How do we tell them apart?

The answer is a very specific system of naming compounds, called *chemical nomenclature*. By following the rules of nomenclature, each and every compound has its own unique name, and each name refers to one and only one compound. Here, we will start with relatively simple molecules that have only two elements in them, the so-called *binary compounds*:

1. Identify the elements in the molecule from its formula.
2. Begin the name with the element name of the first element. If there is more than one atom of this element in the molecular formula, use a numerical prefix to indicate the number of atoms, as listed in Table 3.5.2. *Do not use the prefix* mono- *if there is only one atom of the first element.*
3. Name the second element by using three pieces:
   - a numerical prefix indicating the number of atoms of the second element, plus
   - the stem of the element name (e.g., *ox* for oxygen, *chlor* for chlorine, etc.), plus
   - the suffix *-ide*
4. Combine the two words, leaving a space between them.

Table 3.5.2: Numerical Prefixes Used in Naming Molecular Compounds

The Number of Atoms of an Element	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

Let us see how these steps work for a molecule whose molecular formula is SO_2, which has one sulfur atom and two oxygen atoms—this completes step 1. According to step 2, we start with the name of the first element—sulfur. Remember, we do not use the *mono-* prefix for the first element. Now for step 3, we combine the numerical prefix *di-* (see Table 3.5.2) with the stem *ox-* and the suffix *-ide*, to make *dioxide*. Bringing these two words together, we have the unique name for this compound—sulfur dioxide.

Why all this trouble? There is another common compound consisting of sulfur and oxygen whose molecular formula is SO_3, so the compounds need to be distinguished. SO_3 has three oxygen atoms in it, so it is a different compound with different chemical and physical properties. The system of chemical nomenclature is designed to *give this compound its own unique name*. Its name, if you go through all the steps, is sulfur trioxide. Different compounds have different names.

In some cases, when a prefix ends in *a* or *o;* and the element name begins with *o,* we drop the *a* or *o* on the prefix. So we see *monoxide* or *pentoxide*, rather than *monooxide* or *pentaoxide* in molecule names.

Example 3.5.1

Name each molecule.

a. PF_3
b. CO

c. Se_2Br_2

Solution

a. A molecule with a single phosphorus atom and three fluorine atoms is called phosphorus trifluoride.
b. A compound with one carbon atom and one oxygen atom is properly called carbon monoxide, not carbon monooxide.
c. There are two atoms of each element, selenium and bromine. According to the rules, the proper name here is *diselenium dibromide*.

Exercise 3.5.1

Name each molecule.

a. SF_4
b. P_2S_5

Answer a available at the end of this section.

Answer b available at the end of this section.

One great thing about this system is that it works both ways. From the name of a compound, you should be able to determine its molecular formula. Simply list the element symbols, with a numerical subscript if there is more than one atom of that element, in the order of the name (we do not use the subscript 1 if there is only one atom of the element present; 1 is implied). From the name *nitrogen trichloride*, you should be able to get NCl_3 as the formula for this molecule. From the name *diphosphorus pentoxide*, you should be able to get the formula P_2O_5 (note the numerical prefix on the first element, indicating there is more than one atom of phosphorus in the formula).

Example 3.5.2

Give the formula for each molecule.

a. carbon tetrachloride
b. silicon dioxide
c. trisilicon tetranitride

Solution

a. The name *carbon tetrachloride* implies one carbon atom and four chlorine atoms, so the formula is CCl_4.
b. The name *silicon dioxide* implies one silicon atom and two oxygen atoms, so the formula is SiO_2.
c. We have a name that has numerical prefixes on both elements. *Tri-* means three, and *tetra-* means four, so the formula of this compound is Si_3N_4.

Exercise 3.5.2

Give the formula for each molecule.

a. disulfur difluoride
b. iodine pentabromide

Answer a available at the end of this section.

Answer b available at the end of this section.

Some simple molecules have common names that we use as part of the formal system of chemical nomenclature. For example, H_2O is given the name *water*, not *dihydrogen monoxide*. NH_3 is called *ammonia*, while CH_4 is called *methane*. We will occasionally see other molecules that have common names; we will point them out as they occur.

Key Takeaways

- Molecules are groups of atoms that behave as a single unit.
- Some elements exist as molecules: hydrogen, oxygen, sulfur, and so forth.
- There are rules that can express a unique name for any given molecule, and a unique formula for any given name.

Answers

Exercise 3.5.1

sulfur tetrafluoride

Exercise 3.5.1

diphosphorus pentasulfide

Exercise 3.5.2

S_2F_2

Exercise 3.5.2

IBr_5

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4: CHEMICAL STRUCTURES AND INTERACTIONS (UNDER CONSTRUCTION)

4.1: Chemical Bonding

4.2: Lewis Theory of Bonding

4.3: Lewis Structures of Ionic Compounds

4.4: Lewis Structures of Covalent Compounds and Ions

4.5: VSEPR Theory

4.6: Polarity

4.7: Intermolecular Forces

4.8: Impact of Intermolecular Forces on Physical Properties

4.9: Impact of Solutes on Physical Properties

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4.1: Chemical Bonding

Learning Objectives

- Explain the bonding nature of ionic compounds.
- Relate microscopic bonding properties to macroscopic solid properties.

Neutral atoms and their associated ions have very different physical and chemical properties. Sodium *atoms* form sodium metal, a soft, silvery-white metal that burns vigorously in air and reacts explosively with water. Chlorine *atoms* form chlorine gas, Cl_2, a yellow-green gas that is extremely corrosive to most metals and very poisonous to animals and plants. The vigorous reaction between the elements sodium and chlorine forms the white, crystalline compound sodium chloride, common table salt, which contains sodium *cations* and chloride *anions* (Figure $4.1.1$). The compound composed of these ions exhibits properties entirely different from the properties of the elements sodium and chlorine. Chlorine is poisonous, but sodium chloride is essential to life; sodium atoms react vigorously with water, but sodium chloride simply dissolves in water.



Figure $4.1.1$ (a) Sodium is a soft metal that must be stored in mineral oil to prevent reaction with air or water. (b) Chlorine is a pale yellow-green gas. (c) When combined, they form white crystals of sodium chloride (table salt). (credit a: modification of work by "Jurii"/Wikimedia Commons) Commons)

Ionic Bonds

When two atoms approach each other, they have the potential to bond (or connect). If a metal and a nonmetal interact, then an ionic bond will result. These types of bonds involve the metal donating it(s) valence electron(s) to a nonmetal. As the electronic transfer occurs, both atoms will achieve more stable confirmations. The end result will be a less reactive compound. These type of species are composed of both cations and anions. In addition, they are crystalline and solid in nature (Figure $4.1.2$). A few examples of real-world ionic compounds would be NaCl (table salt) and NaF (active ingredient in toothpaste).

The formula for table salt is NaCl. It is the result of Na^+ ions and Cl^- ions bonding together (Figure $4.1.3$). If sodium metal and chlorine gas mix under the right conditions, they will form salt. The sodium loses an electron, and the chlorine gains that electron. This reaction is highly favorable because of the electrostatic attraction between the particles. In the process, a great amount of light and heat is released. The resulting salt is mostly unreactive — it is stable. It will not undergo any explosive reactions, unlike



the sodium and chlorine that it is made of.

Figure 4.1.2 *NaCl crystals.*

(Public Domain; NASA).

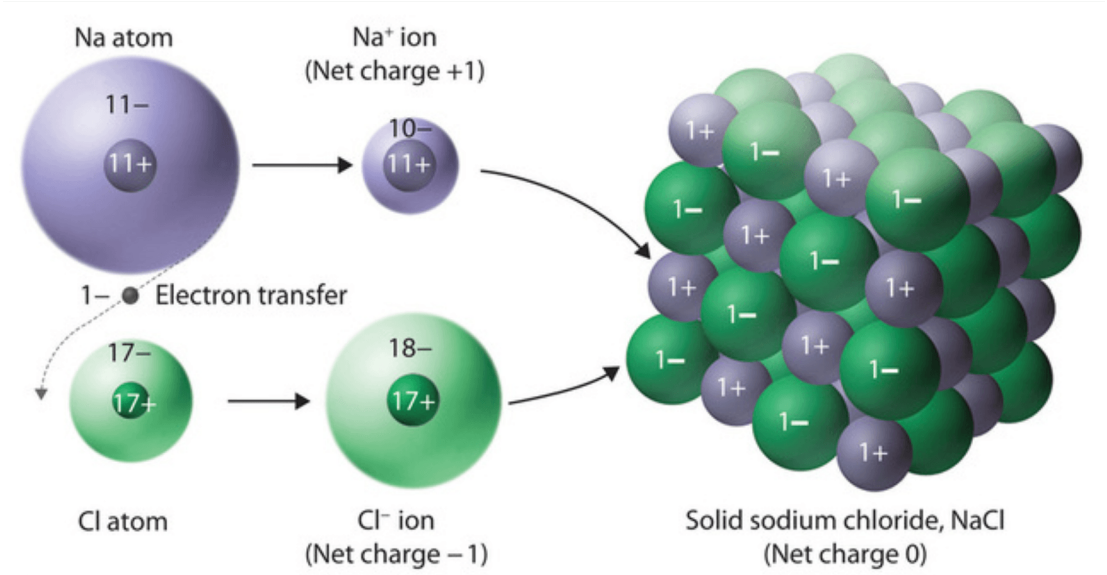
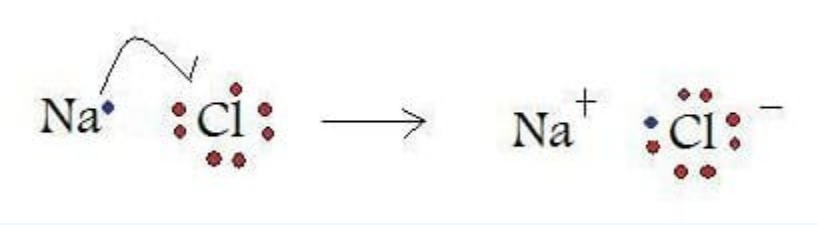


Figure 4.1.3 The formation of sodium chloride from the attraction of sodium and chloride ions.

The reaction is represented with Lewis dot symbols below.

Chloride Salts



The sodium atom is donating its 1 valence electron to the chlorine atom. This creates a sodium cation and a chlorine anion. Notice that the net charge of the resulting compound is 0.

Summary

- An ionic bond is formed when a metal donates it(s) valence electron(s) to a nonmetal.
- The resulting ionic compound is more stable and less reactive.

References

1. Vollhardt, K. Peter C., and Neil E. Schore. Organic Chemistry Structure and Function. New York: W. H. Freeman, 2007.
2. Petrucci, Ralph H. General Chemistry: Principles and Modern Applications. Upper Saddle River, NJ: Pearson Education, 2007.
3. Brown, Theodore L., Eugene H. Lemay, and Bruce E. Bursten. Chemistry: The Central Science. 6th ed. Englewood Cliffs, NJ: Prentice Hall, 1994.

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- TextMap: Chemistry a Central Science (Brown et al.)

- Marisa Alviar-Agnew (Sacramento City College)

Learning Objectives

- Define covalent bond.
- Illustrate covalent bond formation with Lewis electron dot diagrams.

Ionic bonding typically occurs when it is easy for one atom to lose one or more electrons and another atom to gain one or more electrons. However, some atoms won't give up or gain electrons easily. Yet they still participate in compound formation. How? There is another mechanism for obtaining a complete valence shell: *sharing* electrons. When electrons are shared between two atoms, they make a bond called a **covalent bond**.

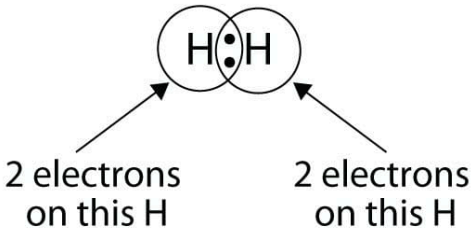
Chemists frequently use Lewis diagrams to represent covalent bonding in molecular substances. For example, the Lewis diagrams of two separate hydrogen atoms are as follows:



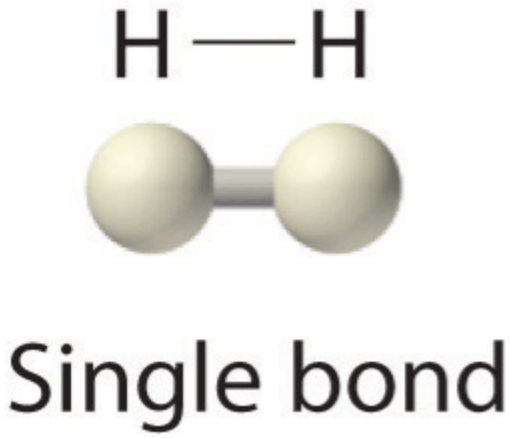
The Lewis diagram of two hydrogen atoms sharing electrons looks like this:



We can use circles to show that each H atom has two electrons around the nucleus, completely filling each atom's valence shell:



Because each H atom has a filled valence shell, this bond is stable, and we have made a diatomic hydrogen molecule. (This explains why hydrogen is one of the diatomic elements.) For simplicity's sake, it is not unusual to represent the covalent bond with a dash, instead of with two dots:



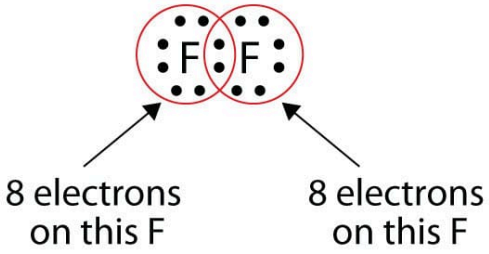
Because two atoms are sharing one pair of electrons, this covalent bond is called a **single bond**. As another example, consider fluorine. F atoms have seven electrons in their valence shell:



These two atoms can do the same thing that the H atoms did; they share their unpaired electrons to make a covalent bond.



Note that each F atom has a complete octet around it now:



We can also write this using a dash to represent the shared electron pair:

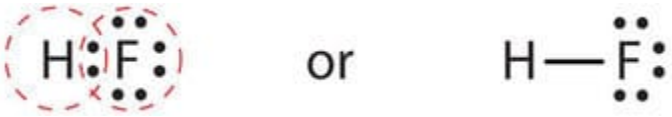


There are two different types of electrons in the fluorine diatomic molecule. The **bonding electron pair** makes the covalent bond. Each F atom has three other pairs of electrons that do not participate in the bonding; they are called **lone pair electrons**. Each F atom has one bonding pair and three lone pairs of electrons.

Covalent bonds can be made between different elements as well. One example is HF. Each atom starts out with an odd number of electrons in its valence shell:



The two atoms can share their unpaired electrons to make a covalent bond:



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

Example 4.1.1:

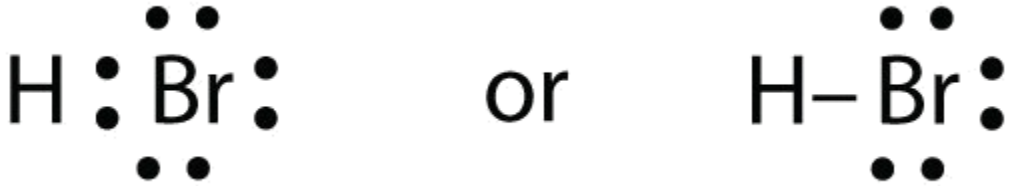
Use Lewis electron dot diagrams to illustrate the covalent bond formation in HBr.

Solution

HBr is very similar to HF, except that it has Br instead of F. The atoms are as follows:



The two atoms can share their unpaired electron:



Exercise 4.1.1

Use Lewis electron dot diagrams to illustrate the covalent bond formation in Cl_2.

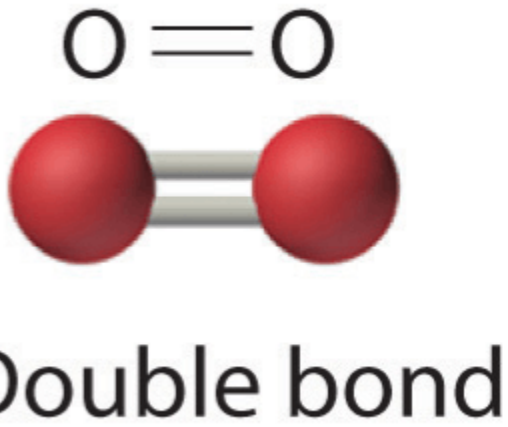
Answer: available at the end of this section.

When working with covalent structures, it sometimes looks like you have leftover electrons. You apply the rules you learned so far and there are still some electrons hanging out there unattached. You can't just leave them there. So where do you put them?

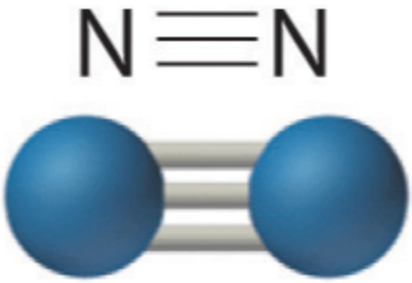
Multiple Covalent Bonds

The sharing of a pair of electrons represents a single covalent bond, usually just referred to as a ***single bond***. However, in many molecules atoms attain complete octets by sharing more than one pair of electrons between them:

- Two electron pairs shared a ***double bond***



- Three electron pairs shared a ***triple bond***



Triple bond

Because each nitrogen contains 5 valence electrons, they need to share 3 pairs to each achieve a valence octet. N_2 is fairly inert, due to the strong triple bond between the two nitrogen atoms.



Summary

- Covalent bonds are formed when atoms share electrons.
- Lewis electron dot diagrams can be drawn to illustrate covalent bond formation.
- Double bonds or triple bonds between atoms may be necessary to properly illustrate the bonding in some molecules.

Contributors and Attributions

-
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Answers

Exercise 4.1.1



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4.2: Lewis Theory of Bonding

Learning Objective

- Draw a Lewis electron dot diagram for an atom.
- Know the importance of Lewis dot in bonding.

At the beginning of the 20th century, an American physical chemist G. N. Lewis (1875–1946) devised a system of symbols—now called *Lewis electron dot symbols* (often shortened to *Lewis dot symbols*) that can be used for predicting the number of bonds formed by most elements in their compounds.

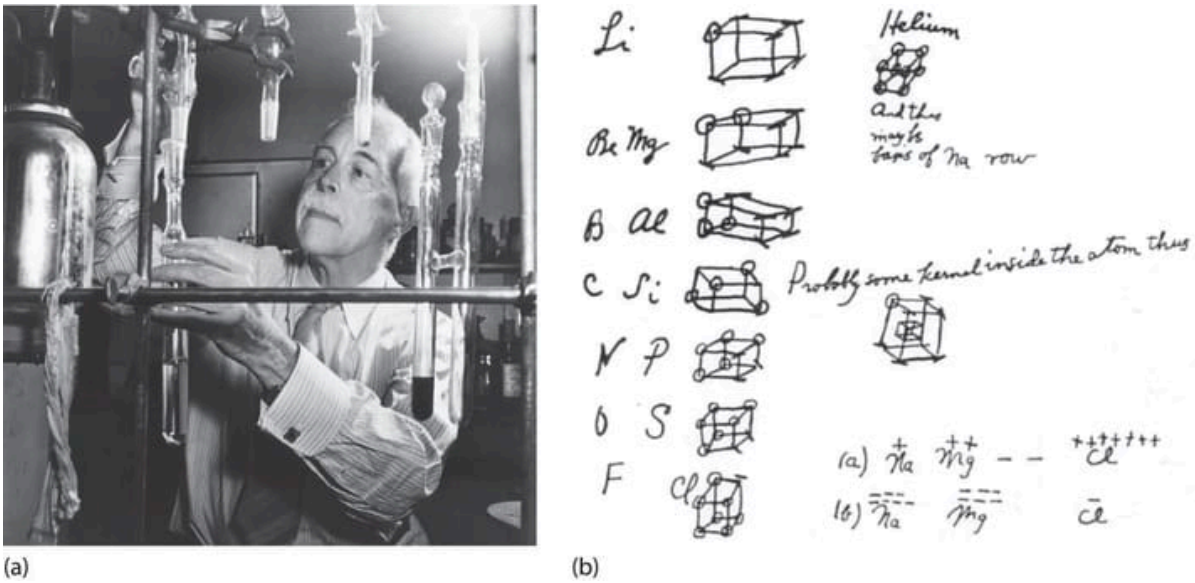
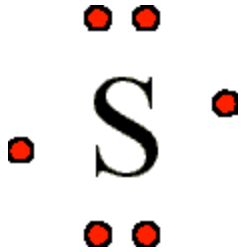


Figure 4.2.1 G. N. Lewis and the Octet Rule. (a) Lewis is working in the laboratory. (b) In Lewis's original sketch for the octet rule, he initially placed the electrons at the corners of a cube rather than placing them as we do now.

Lewis Dot symbols

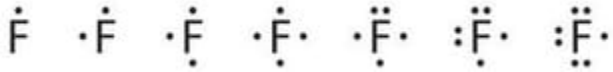
- convenient representation of valence electrons
- allows scientists to keep track of valence electrons during bond formation
- consists of the atomic symbol for the element plus a dot for each **valence** electron

To write an element's Lewis dot symbol, we place dots representing its valence electrons, one at a time, around the element's chemical symbol. Up to four dots are placed above, below, to the left, and to the right of the symbol (in any order, as long as elements with four or fewer valence electrons have no more than one dot in each position). The next dots, for elements with more than four valence electrons, are again distributed one at a time, each paired with one of the first four. For example, the element sulfur has six valence electrons (note roman numeral above group on the periodic table) and its Lewis symbol would be:



S two dots above two dots below and 1 dot left and right

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



Dot symbols for fluorine

Lewis eventually published his theory of chemical bonding in 1916. Formation of chemical bonds to complete the requirement of eight electrons for the atom becomes a natural tendency. Lewis dot symbols of the first two periods are given here to illustrate this point. In fact, the entire group (column) of elements have the same Lewis dot symbols, because they have the same number of valence electrons.

By going through the periodic table, we see that the Lewis electron dot symbols of atoms will never have more than eight dots around the atomic symbo see Table 4.2.1.

Table 4.2.1 *Lewis Dot Symbols of the Main Group Elements*

H·							He:
Li·	·Be ·	**\(\ce{ \cdot \underset{\:}{\overset{\Large{\cdot}}{B}} \cdot}\)**	·Ċ ·	:Ṅ·	:Ȯ·	:F̈·	:N̈e:
Na K Rb Cs	Mg Ca Sr Ba	Al Ga In Tl	Si Ge Sn Pb	P As Sb Bi	S Se Te Po	Cl Br I At	Ar Kr At Rn

Example 4.2.1: Lewis Dot Diagrams

What is the Lewis electron dot diagram for each element?

a. aluminum
b. selenium

Solution

a. The valence electron configuration for aluminum is $3s^2 3p^1$. So it would have three dots around the symbol for aluminum, two of them paired to represent the 3s electrons:

$$\dot{Al}:$$

2. The valence electron configuration for selenium is $4s^2 4p^4$. In the highest-numbered shell, the n = 4 shell, there are six electrons. Its electron dot diagram is as follows:

$$\cdot\dot{\underset{\cdot\cdot}{Se}}:$$

Exercise 4.2.1

What is the Lewis electron dot diagram for each element?

a. phosphorus
b. argon

Answer a available at the end of this section.

Summary

- Lewis electron dot diagrams use dots to represent valence electrons around an atomic symbol.
- Lewis dot symbols can be used to predict the number of bonds formed by elements in a compound.

Contributors and Attributions

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Interactive periodic table of the elements, in pictures and words

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Learning Objectives

- Know the classification of the different elements.
- Define valence electrons.
- Describe how the number of valence electrons are related the arrangement of the main group elements and its chemical behavior.

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). In and atomic , the **main group** is the group of elements whose lightest members are represented by , , , , , , , and as arranged in the of the elements. The main group includes the elements (except , which is sometimes not included) in groups 1 and 2, and groups 13 to 18. Main-group elements (with some of the lighter) are the most abundant elements on , in the , and in the . They are sometimes also called the **representative elements**. In older nomenclature the main-group elements are groups IA and IIA, and groups IIIB to 0 (CAS groups IIIA to VIIIA). Group 12 is labelled as group IIB in both systems. Group 3 is labelled as group

IIIA in the older nomenclature (CAS group IIIB).

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.

One way to categorize the elements of the periodic table is shown in Figure 4.2.1. The first two columns on the left and the last six columns on the right as mentioned earlier are 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**.

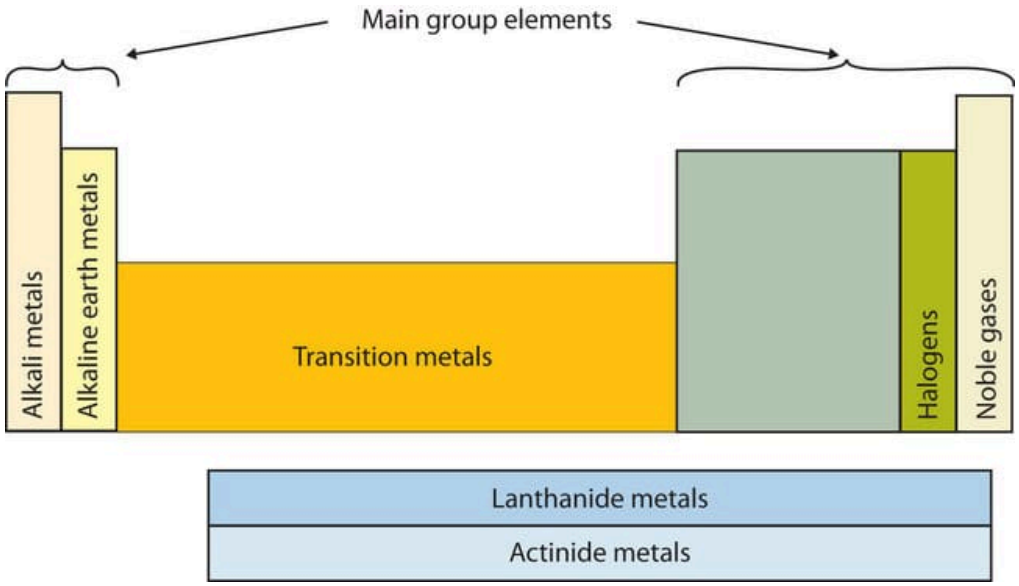


Figure 4.2.1: Special names for sections of the periodic table.

Family Features: Outer Electron Configurations

Valence Electrons

The valence shell is the outermost shell of an atom in its uncombined state, which contains the electrons most likely to account for the nature of any reactions involving the atom and of the bonding interactions it has with other atoms. **Valence electrons** are the outer-shell electrons of an atom. These are electrons that can participate in the formation of a chemical bond. The presence of valence electrons can determine the element's chemical properties and whether it may bond with other elements.

An atom with a closed shell of valence electrons tends to be chemically inert. An atom with one or two valence electrons more than a closed shell is highly reactive, because the extra valence electrons are easily removed to form a positive ion. An atom with one or two valence electrons fewer than a closed shell is also highly reactive, because of a tendency either to gain the missing valence electrons (thereby forming a negative ion), or to share valence electrons (thereby forming a covalent bond).

The number of valence electrons of an element can be determined by the periodic table group (vertical column) in which the element is categorized (Table 4.2.1). With the exception of groups 3–12 (the transition metals), the units digit of the group number identifies how many valence electrons are associated with a neutral atom of an element listed under that particular column.

Table 4.2.1: Number of Valence Electrons for Main Periodic Table Groups.

Periodic table group	Valence electrons
Group 1 (I) (alkali metals)	1
Group 2 (II) (alkaline earth metals)	2
Groups 3-12 (transition metals)	2*

Group 13 (III) (boron group)	3
Group 14 (IV) (carbon group)	4
Group 15 (V) (pnictogens)	5
Group 16 (VI) (chalcogens)	6
Group 17 (VII) (halogens)	7
Group 18 (VIII or 0) (noble gases)	8**

**** The general method for counting valence electrons is generally not useful for transition metals.***

***** Except for helium, which has only two valence electrons.***

Valence Electrons and the Periodic Table

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 4.2.1 *Finding the valence electron for an element.*

Family Groups

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.

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 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 4.2.1: 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 4.2.1: Groups

Provide the family or group name of each element.

a. F
b. Ca
c. Kr

Answer a: available at the end of this section.

Metals and Nonmetals

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 4.2.2. 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 4.2.2, 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.

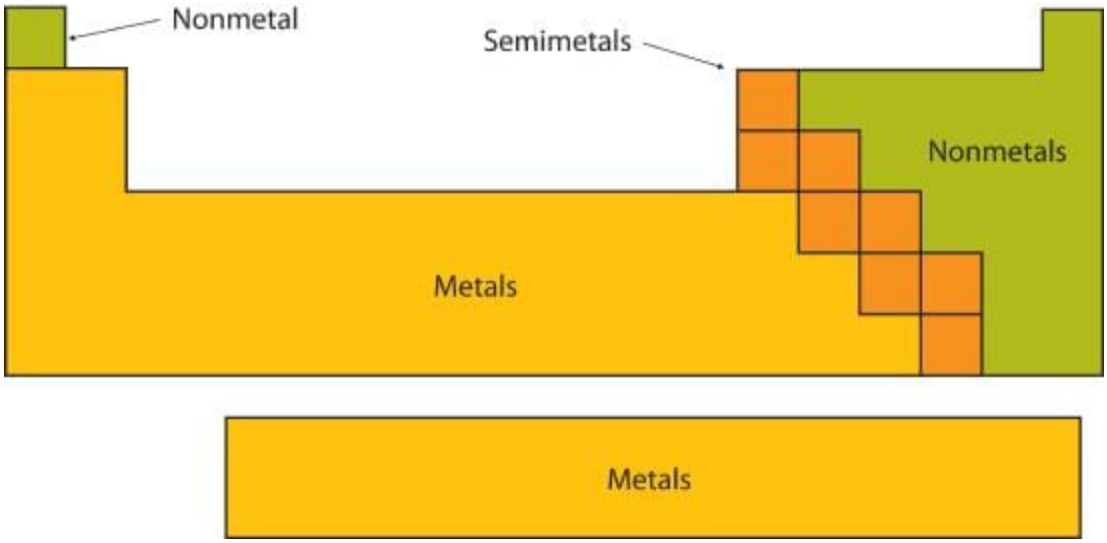


Figure 4.2.2: Types of Elements. Elements are either metals, nonmetals, or metalloids (or semi metals). Each group is located in a different part of the periodic table.

Example 4.2.2

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

a. Se
b. Mg
c. Ge

Solution

a. In Figure 4.2.1, selenium lies above and to the right of the diagonal line marking the boundary between metals and nonmetals, so it should be a nonmetal.
b. Magnesium lies to the left of the diagonal line marking the boundary between metals and nonmetals, so it should be a metal.
c. Germanium lies within the diagonal line marking the boundary between metals and nonmetals, so it should be a metalloid.

Exercise 4.2.2

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

Answer available at the end of this section.

Example 4.2.3: 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 4.2.3: Classification of Elements

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

a. F
b. U
c. Cu

Answer a: available at the end of this section.

Alkali Metals

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 4.2.2 *Alkali metals in water.*

Summary

- Valence electrons are the outer-shell electrons of an atom. These are electrons that can participate in the formation of a chemical bond.
- 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.
- An important grouping of elements in the periodic table are the main group elements, the transition metals, and the inner transition metals (the lanthanides, and the actinides).
- 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.

References

1. Petrucci, Ralph H., William S. Harwood, F. G. Herring, and Jeffrey D. Madura. General Chemistry: Principles and Modern Applications. 9th ed. Upper Saddle River: Pearson Education, Inc., 2007.
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.

Contributors and Attributions

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- Wikipedia

Answers

Exercise 4.2.1



Exercise 4.2.1**: Groups**

Fluorine is a halogen (Group 17)
Calcium is a alkaline earth metal (Group 2)
Krypton is a noble gas (Group 18)

Exercise 4.2.2

Indium is a metal.

Exercise 4.2.3**: Classification of Elements**

Fluorine is a nonmetal.
Uranium is a metal (and a inner transition metal too)
Copper is a metal (and a transition metal too)

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4.3: Lewis Structures of Ionic Compounds

Learning Objectives

- State the octet rule.
- Define cations and anions.
- Define *ionic bond*.
- Draw Lewis structures for ionic compounds.

The Octet Rule: The Drive for Eight

The **octet rule** refers to the tendency of atoms to prefer to have eight electrons in the *valence shell*. When atoms have fewer than eight electrons, they tend to react and form more stable compounds. When discussing the octet rule, we do not consider d or f electrons. Only the s and p electrons are involved in the octet rule, making it useful for the *main group elements* (elements not in the transition metal or inner-transition metal blocks); an octet in these atoms corresponds to an electron configurations ending with s^2p^6.

Note

The noble gases rarely form compounds. They have the most stable configuration (full octet, no charge), so they have no reason to react and change their configuration. All other elements attempt to gain, lose, or share electrons to achieve a noble gas configuration.

Most atoms do not have eight electrons in their valence electron shell. Some atoms have only a few electrons in their outer shell, while some atoms 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 as a result 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.

$$\mathrm{Na} \rightarrow \mathrm{Na^+} + \mathrm{e^-}$$

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. F

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.)

$$\mathrm{e^-} + \mathrm{Cl} \longrightarrow \mathrm{Cl^-}$$

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.

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

Note the usefulness of the periodic table in predicting likely ion formation and charge (Figure 4.3.1). Moving from the far left to the right on the periodic table, main-group elements tend to form cations with a charge equal to the group number. That is, group 1 elements form 1+ ions; group 2 elements form 2+ ions, and so on. Moving from the far right to the left on the periodic table, elements often form anions with a negative charge equal to the number of groups moved left from the noble gases. For example, group 17 elements (one group left of the noble gases) form 1– ions; group 16 elements (two groups left) form 2– ions, and so on. This trend can be used as a guide in many cases, but its predictive value decreases when moving toward the center of the periodic table. In fact, transition metals and some other metals often exhibit variable charges that are not predictable by their location in the table. For example, copper can form ions with a 1+ or 2+ charge, and iron can form ions with a 2+ or 3+ charge.

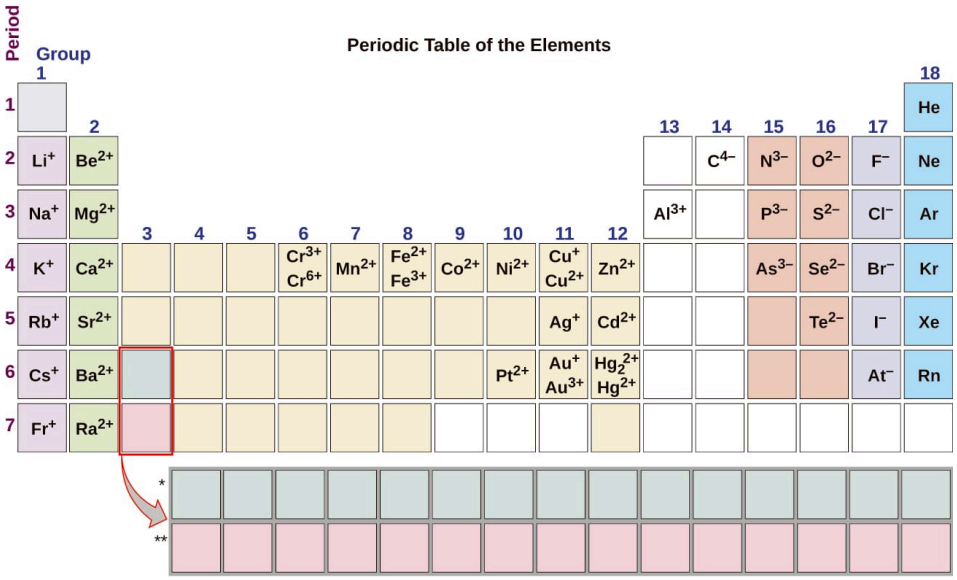


Figure 4.3.1: Some elements exhibit a regular pattern of ionic charge when they form ions.

Group one of the periodic table contains L i superscript plus sign in period 2, N a superscript plus sign in period 3, K superscript plus sign in period 4, R b superscript plus sign in period 5, C s superscript plus sign in period 6, and F r superscript plus sign in period 7. Group two contains B e superscript 2 plus sign in period 2, M g superscript 2 plus sign in period 3, C a superscript 2 plus sign in period 4, S r superscript 2 plus sign in period 5, B a superscript 2 plus sign in period 6, and R a superscript 2 plus sign in period 7. Group six contains C r superscript 3 plus sign and C r superscript 6 plus sign in period 4. Group seven contains M n superscript 2 plus sign in period 4. Group eight contains F e superscript 2 plus sign and F e superscript 3 plus sign in period 4. Group nine contains C o superscript 2 plus sign in period 4. Group ten contains N i superscript 2 plus sign in period 4, and P t superscript 2 plus sign in period 6. Group 11 contains C U superscript plus sign and C U superscript 2 plus sign in period 4, A g superscript plus sign in period 5, and A u superscript plus sign and A u superscript 3 plus sign in period 6. Group 12 contains Z n superscript 2 plus sign in period 4, C d superscript 2 plus sign in period 5, and H g subscript 2 superscript 2 plus sign and H g superscript 2 plus sign in period 6. Group 13 contains A l superscript 3 plus sign in period 3. Group 14 contains C superscript 4 negative sign in period 2. Group 15 contains N superscript 3 negative sign in period 2, P superscript 3 negative sign in period 3, and A s superscript 3 negative sign in period 4. Group 16 contains O superscript 2 negative sign in period 2, S superscript 2 negative sign in period 3, S e superscript 2 negative sign in period 4 and T e superscript 2 negative sign in period 5. Group 17 contains F superscript negative sign in period 2, C l superscript negative sign in period 3, B r superscript negative sign in period 4, I superscript negative sign in period 5, and A t superscript negative sign in period 6. Group 18 contains H e in period 1, N e in period 2, A r in period 3, K r in period 4, X e in period 5 and R n in period 6.

Lewis Symbols for Ionic Compounds

In the section above, we saw how ions are formed by losing electrons to make cations or by gaining electrons to form anions. The astute reader may have noticed something: Many of the ions that form have eight electrons in their valence shell. Either atoms gain enough electrons to have eight electrons in the valence shell and become the appropriately charged anion, or they lose the electrons in their original valence shell; the *lower* shell, now the valence shell, has eight electrons in it, so the atom becomes positively charged. For whatever reason, having eight electrons in a valence shell is a particularly energetically stable arrangement of electrons. When atoms form compounds, the octet rule is not always satisfied for all atoms at all times, but it is a very good

rule of thumb for understanding the kinds of bonding arrangements that atoms can make.

It is not impossible to violate the octet rule. Consider lithium: in its elemental form, it has one valence electron and is stable. It is rather reactive, however, and does not require a lot of energy to remove that electron to make the Li^+ ion. We *could* remove another electron by adding even more energy to the ion, to make the Li^{2+} ion. However, that requires much more energy than is normally available in chemical reactions, so sodium stops at a 1+ charge after losing a single electron. It turns out that the Li^+ ion has a complete octet in its new valence shell, the $n = 2$ shell, which satisfies the octet rule. The octet rule is a result of trends in energies and is useful in explaining why atoms form the ions that they do.

Now consider an Li atom in the presence of a Cl atom. The two atoms have these Lewis electron dot diagrams and electron configurations:

$$\mathbf{Li}\cdot \qquad \cdot\ddot{\mathbf{Cl}}:$$

$$[Ne]\,3s^1 \qquad [Ne]\,3s^2 3p^5$$

For the Li atom to obtain an octet, it must lose an electron; for the Cl atom to gain an octet, it must gain an electron. An electron transfers from the Na atom to the Cl atom:

$$\mathbf{Li}\cdot \rightsquigarrow \cdot\ddot{\mathbf{Cl}}:$$

resulting in two ions—the Na^+ ion and the Cl^- ion:

$$\mathbf{Li}\cdot^{+} \qquad :\ddot{\mathbf{Cl}}:^{-}$$

$$[Ne] \qquad [Ne]\,3s^2 3p^6$$

Both species now have complete octets, and the electron shells are energetically stable. From basic physics, we know that opposite charges attract. This is what happens to the Na^+ and Cl^- ions:

$$\mathbf{Li}\cdot^{+} + :\ddot{\mathbf{Cl}}:^{-} \rightarrow Na^+Cl^- \;\; or \;\; NaCl$$

where we have written the final formula (the formula for sodium chloride) as per the convention for ionic compounds, without listing the charges explicitly. The attraction between oppositely charged ions is called an **ionic bond,** and it is one of the main types of chemical bonds in chemistry. Ionic bonds are caused by electrons *transferring* from one atom to another.

In electron transfer, the number of electrons lost must equal the number of electrons gained. We saw this in the formation of NaCl. A similar process occurs between Mg atoms and O atoms, except in this case two electrons are transferred:



The two ions each have octets as their valence shell, and the two oppositely charged particles attract, making an ionic bond:

$$\mathbf{Mg}^{2+} + \left[:\ddot{\mathbf{O}}:\right]^{2-} \qquad Mg^{2+}O^{2-} \; or \; MgO$$

Remember, in the final formula for the ionic compound, we do not write the charges on the ions.

What about when an Na atom interacts with an O atom? The O atom needs two electrons to complete its valence octet, but the Na atom supplies only one electron:

$$\mathbf{Na}\cdot \rightsquigarrow \cdot\ddot{\mathbf{O}}:$$

The O atom still does not have an octet of electrons. What we need is a second Na atom to donate a second electron to the O atom:



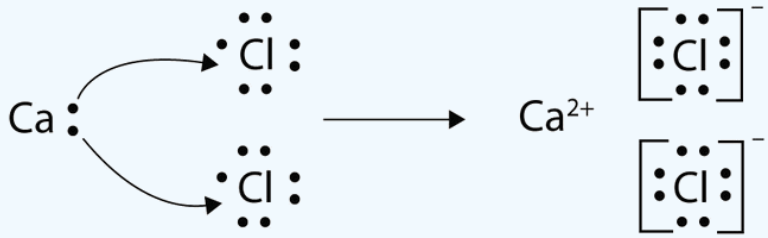
These three ions attract each other to give an overall neutral-charged ionic compound, which we write as Na_2O. The need for the number of electrons lost being equal to the number of electrons gained explains why ionic compounds have the ratio of cations to anions that they do. This is required by the law of conservation of matter as well.

Example 4.3.1: Synthesis of calcium chloride from Elements

With arrows, illustrate the transfer of electrons to form calcium chloride from Ca atoms and Cl atoms.

Solution

A Ca atom has two valence electrons, while a Cl atom has seven electrons. A Cl atom needs only one more to complete its octet, while Ca atoms have two electrons to lose. Thus we need two Cl atoms to accept the two electrons from one Ca atom. The transfer process looks as follows:

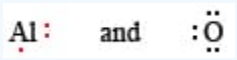


The oppositely charged ions attract each other to make $CaCl_2$.

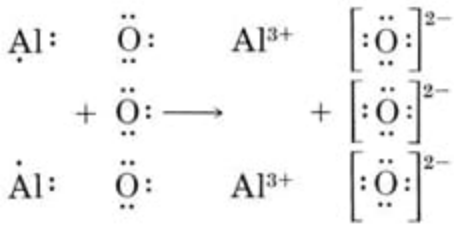
Example 4.3.2: Ionic Formula

Find the formula of the ionic compound formed from O and Al.

Solution: We first write down Lewis diagrams for each atom involved:



We now see that each O atom needs 2 electrons to make up an octet, while each Al atom has 3 electrons to donate. In order that the same number of electrons would be donated as accepted, we need 2 Al atoms ($2 \times 3e^-$ donated) and 3 O atoms ($3 \times 2e^-$ accepted). The whole process is then



The resultant oxide consists of aluminum ions, Al^{3+}, and oxide ions, O^{2-}, in the ratio of 2:3. The formula is Al_2O_3.

Exercise 4.3.1

With arrows, illustrate the transfer of electrons to form potassium sulfide from K atoms and S atoms.

Answer: available at the end of this section.

Summary

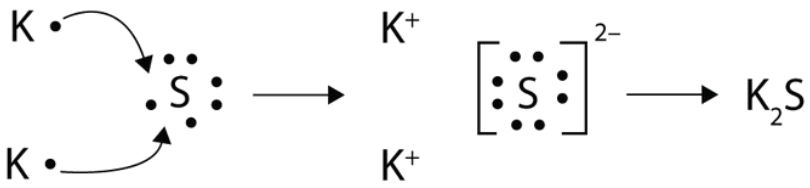
- The tendency to form species that have eight electrons in the valence shell is called the octet rule.
- The attraction of oppositely charged ions caused by electron transfer is called an ionic bond.
- The strength of ionic bonding depends on the magnitude of the charges and the sizes of the ions.

Contributors and Attributions

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- OpenSTAX

Answers

Exercise 4.3.1



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4.4: Lewis Structures of Covalent Compounds and Ions

Learning Objective

- Write Lewis structures for molecules and ions.
- Know the exceptions to the octet rule.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions:

How-to: Constructing Lewis electron structures

1. Determine the total number of valence electrons in the molecule or ion.

- Add together the valence electrons from each atom. (Recall that the number of valence electrons is indicated by the position of the element in the periodic table.)
- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

  For CO_3^{2-}, for example, we add two electrons to the total because of the –2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and each atom has an octet of electrons except for hydrogen (with two electrons).

The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 4.4.1: Water

Write the Lewis Structure for H_2O.

Solution

Example 4.4.1: steps for writing the Lewis structure of water

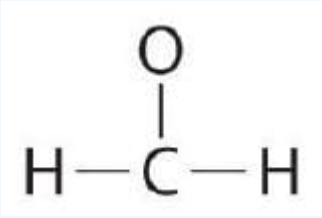
Steps for Writing Lewis Structures	Example 4.4.1:
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.
2. Arrange the atoms to show specific connections.	**H O H** Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives **H -O- H** with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary
7. Final check	The Lewis structure gives oxygen an octet and each hydrogen two electrons,

Example 4.4.2

Write the Lewis structure for the CH_2O molecule

Solution

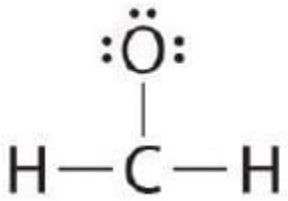
Steps are writing the Lewis structure of the CH_2O molecule

Steps for Writing Lewis Structures	Example 4.4.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has one valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following: 

Six electrons are used, and 6 are left over.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:



Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.

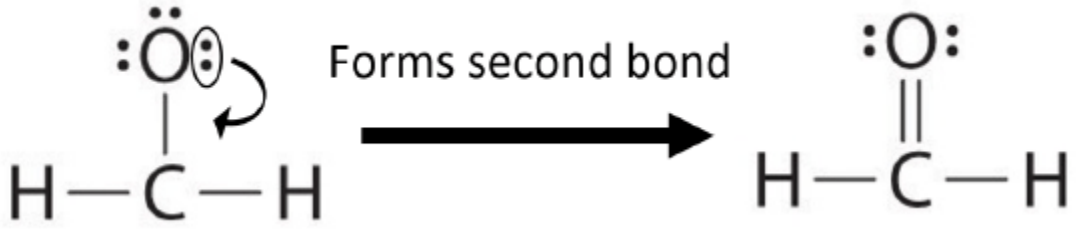
5. If any electrons are left over, place them on the central atom.

Not necessary

There are no electrons left to place on the central atom.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond:



7. Final check Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 4.4.1

Write Lewis electron structures for CO_2 and SCl_2, a vile-smelling, unstable red liquid that is used in the manufacture of rubber.

Answer CO2 available at the end of this section.

The United States Supreme Court has the unenviable task of deciding what the law is. This responsibility can be a major challenge when there is no clear principle involved or where there is a new situation not encountered before. Chemistry faces the same challenge in extending basic concepts to fit a new situation. Drawing of Lewis structures for polyatomic ions uses the same approach, but tweaks the process a little to fit a somewhat different set of circumstances.

Lewis Structures for Polyatomic Ions

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, NH_4^+, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia (NH_3) molecule in a coordinate covalent bond.

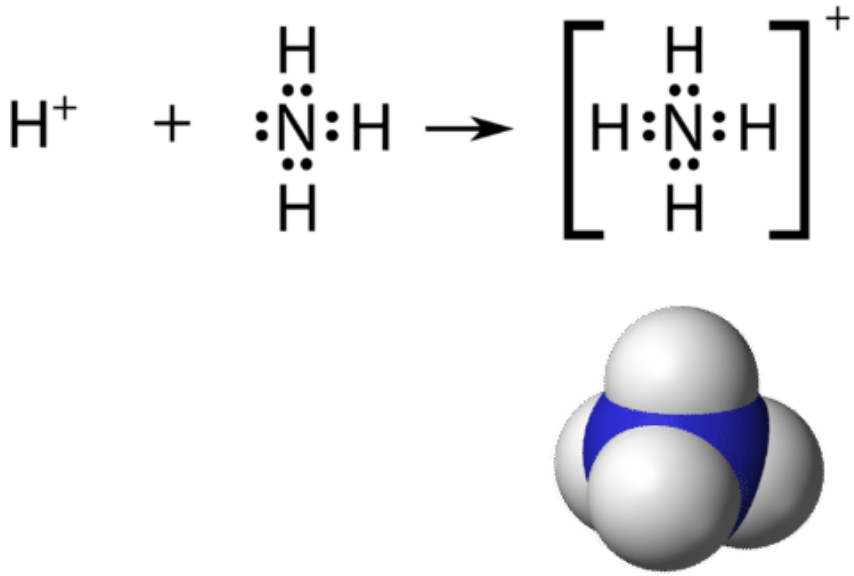


Figure 4.4.3: The ammonium ion.

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$1\ \mathrm{N}$ atom $= 5$ valence electrons

$4\ \mathrm{H}$ atoms $= 4 \times 1 = 4$ valence electrons

subtract 1 electron for the $1+$charge of the ion

total of 8 valence electrons in the ion

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside the brackets.

Exercise 4.4.2

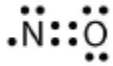
Draw the Lewis electron dot structure for the sulfate ion.

Answer available at the end of this section.

Exceptions to the Octet Rule

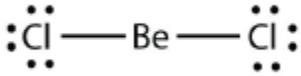
As important and useful as the octet rule is in chemical bonding, there are some well-known violations. This does not mean that the octet rule is useless—quite the contrary. As with many rules, there are exceptions, or violations.

There are three violations to the octet rule. Odd-electron molecules represent the first violation to the octet rule. Although they are few, some stable compounds have an odd number of electrons in their valence shells. With an odd number of electrons, at least one atom in the molecule will have to violate the octet rule. Examples of stable odd-electron molecules are NO, NO_2, and ClO_2. The Lewis electron dot diagram for NO is as follows:

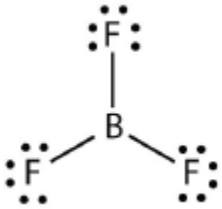


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:

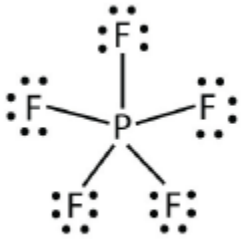


Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell.

These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

Example 4.4.3: Octet Violations

Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

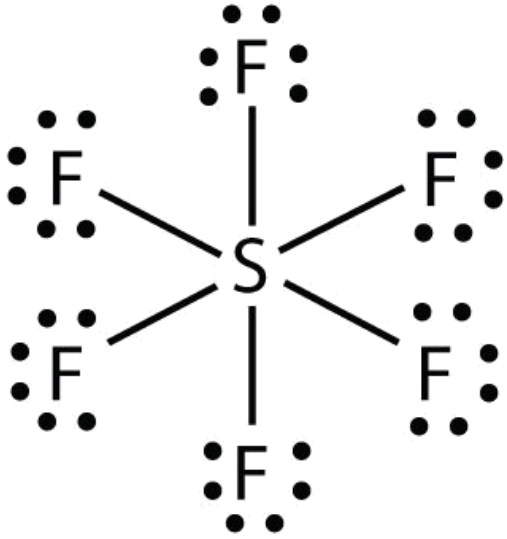
a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:



b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



Exercise 4.4.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer: available at the end of this section.

The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

- Lewis dot symbols (Lewis structures) provide a simple rationalization of why elements form compounds with the observed stoichiometries.
- In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms.
- Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds.
- There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules

Contributors and Attributions

- Modified by Joshua Halpern (Howard University)
-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Learning Objective

- Write Lewis structures for molecules and ions.
- Know the exceptions to the octet rule.

The following procedure can be used to construct Lewis electron structures for more complex molecules and ions:

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- If the species is a polyatomic ion, remember to add or subtract the number of electrons necessary to give the total charge on the ion.

  For CO_3^{2-}, for example, we add two electrons to the total because of the –2 charge.

2. Arrange the atoms to show specific connections.

- When there is a central atom, it is usually the least electronegative element in the compound. Chemists usually list this central atom first in the chemical formula (as in CCl_4 and CO_3^{2-}, which both have C as the central atom), which is another clue to the compound's structure.
- Hydrogen and the halogens are almost always connected to only one other atom, so they are usually *terminal* rather than central.

3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.

- In H_2O, for example, there is a bonding pair of electrons between oxygen and each hydrogen.

4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).

- These electrons will usually be lone pairs.

5. If any electrons are left over, place them on the central atom.

- We will explain later that some atoms are able to accommodate more than eight electrons.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

- This will not change the number of electrons on the terminal atoms.

7. Final check

- Always make sure all valence electrons are accounted for and each atom has an octet of electrons except for hydrogen (with two electrons).

The central atom is usually the least electronegative element in the molecule or ion; hydrogen and the halogens are usually terminal.

Now let's apply this procedure to some particular compounds, beginning with one we have already discussed.

Example 4.4.1: Water

Write the Lewis Structure for H_2O.

Solution

Example 4.4.1: steps for writing the Lewis structure of water

Steps for Writing Lewis Structures	Example 4.4.1:
1. Determine the total number of valence electrons in the molecule or ion.	Each H atom (group 1) has 1 valence electron, and the O atom (group 16) has 6 valence electrons, for a total of 8 valence electrons.
2. Arrange the atoms to show specific connections.	**H O H** Because H atoms are almost always terminal, the arrangement within the molecule must be HOH.

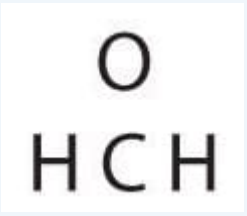
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3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond. 4. Beginning with the terminal atoms, add enough electrons to each atom to give each atom an octet (two for hydrogen).	Placing one bonding pair of electrons between the O atom and each H atom gives **H -O- H** with 4 electrons left over. Each H atom has a full valence shell of 2 electrons.
5. If any electrons are left over, place them on the central atom.	Adding the remaining 4 electrons to the oxygen (as two lone pairs) gives the following structure: 
6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.	Not necessary
7. Final check	The Lewis structure gives oxygen an octet and each hydrogen two electrons,

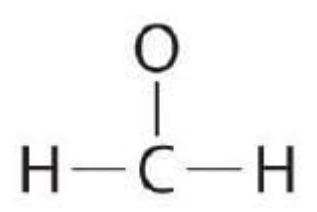
Example 4.4.2

Write the Lewis structure for the CH_2O molecule

Solution

Steps are writing the Lewis structure of the CH_2O molecule

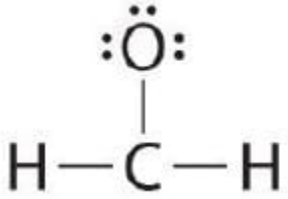
Steps for Writing Lewis Structures	Example 4.4.2
1. Determine the total number of valence electrons in the molecule or ion.	Each hydrogen atom (group 1) has one valence electron, carbon (group 14) has 4 valence electrons, and oxygen (group 16) has 6 valence electrons, for a total of [(2)(1) + 4 + 6] = 12 valence electrons.
2. Arrange the atoms to show specific connections.	 Because carbon is less electronegative than oxygen and hydrogen is normally terminal, C must be the central atom.

Steps for Writing Lewis Structures	Example 4.4.2
3. Place a bonding pair of electrons between each pair of adjacent atoms to give a single bond.	Placing a bonding pair of electrons between each pair of bonded atoms gives the following: 

Six electrons are used, and 6 are left over.

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Adding all 6 remaining electrons to oxygen (as three lone pairs) gives the following:



Although oxygen now has an octet and each hydrogen has 2 electrons, carbon has only 6 electrons.

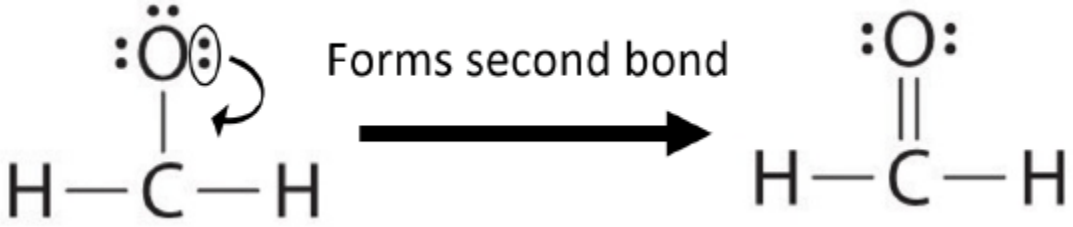
5. If any electrons are left over, place them on the central atom.

Not necessary

There are no electrons left to place on the central atom.

6. If the central atom has fewer electrons than an octet, use lone pairs from terminal atoms to form multiple (double or triple) bonds to the central atom to achieve an octet.

To give carbon an octet of electrons, we use one of the lone pairs of electrons on oxygen to form a carbon–oxygen double bond:



7. Final check Both the oxygen and the carbon now have an octet of electrons, so this is an acceptable Lewis electron structure. The O has two bonding pairs and two lone pairs, and C has four bonding pairs. This is the structure of formaldehyde, which is used in embalming fluid.

Exercise 4.4.1

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Lewis Structures for Polyatomic Ions

Recall that a polyatomic ion is a group of atoms that are covalently bonded together and which carry an overall electrical charge. The ammonium ion, $\mathrm{NH_4}^+$, is formed when a hydrogen ion (H^+) attaches to the lone pair of an ammonia $(\mathrm{NH_3})$ molecule in a coordinate covalent bond.

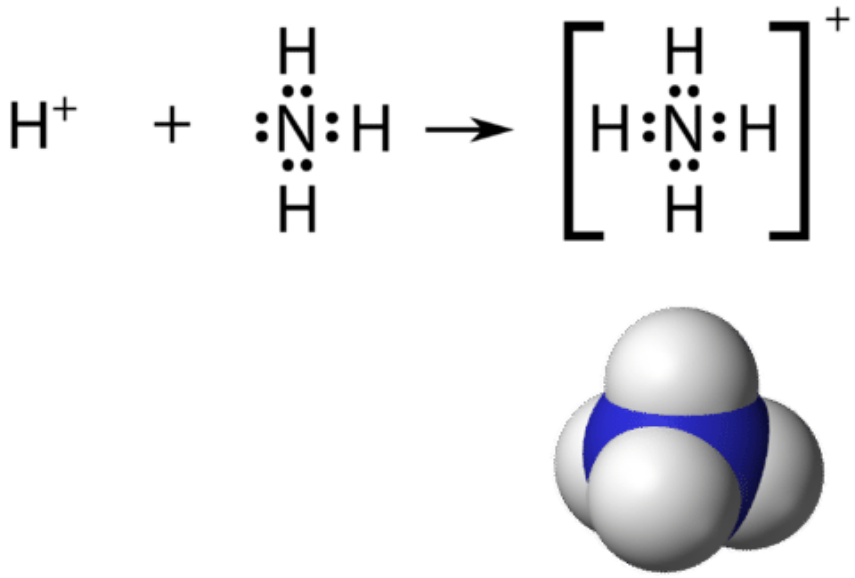


Figure 4.4.3: The ammonium ion.

When drawing the Lewis structure of a polyatomic ion, the charge of the ion is reflected in the number of total valence electrons in the structure. In the case of the ammonium ion:

$1\ \mathrm{N}$ atom $= 5$ valence electrons

$4\ \mathrm{H}$ atoms $= 4 \times 1 = 4$ valence electrons

subtract 1 electron for the $1+$charge of the ion

total of 8 valence electrons in the ion

It is customary to put the Lewis structure of a polyatomic ion into a large set of brackets, with the charge of the ion as a superscript outside the brackets.

Exercise 4.4.2

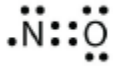
Draw the Lewis electron dot structure for the sulfate ion.

Answer available at the end of this section.

Exceptions to the Octet Rule

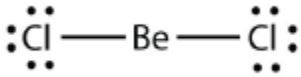
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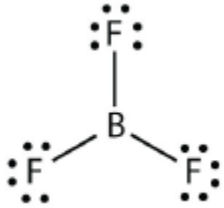


Although the O atom has an octet of electrons, the N atom has only seven electrons in its valence shell. Although NO is a stable compound, it is very chemically reactive, as are most other odd-electron compounds.

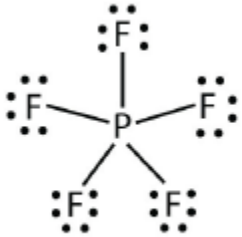
Electron-deficient molecules represent the second violation to the octet rule. These stable compounds have less than eight electrons around an atom in the molecule. The most common examples are the covalent compounds of beryllium and boron. For example, beryllium can form two covalent bonds, resulting in only four electrons in its valence shell:



Boron commonly makes only three covalent bonds, resulting in only six valence electrons around the B atom. A well-known example is BF_3:



The third violation to the octet rule is found in those compounds with more than eight electrons assigned to their valence shell. These are called expanded valence shell molecules. Such compounds are formed only by central atoms in the third row of the periodic table or beyond that have empty *d* orbitals in their valence shells that can participate in covalent bonding. One such compound is PF_5. The only reasonable Lewis electron dot diagram for this compound has the P atom making five covalent bonds:



Formally, the P atom has 10 electrons in its valence shell.

Example 4.4.3: Octet Violations

Identify each violation to the octet rule by drawing a Lewis electron dot diagram.

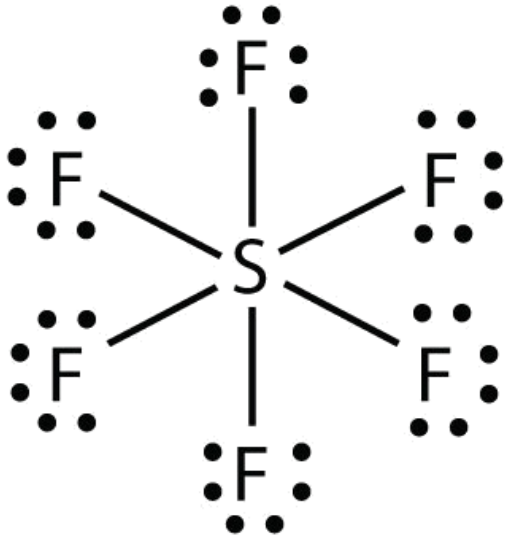
a. ClO
b. SF_6

Solution

a. With one Cl atom and one O atom, this molecule has 6 + 7 = 13 valence electrons, so it is an odd-electron molecule. A Lewis electron dot diagram for this molecule is as follows:



b. In SF_6, the central S atom makes six covalent bonds to the six surrounding F atoms, so it is an expanded valence shell molecule. Its Lewis electron dot diagram is as follows:



Exercise 4.4.3: Xenon Difluoride

Identify the violation to the octet rule in XeF_2 by drawing a Lewis electron dot diagram.

Answer: available at the end of this section.

The Xe atom has an expanded valence shell with more than eight electrons around it.

Summary

- Lewis dot symbols (Lewis structures) provide a simple rationalization of why elements form compounds with the observed stoichiometries.

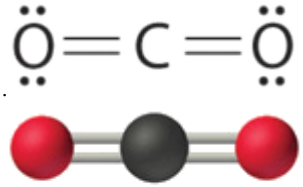
- In Lewis electron structures, we encounter **bonding pairs**, which are shared by two atoms, and **lone pairs**, which are not shared between atoms.
- Lewis structures for polyatomic ions follow the same rules as those for other covalent compounds.
- There are three violations to the octet rule: odd-electron molecules, electron-deficient molecules, and expanded valence shell molecules

Contributors and Attributions

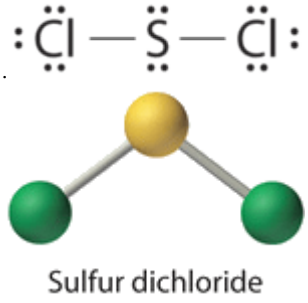
- Modified by Joshua Halpern (Howard University)
-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Answers

Exercise 4.4.1

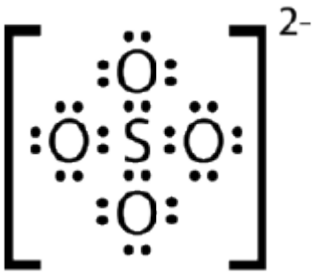


Carbon dioxide



Sulfur dichloride

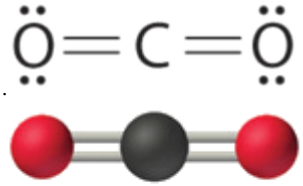
Exercise 4.4.2



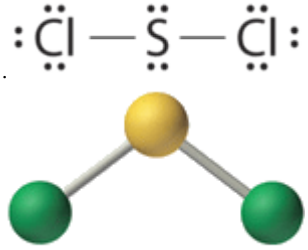
Exercise 4.4.3**: Xenon Difluoride**



Exercise 4.4.1

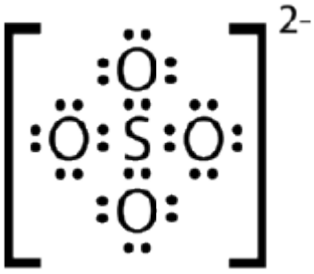


Carbon dioxide



Sulfur dichloride

Exercise 4.4.2



Exercise 4.4.3**: Xenon Difluoride**



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4.5: VSEPR Theory

Learning Objective

- Determine the shape of simple molecules and polyatomic ions.

Molecules have shapes. There is an abundance of experimental evidence to that effect—from their physical properties to their chemical reactivity. Small molecules—molecules with a single central atom—have shapes that can be easily predicted. The basic idea in molecular shapes is **called valence shell electron pair repulsion (VSEPR).** It basically says that electron pairs, being composed of negatively charged particles, repel each other to get as far away from each other as possible. VSEPR makes a distinction between *electron group geometry*, which expresses how electron groups (bonds and nonbonding electron pairs) are arranged, and *molecular geometry*, which expresses how the atoms in a molecule are arranged. However, the two geometries are related.

There are two types of **electron groups**: any type of bond—single, double, or triple—and lone electron pairs. When applying VSEPR to simple molecules, the first thing to do is to count the number of electron groups around the central atom. Remember that a multiple bond counts as only *one* electron group.

Any molecule with only two atoms is linear. A molecule whose central atom contains only two electron groups orients those two groups as far apart from each other as possible—180° apart. When the two electron groups are 180° apart, the atoms attached to those electron groups are also 180° apart, so the overall molecular shape is linear. Examples include BeH_2 and CO_2:



Figure 4.5.1 *Beryllium hydride and carbon dioxide bonding.*

The two molecules, shown in the figure below in a "ball and stick" model.

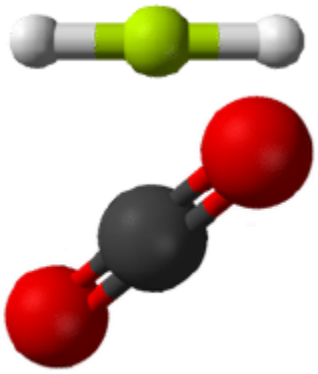


Figure 4.5.2 *Beryllium hydride and carbon dioxide models.*

A molecule with three electron groups orients the three groups as far apart as possible. They adopt the positions of an equilateral triangle—120° apart and in a plane. The shape of such molecules is *trigonal planar*. An example is BF_3:

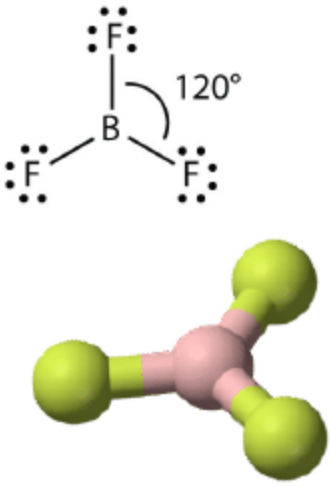


Figure 4.5.3 *Boron trifluoride bonding.*

Some substances have a trigonal planar electron group distribution but have atoms bonded to only two of the three electron groups. An example is GeF_2:

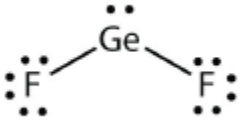


Figure 4.5.4: Germanium *difluoride bonding.*

From an electron group geometry perspective, GeF_2 has a trigonal planar shape, but its real shape is dictated by the positions of the atoms. This shape is called *bent* or *angular*.

A molecule with four electron groups about the central atom orients the four groups in the direction of a tetrahedron, as shown in Figure 4.5.1 Tetrahedral Geometry. If there are four atoms attached to these electron groups, then the molecular shape is also *tetrahedral*. Methane (CH_4) is an example.

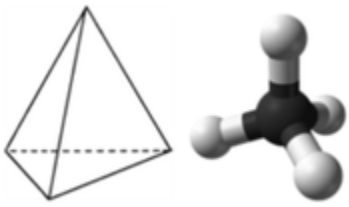


Figure 4.5.5 *Tetrahedral structure of methane.*

This diagram of CH_4 illustrates the standard convention of displaying a three-dimensional molecule on a two-dimensional surface. The straight lines are in the plane of the page, the solid wedged line is coming out of the plane toward the reader, and the dashed wedged line is going out of the plane away from the reader.

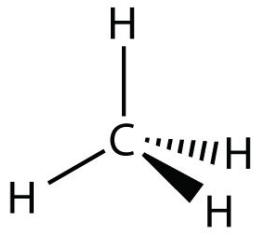


Figure 4.5.6 *Methane bonding.*

NH_3 is an example of a molecule whose central atom has four electron groups but only three of them are bonded to surrounding atoms.

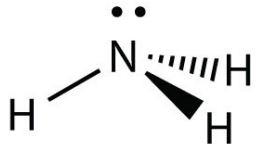


Figure 4.5.7 *Ammonia bonding.*

Although the electron groups are oriented in the shape of a tetrahedron, from a molecular geometry perspective, the shape of NH_3 is *trigonal pyramidal*.

H_2O is an example of a molecule whose central atom has four electron groups but only two of them are bonded to surrounding atoms.



Figure 4.5.8 *Water bonding.*

Although the electron groups are oriented in the shape of a tetrahedron, the shape of the molecule is *bent* or *angular*. A molecule with four electron groups about the central atom but only one electron group bonded to another atom is linear because there are only two atoms in the molecule.

Double or triple bonds count as a single electron group. The Lewis electron dot diagram of formaldehyde (CH_2O) is shown in Figure 4.5.9.



Figure 4.5.9 *Lewis electron dot diagram of formaldehyde.*

The central C atom has three electron groups around it because the double bond counts as one electron group. The three electron groups repel each other to adopt a trigonal planar shape.

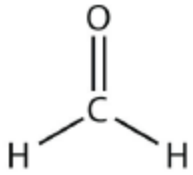


Figure 4.5.10: Formaldehyde bonding.

(The lone electron pairs on the O atom are omitted for clarity.) The molecule will not be a perfect equilateral triangle because the C–O double bond is different from the two C–H bonds, but both planar and triangular describe the appropriate approximate shape of this molecule. Figure 4.5.11 illustrates several representations of the water, ammonia, and methane molecules.

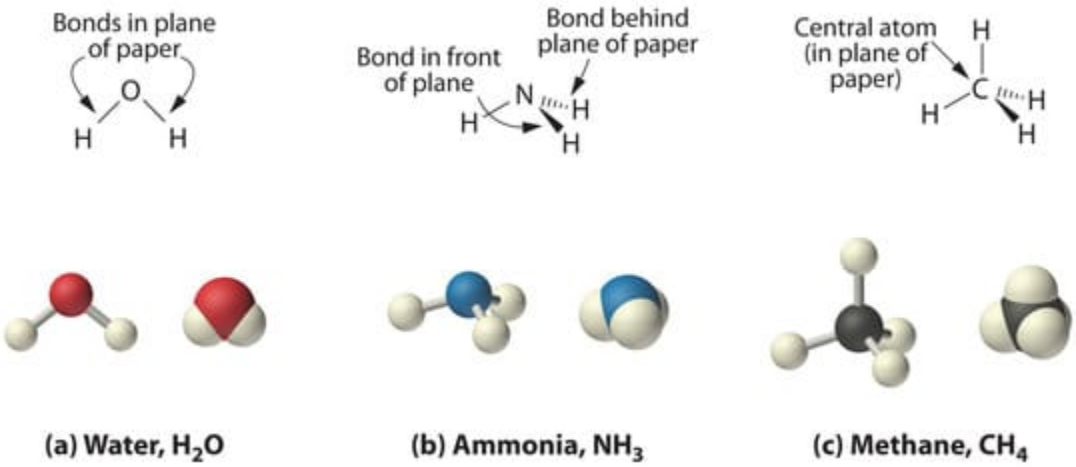


Figure 4.5.11 The three-dimensional structures, ball and stick models,and space filling models of water, ammonia, and methane. *(a) Water is a V-shaped molecule, in which all three atoms lie in a plane. (b) In contrast, ammonia has a pyramidal structure, in which the three hydrogen atoms form the base of the pyramid and the nitrogen atom is at the vertex. (c) The four hydrogen atoms of methane form a tetrahedron; the carbon atom lies in the center.*

Table 4.5.1 summarizes the shapes of molecules based on their number of electron groups and surrounding atoms and the common shapes are shown in figure 4.5.1.

Table 4.5.1: Summary of Molecular Shapes

Number of Electron Groups on Central Atom	Number of Bonding Groups	Number of Lone Pairs	Electron Geometry	Molecular Shape
2	2	0	linear	linear
3	3	0	trigonal planar	trigonal planar
3	2	1	trigonal planar	bent
4	4	0	tetrahedral	tetrahedral
4	3	1	tetrahedral	trigonal pyramidal
4	2	2	tetrahedral	bent

Figure 4.5.12: Common structures for molecules and Ppolyatomic ions that consist of a central atom bonded to two or three other atoms.

Example 4.5.1:

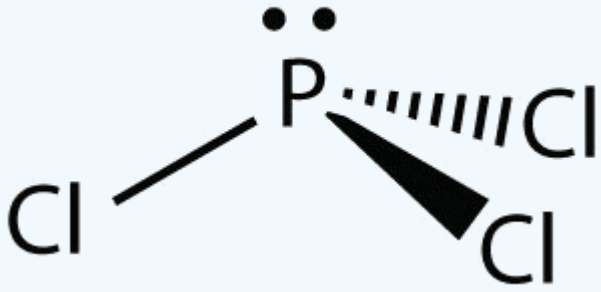
What is the approximate shape of each molecule?

a. PCl_3
b. NOF

Solution

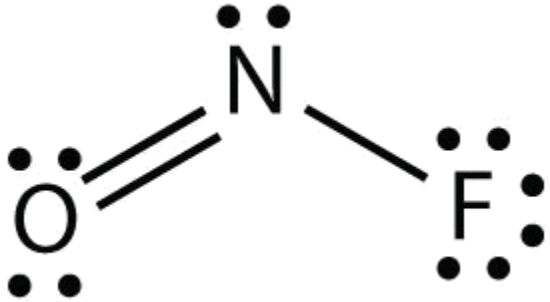
The first step is to draw the Lewis structure of the molecule.

1. For PCl_3, the electron dot diagram is as follows:



The lone electron pairs on the Cl atoms are omitted for clarity. The P atom has four electron groups with three of them bonded to surrounding atoms, so the molecular shape is trigonal pyramidal.

The electron dot diagram for NOF is as follows:



The N atom has three electron groups on it, two of which are bonded to other atoms. The molecular shape is bent.

Exercise 4.5.1

What is the approximate molecular shape of CH_2Cl_2?

Answer available at the end of this section.

Exercise 4.5.2

Ethylene (C_2H_4) has two central atoms. Determine the geometry around each central atom and the shape of the overall molecule. Hint, hydrogen is a terminal atom.

Answer available at the end of this section.

Summary

The approximate shape of a molecule can be predicted from the number of electron groups and the number of surrounding atoms.

Contributors and Attributions

-
- TextMap: Beginning Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Answers

Exercise 4.5.1

Tetrahedral

Exercise 4.5.2

Trigonal planar about both central C atoms

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4.6: Polarity

Learning Objectives

- Define electronegativity
- Determine the polarity of a covalent bond.

Electronegativity

The ability of an atom in a molecule to attract shared electrons is called **electronegativity**. When two atoms combine, the difference between their electronegativities is an indication of the type of bond that will form. If the difference between the electronegativities of the two atoms is small, neither atom can take the shared electrons completely away from the other atom and the bond will be covalent. If the difference between the electronegativities is large, the more electronegative atom will take the bonding electrons completely away from the other atom (electron transfer will occur) and the bond will be ionic. This is why metals (low electronegativities) bonded with nonmetals (high electronegativities) typically produce ionic compounds.

A bond may be so polar that an electron actually transfers from one atom to another, forming a true ionic bond. How do we judge the degree of polarity? Scientists have devised a scale called **electronegativity**, a scale for judging how much atoms of any element attract electrons. Electronegativity is a unitless number; the higher the number, the more an atom attracts electrons. A common scale for electronegativity is shown in Figure 4.6.1.

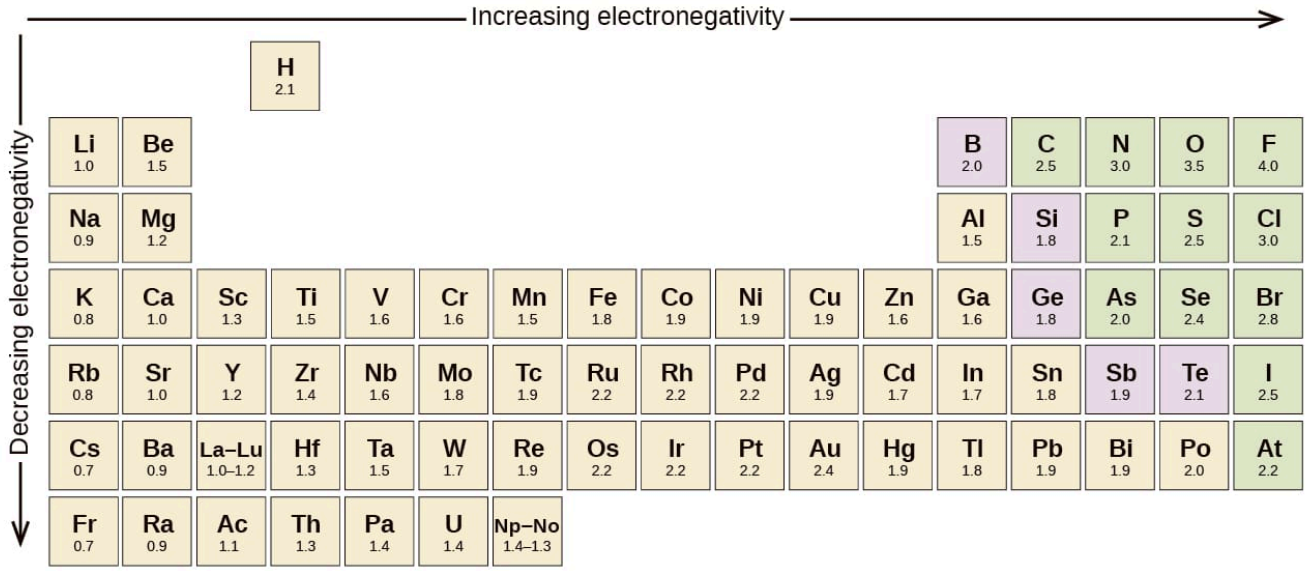


Figure 4.6.1 Electronegativities of the Elements. Electronegativities are used to determine the polarity of covalent bonds.

Bond Polarity

The polarity of a covalent bond can be judged by determining the *difference* of the electronegativities of the two atoms involved in the covalent bond, as summarized in the following table:

Table 4.6.1:Electronegativity differences and bond type

Electronegativity Difference	Bond Type
0	nonpolar covalent
0–0.4	slightly polar covalent

Electronegativity Difference	Bond Type
0.5–2.0	definitely polar covalent
>2.0	likely ionic

Nonpolar Covalent Bonds

A bond in which the electronegativity difference is less than 1.9 is considered to be mostly covalent in character. However, at this point we need to distinguish between two general types of covalent bonds. A **nonpolar covalent bond** is a covalent bond in which the bonding electrons are shared equally between the two atoms. In a nonpolar covalent bond, the distribution of electrical charge is balanced between the two atoms.

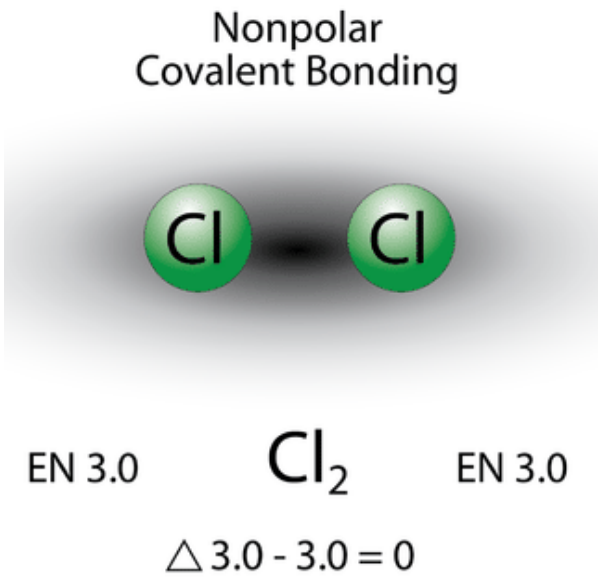


Figure 4.6.2 A nonpolar covalent bond is one in which the distribution of electron density between the two atoms is equal.

The two chlorine atoms share the pair of electrons in the single covalent bond equally, and the electron density surrounding the $\mathrm{Cl_2}$ molecule is symmetrical. Also note that molecules in which the electronegativity difference is very small (<0.5) are also considered nonpolar covalent. An example would be a bond between chlorine and bromine ($\Delta \mathrm{EN} = 3.0 - 2.8 = 0.2$).

Polar Covalent Bonds

A bond in which the electronegativity difference between the atoms is between 0.5 and 2.0 is called a polar covalent bond. A **polar covalent bond** is a covalent bond in which the atoms have an unequal attraction for electrons and so the sharing is unequal. In a polar covalent bond, sometimes simply called a polar bond, the distribution of electrons around the molecule is no longer symmetrical.

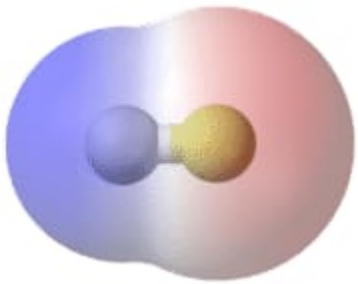


Figure 4.6.3 In the polar covalent bond of HF, the electron density is unevenly distributed. There is a higher density (red) near the fluorine atom, and a lower density (blue) near the hydrogen atom.

An easy way to illustrate the uneven electron distribution in a polar covalent bond is to use the Greek letter delta (δ).

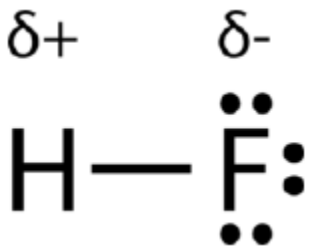


Figure $4.6.4$ Use of δ to indicate partial charge.

The atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge. The delta symbol is used to indicate that the quantity of charge is less than one. A crossed arrow can also be used to indicate the direction of greater electron density.

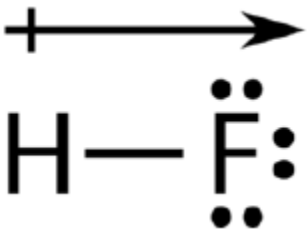


Figure $4.6.5$ Use of crossed arrow to indicate polarity

YouTube Video

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View this YouTube Video online

Electronegativity differences in bonding using Pauling scale. Using differences in electronegativity to classify bonds as covalent, polar covalent, or ionic.

Example $4.6.1$: Bond Polarity

What is the polarity of each bond?

a. C–H
b. O–H

Solution

Using Figure $4.6.1$, we can calculate the difference of the electronegativities of the atoms involved in the bond.

a. For the C–H bond, the difference in the electronegativities is 2.5 – 2.1 = 0.4. Thus we predict that this bond will be non polar covalent.
b. For the O–H bond, the difference in electronegativities is 3.5 – 2.1 = 1.4, so we predict that this bond will be definitely polar covalent.

Exercise 4.6.1

What is the polarity of each bond?

a. Rb–F
b. P–Cl

Answer a available at the end of this section.

Summary

- The type of bond (polar covalent,non polar covalent or ionic) between two atoms is determined by the differences in electronegativity.
- For atoms sharing a polar covalent bond, the atom with the greater electronegativity acquires a partial negative charge, while the atom with the lesser electronegativity acquires a partial positive charge.

Contributors and Attributions

-

- StackExchange (thomij).
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Learning Objective

- Determine if a molecule is polar or nonpolar.

Molecular Polarity

To determine if a molecule is polar or nonpolar, it is frequently useful to look at Lewis structures. **Nonpolar compounds** will be symmetric, meaning all of the sides around the central atom are identical - bonded to the same element with no unshared pairs of electrons. Notice that a tetrahedral molecule such as CCl_4 is nonpolar Figure $(4.6.1$. Another non polar molecule shown below is boron trifluoride, BF_3. BF_3 is a trigonal planar molecule and all three peripheral atoms are the same.

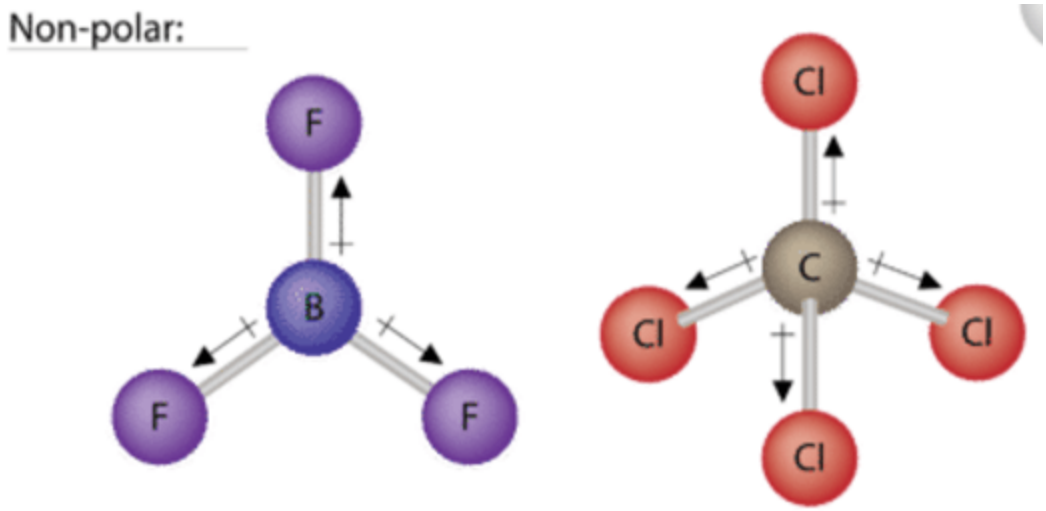
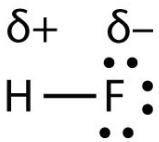


Figure 4.6.1 Some examples of nonpolar molecules based on molecular geometry (BF_3 and CCl_4).

Polar molecules are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded. This works pretty well - as long as you can visualize the molecular geometry. That's the hard part. To know how the bonds are oriented in space, you have to have a strong grasp of Lewis structures and VSEPR theory. Assuming you do, you can look at the structure of each one and decide if it is polar or not - *whether or not you know the individual atom electronegativity*. This is because you know that all bonds between dissimilar elements are polar, and in these particular examples, it doesn't matter which direction the dipole moment vectors are pointing (out or in).

A **polar molecule** is a molecule in which one end of the molecule is slightly positive, while the other end is slightly negative. A diatomic molecule that consists of a polar covalent bond, such as HF, is a polar molecule.



As mentioned in section 4.7, because the electrons in the bond are nearer to the F atom, this side of the molecule takes on a partial negative charge, which is represented by δ– (δ is the lowercase Greek letter delta). The other side of the molecule, the H atom, adopts a partial positive charge, which is represented by δ+. The two electrically charged regions on either end of the molecule are called poles, similar to a magnet having a north and a south pole. A molecule with two poles is called a **dipole** (see figure below). Hydrogen fluoride is a dipole.



Figure 4.6.2 A dipole is any molecule with a positive end and a negative end, resulting from unequal distribution of electron density throughout the molecule.

For molecules with more than two atoms, the molecular geometry must also be taken into account when determining if the molecule is polar or nonpolar. The figure below shows a comparison between carbon dioxide and water. Carbon dioxide $(\mathrm{CO_2})$ is a linear molecule. The oxygen atoms are more electronegative than the carbon atom, so there are two individual dipoles pointing outward from the C atom to each O atom. However, since the dipoles are of equal strength and are oriented this way, they cancel out and the overall molecular polarity of $\mathrm{CO_2}$ is zero.

Water is a bent molecule because of the two lone pairs on the central oxygen atom. The individual dipoles point from the H atoms toward the O atom. Because of the shape, the dipoles do not cancel each other out and the water molecule is polar. In the figure

below, the net dipole is shown in blue and points upward.

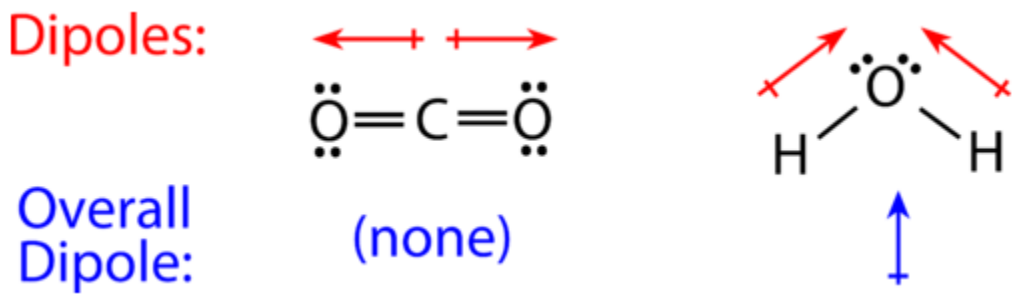


Figure 4.6.3 *The molecular geometry of a molecule affects its polarity. Each CO bond has a dipole moment, but they point in opposite directions so that the net CO2 molecule is nonpolar. In contrast, water is polar because the OH bond moments do not cancel out.*

Three other polar molecules are shown below with the arrows pointing to the more electron dense atoms. Just like the water molecule, none of the bond moments cancel out.

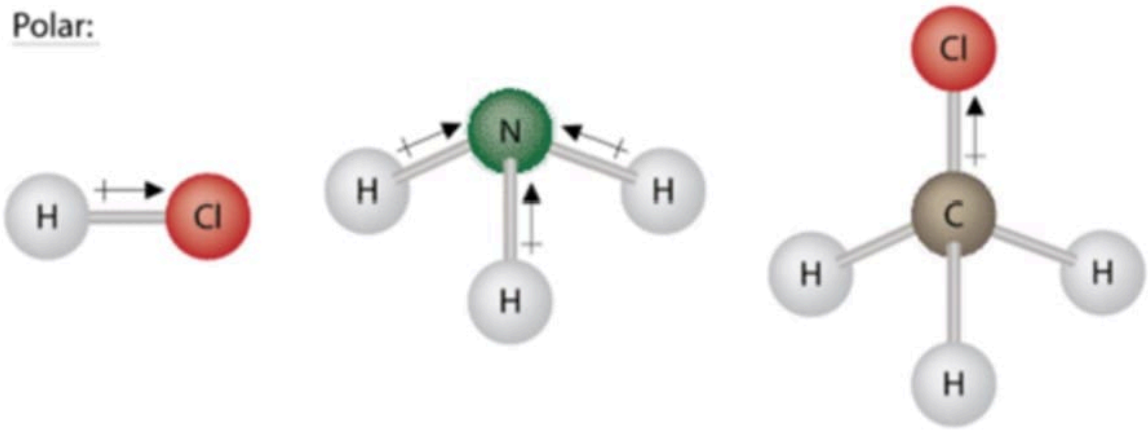


Figure 4.6.4 Some examples of polar molecules based on molecular geometry (HCl, NH3 and CH3Cl).

To summarize, to be polar, a molecule must:

1. Contain at least one polar covalent bond.
2. Have a molecular structure such that the sum of the vectors of each bond dipole moment does not cancel.

Steps to Identify Polar Molecules

1. Draw the Lewis structure
2. Figure out the geometry (using VSEPR theory)
3. Visualize or draw the geometry
4. Find the net dipole moment (you don't have to actually do calculations if you can visualize it)
5. If the net dipole moment is zero, it is non-polar. Otherwise, it is polar.

Example 4.6.1:

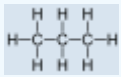
Label each of the following as polar or nonpolar.

a. Water, H_2O: 

b. Methanol, CH_3OH: 

c. Hydrogen Cyanide, HCN: 

d. Oxygen, O_2: 

e. Propane, C_3H_8: 

Solution

a. Water is polar. Any molecule with lone pairs of electrons around the central atom is polar.
b. Methanol is polar. This is not a symmetric molecule. The $-\mathrm{OH}$ side is different from the other 3 $-\mathrm{H}$ sides.
c. Hydrogen cyanide is polar. The molecule is not symmetric. The nitrogen and hydrogen have different electronegativities, creating an uneven pull on the electrons.
d. Oxygen is nonpolar. The molecule is symmetric. The two oxygen atoms pull on the electrons by exactly the same amount.
e. Propane is nonpolar, because it is symmetric, with H atoms bonded to every side around the central atoms and no unshared pairs of electrons.

Exercise $4.6.1$

Label each of the following as polar or nonpolar.

a. SO_3

b. NH_3

Answer available at the end of this section.

Summary

- Non polar molecules are symmetric with no unshared electrons.
- Polar molecules are asymmetric, either containing lone pairs of electrons on a central atom or having atoms with different electronegativities bonded.

Contributors and Attributions

-
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Answers

Exercise $4.6.1$
likely ionic
polar covalent

Exercise $4.6.1$
non polar
polar

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4.7: Intermolecular Forces

Learning Objectives

- Describe the types of intermolecular forces possible between atoms or molecules in condensed phases (dispersion forces, dipole-dipole attractions, and hydrogen bonding)
- Identify the types of intermolecular forces experienced by specific molecules based on their structures

Under appropriate conditions, the attractions between all gas molecules will cause them to form liquids or solids. This is due to intermolecular forces, not *intra*molecular forces. *Intra*molecular forces are those *within* the molecule that keep the molecule together, for example, the bonds between the atoms. *Inter*molecular forces are the attractions *between* molecules, which determine many of the physical properties of a substance. Figure 4.7.1 illustrates these different molecular forces. The strengths of these attractive forces vary widely, though usually the IMFs between small molecules are weak compared to the intramolecular forces that bond atoms together within a molecule. For example, to overcome the IMFs in one mole of liquid HCl and convert it into gaseous HCl requires only about 17 kilojoules. However, to break the covalent bonds between the hydrogen and chlorine atoms in one mole of HCl requires about 25 times more energy—430 kilojoules.

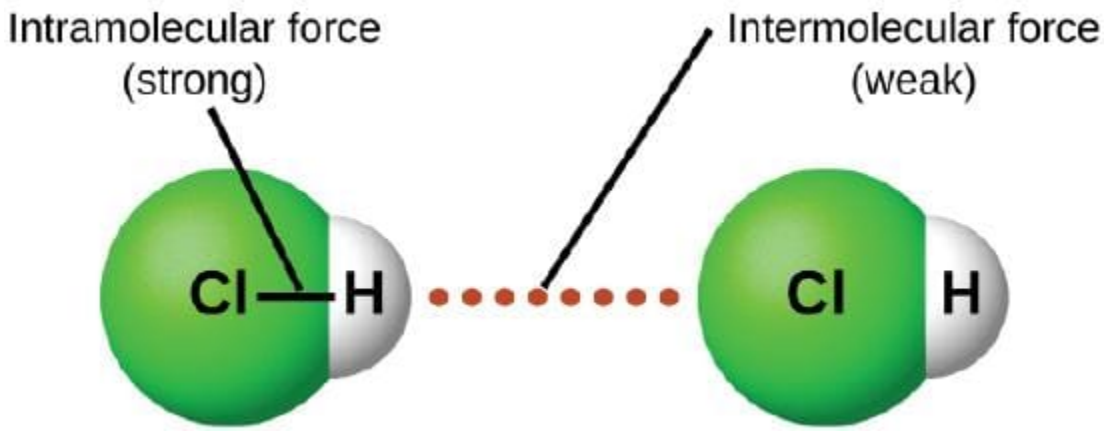


Figure 4.7.1 Intramolecular forces keep a molecule intact. Intermolecular forces hold multiple molecules together and determine many of a substance's properties.

An image is shown in which two molecules composed of a green sphere labeled "C l" connected on the right to a white sphere labeled "H" are near one another with a dotted line labeled "Intermolecular force (weak)" drawn between them. A line connects the two spheres in each molecule and the line is labeled "Intramolecular force (strong)."

Dipole–Dipole Forces

Dipole-dipole forces *are the attractive forces that occur between polar molecules* (see figure below). A molecule of hydrogen chloride has a partially positive hydrogen atom and a partially negative chlorine atom. A collection of many hydrogen chloride molecules will align themselves so that the oppositely charged regions of neighboring molecules are near each other.

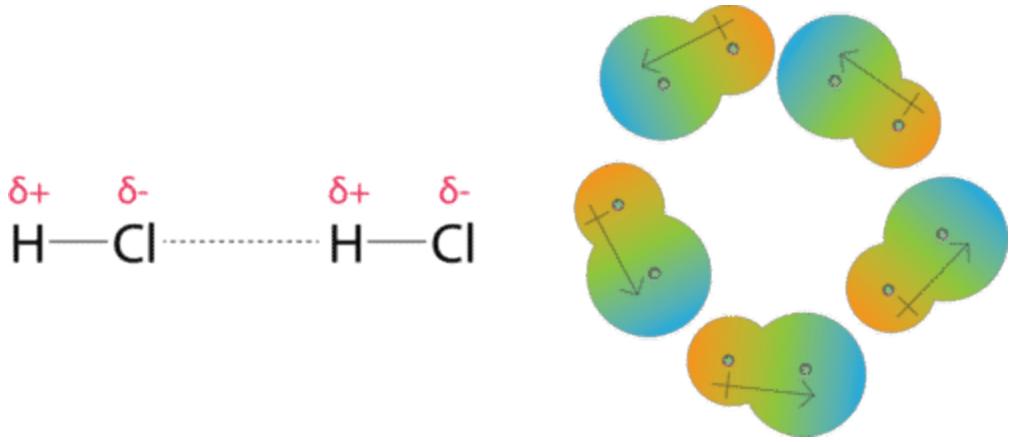


Figure 4.7.2 *Dipole-dipole forces result from the attraction between the positive end of one dipole and the negative end of a neighboring dipole. Dipole-dipole forces are similar to ionic bonds, but because they involve only partial charges, they are much weaker.*

Dipole-dipole forces in hydrogen chloride

The strength of the intermolecular interactions increases as the dipole moment of the molecules increases, as shown in Table 4.7.1.

Table 4.7.1: Relationships between the Dipole Moment and the Boiling Point for Organic Compounds of Similar Molar Mass

Compound	Molar Mass (g/mol)	Dipole Moment (D)	Boiling Point (K)
C_3H_6 (cyclopropane)	42	0	240
CH_3OCH_3 (dimethyl ether)	46	1.30	248
CH_3CN (acetonitrile)	41	3.9	355

The attractive energy between two ions is proportional to 1/r, whereas the attractive energy between two dipoles is proportional to 1/r6.

Example 4.7.1 Dipole-Dipole Forces and Their Effects

Predict which will have the higher boiling point: N_2 or CO. Explain your reasoning.

Solution

CO and N_2 are both diatomic molecules with masses of about 28 amu, so they experience similar London dispersion forces. Because CO is a polar molecule, it experiences dipole-dipole attractions. Because N_2 is nonpolar, its molecules cannot exhibit dipole-dipole attractions. The dipole-dipole attractions between CO molecules are comparably stronger than the dispersion forces between nonpolar N_2 molecules, so CO is expected to have the higher boiling point.

A common method for preparing oxygen is the decomposition

Exercise 4.7.1

Predict which will have the higher boiling point: ICl or $\mathrm{Br_2}$. Explain your reasoning.

Answer available at the end of this section.

Dispersion Forces

Dispersion forces are the weakest of all intermolecular forces. They are often called London forces after Fritz London (1900 -

1954), who first proposed their existence in 1930. **London dispersion forces** *are intermolecular forces that occur between all atoms and molecules due to the random motion of electrons.*

For example, the electron cloud of a helium atom contains two electrons, and, when averaged over time, these electrons will distribute themselves evenly around the nucleus. However, at any given moment, the electron distribution may be uneven, resulting in an *instantaneous dipole*. This weak and temporary dipole can subsequently influence neighboring helium atoms through electrostatic attraction and repulsion. The formation of an induced dipole is illustrated below.



Figure 4.7.3 *Random fluctuations in the electron density within the electron cloud of a helium atom results in a short-lived ("instantaneous") dipole. The attractive force between instantaneous dipoles and the resulting induced dipoles in neighboring molecules is called the London dispersion force.*

From left to right: instantaneous uneven distribution of electrons in helium atom; nonpolar helium atom; instantaneous dipole; induce dipole on neighboring helium atom press resulting in an attractive force

Dispersion forces that develop between atoms in different molecules can attract the two molecules to each other. The forces are relatively weak, however, and become significant only when the molecules are very close. Larger and heavier atoms and molecules exhibit stronger dispersion forces than do smaller and lighter atoms and molecules. F_2 and Cl_2 are gases at room temperature (reflecting weaker attractive forces); Br_2 is a liquid, and I_2 is a solid (reflecting stronger attractive forces). Trends in observed melting and boiling points for the halogens clearly demonstrate this effect, as seen in Table 4.7.1.

Table 4.7.2: Melting and Boiling Points of the Halogens

Halogen	Molar Mass	Atomic Radius	Melting Point	Boiling Point
fluorine, F_2	38 g/mol	72 pm	53 K	85 K
chlorine, Cl_2	71 g/mol	99 pm	172 K	238 K
bromine, Br_2	160 g/mol	114 pm	266 K	332 K
iodine, I_2	254 g/mol	133 pm	387 K	457 K
astatine, At_2	420 g/mol	150 pm	575 K	610 K

The increase in melting and boiling points with increasing atomic/molecular size may be rationalized by considering how the strength of dispersion forces is affected by the electronic structure of the atoms or molecules in the substance. In a larger atom, the valence electrons are, on average, farther from the nuclei than in a smaller atom. Thus, they are less tightly held and can more easily form the temporary dipoles that produce the attraction. The measure of how easy or difficult it is for another electrostatic charge (for example, a nearby ion or polar molecule) to distort a molecule's charge distribution (its electron cloud) is known as polarizability. A molecule that has a charge cloud that is easily distorted is said to be very polarizable and will have large dispersion forces; one with a charge cloud that is difficult to distort is not very polarizable and will have small dispersion forces.

For similar substances, London dispersion forces get stronger with increasing atomic or molecular size.

The strengths of London dispersion forces also depend significantly on molecular shape because shape determines how much of one molecule can interact with its neighboring molecules at any given time. For example, part (b) in Figure 4.7.4 shows 2,2-dimethylpropane (neopentane) and *n*-pentane, both of which have the empirical formula C_5H_{12}. Neopentane is almost spherical, with a small surface area for intermolecular interactions, whereas *n*-pentane has an extended conformation that enables it to come into close contact with other *n*-pentane molecules. As a result, the boiling point of neopentane (9.5°C) is more than 25°C lower than the boiling point of *n*-pentane (36.1°C).

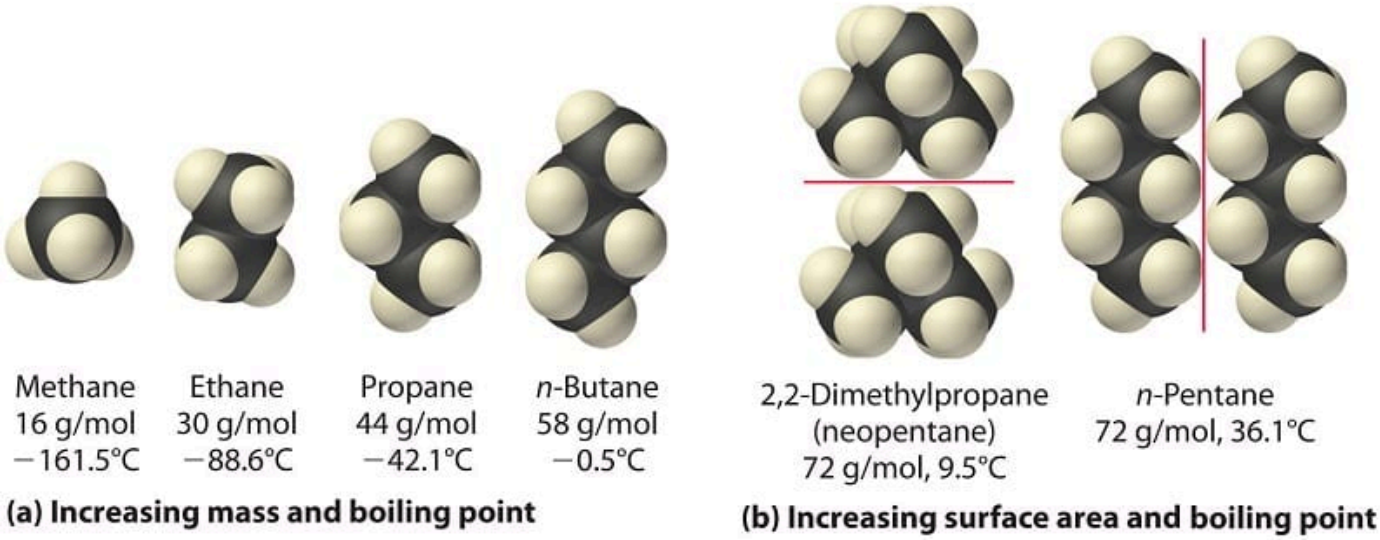


Figure 4.7.4 Mass and Surface Area Affect the Strength of London Dispersion Forces. (a) In this series of four simple alkanes, larger molecules have stronger London forces between them than smaller molecules and consequently higher boiling points. (b) Linear *n*-pentane molecules have a larger surface area and stronger intermolecular forces than spherical neopentane molecules. As a result, neopentane is a gas at room temperature, whereas *n*-pentane is a volatile liquid.

London dispersion forces

All molecules, whether polar or nonpolar, are attracted to one another by London dispersion forces in addition to any other attractive forces that may be present.

Example 4.7.2 London Forces and Their Effects

Order the following compounds of a group 14 element and hydrogen from lowest to highest boiling point: CH_4, SiH_4, GeH_4, and SnH_4. Explain your reasoning.

Given: compounds

Asked for: order of increasing boiling points

Strategyy:

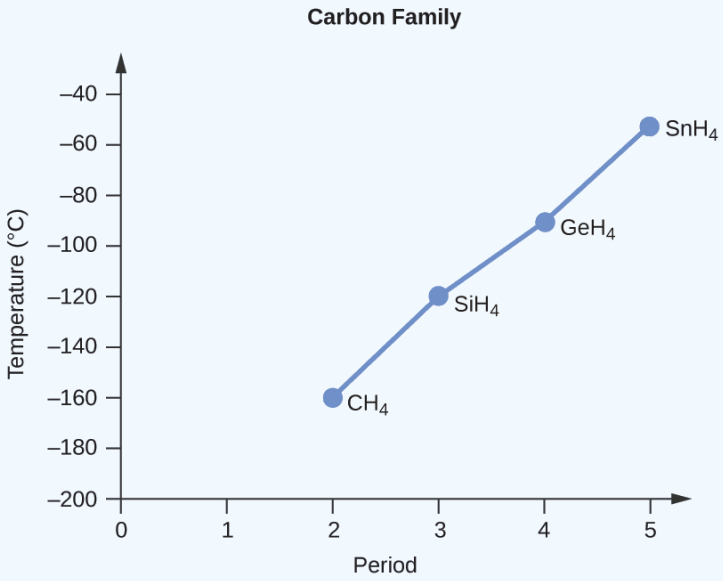
Determine the intermolecular forces in the compound. Also, compare the molar masses and surface area of compounds with similar types of intermolecular force. The substance with the weakest forces and the lowest molar masses and least surface area have the lowest boiling point.

Solution

Applying the skills acquired in the chapter on chemical bonding and molecular geometry, all of these compounds are predicted to be nonpolar, so they may experience only dispersion forces: the smaller the molecule, the less polarizable and the weaker the dispersion forces; the larger the molecule, the larger the dispersion forces. The molar masses of CH_4, SiH_4, GeH_4, and SnH_4 are approximately 16 g/mol, 32 g/mol, 77 g/mol, and 123 g/mol, respectively. Therefore, CH_4 is expected to have the lowest boiling point and SnH_4 the highest boiling point. The ordering from lowest to highest boiling point is expected to be

$$CH_4 < SiH_4 < GeH_4 < SnH_4$$

A graph of the actual boiling points of these compounds versus the period of the group 14 elements shows this prediction to be correct:



A line graph, titled “Carbon Family,” is shown where the y-axis is labeled “Temperature, (degree sign C)” and has values of “negative 200” to “negative 40” from bottom to top in increments of 20. The x-axis is labeled “Period” and has values of “0” to “5” in increments of 1. The first point on the graph is labeled “C H subscript 4” and is at point “2, negative 160.” The second point on the graph is labeled “S i H subscript 4” and is at point “3, negative 120” while the third point on the graph is labeled “G e H subscript 4” and is at point “4, negative 100.” The fourth point on the graph is labeled “S n H subscript 4” and is at point “5, negative 60.”

Exercise 4.7.2

Order the following hydrocarbons from lowest to highest boiling point: C_2H_6, C_3H_8, and C_4H_{10}.

Answer available at the end of this section.

Hydrogen Bonds

The attractive force between water molecules is an unusually strong type of dipole-dipole interaction. Water contains hydrogen atoms that are bound to a highly electronegative oxygen atom, making for very polar bonds. The partially positive hydrogen atom of one molecule is then attracted to the oxygen atom of a nearby water molecule (see figure below).

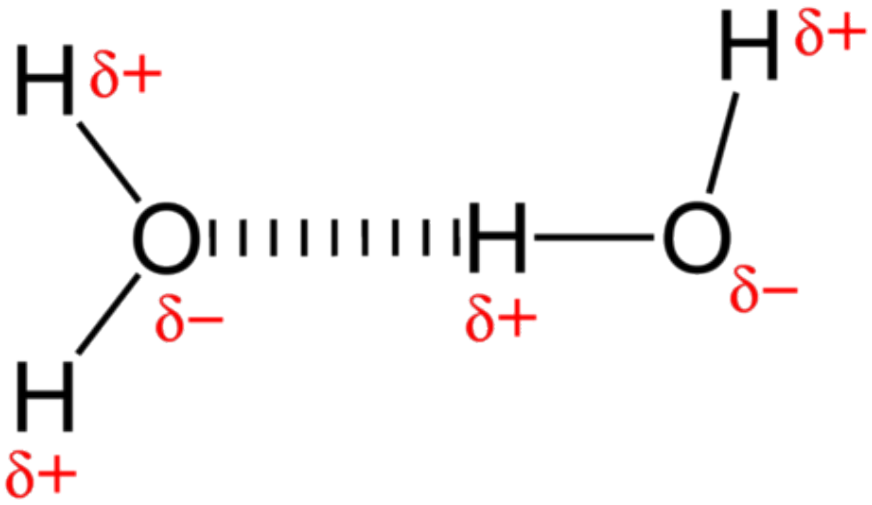


Figure 4.7.5 A hydrogen bond in water occurs between the hydrogen atom of one water molecule and the lone pair of electrons on the oxygen atom of a neighboring water molecule.

A **hydrogen bond** *is an intermolecular attractive force in which a hydrogen atom, that is covalently bonded to a small, highly electronegative atom, is attracted to a lone pair of electrons on an atom in a neighboring molecule*. Hydrogen bonds are very strong compared to other dipole-dipole interactions, but still much weaker than a covalent bond. A typical hydrogen bond is about 5% as strong as a covalent bond.

Hydrogen bonding occurs only in molecules where hydrogen is covalently bonded to one of three elements: fluorine, oxygen, or nitrogen. These three elements are so electronegative that they withdraw the majority of the electron density from the covalent bond with hydrogen, leaving the H atom very electron-deficient. Because the hydrogen atom does not have any electrons other than the ones in the covalent bond, its positively charged nucleus is almost completely exposed, allowing strong attractions to other nearby lone pairs of electrons.

The hydrogen bonding that occurs in water leads to some unusual, but very important properties. Most molecular compounds that have a mass similar to water are gases at room temperature. However, because of the strong hydrogen bonds, water molecules are able to stay condensed in the liquid state. The figure below shows how its bent shape and the presence of two hydrogen atoms per molecule allows each water molecule to hydrogen bond with several other molecules.

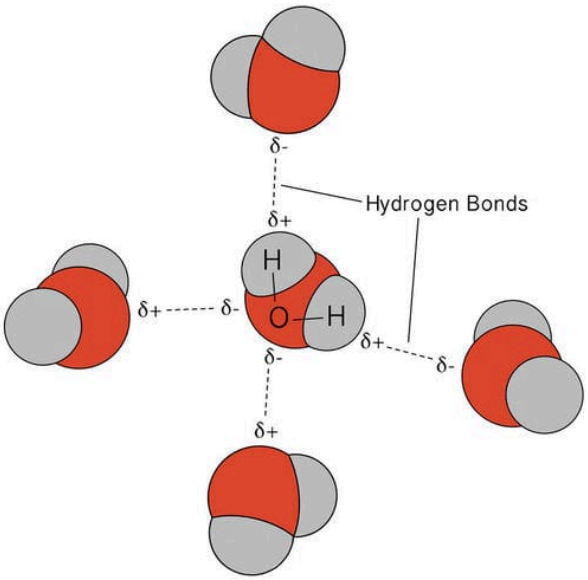


Figure 4.7.6Multiple hydrogen bonds occur simultaneously in water because of its bent shape and the presence of two hydrogen atoms per molecule.

In the liquid state, the hydrogen bonds of water can break and reform as the molecules flow from one place to another. When water is cooled, the molecules begin to slow down. Eventually, when water is frozen to ice, the hydrogen bonds become more rigid and form a well-defined network (Figure 4.7.7). The bent shape of the molecules leads to gaps in the hydrogen bonding network of ice. Ice has the very unusual property that its solid state is less dense than its liquid state. As a result, ice floats in liquid water. Virtually all other substances are denser in the solid state than in the liquid state. Hydrogen bonds also play a very important biological role in the physical structures of proteins and nucleic acids.

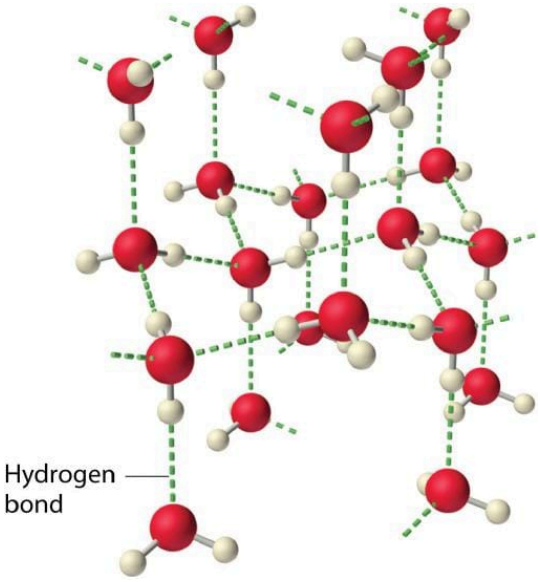


Figure 4.7.7 The Hydrogen-Bonded Structure of Ice.

Hydrogen bond formation requires both a hydrogen bond donor and a hydrogen bond acceptor.

Because ice is less dense than liquid water, rivers, lakes, and oceans freeze from the top down. In fact, the ice forms a protective surface layer that insulates the rest of the water, allowing fish and other organisms to survive in the lower levels of a frozen lake or sea. If ice were denser than the liquid, the ice formed at the surface in cold weather would sink as fast as it formed. Bodies of water would freeze from the bottom up, which would be lethal for most aquatic creatures. The expansion of water when freezing also explains why automobile or boat engines must be protected by “antifreeze” and why unprotected pipes in houses break if they are allowed to freeze.

Strength of Intermolecular Forces

Molecules with hydrogen atoms bonded to electronegative atoms such as O, N, and F tend to exhibit unusually strong intermolecular interactions due to hydrogen bonds as illustrated for the covalent hydrides of elements of groups 14–17 in Figure 4.7.8. These result in much higher boiling points than are observed for substances in which dipole-dipole forces or London dispersion forces dominate. All the polar molecules (with predominantly dipole-dipole forces) in group 16 (H_2S to H_2Te) have lower boiling points than H_2O (with hydrogen bonds). Methane and its heavier congeners in group 14 form a series whose boiling points increase smoothly with increasing molar mass. This is the expected trend in nonpolar molecules, for which London dispersion forces are the exclusive intermolecular forces. In contrast, the hydrides of the lightest members of groups 15–17 have boiling points that are more than 100°C greater than predicted on the basis of their molar masses. The effect is most dramatic for water: if we extend the straight line connecting the points for H_2Te and H_2Se to the line for period 2, we obtain an estimated boiling point of –130°C for water! Imagine the implications for life on Earth if water boiled at –130°C rather than 100°C.

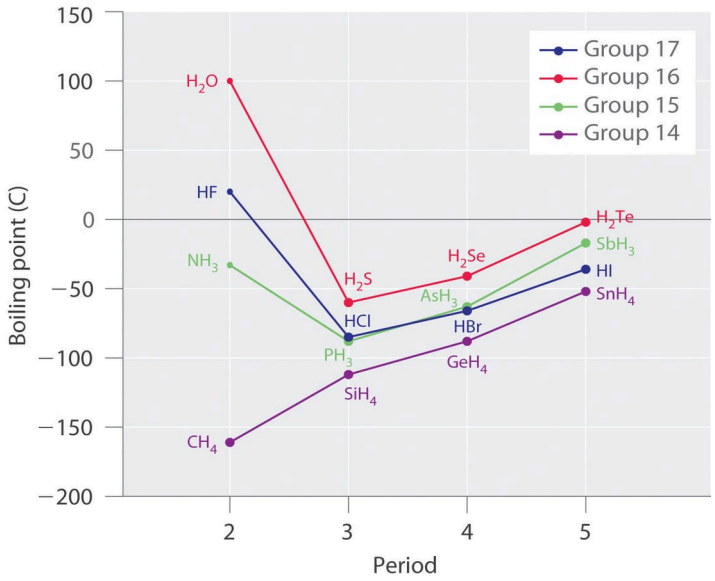


Figure 4.7.8 The Effects of Hydrogen Bonding on Boiling Points. These plots of the boiling points of the covalent hydrides of the elements of groups 14–17 show that the boiling points of the lightest members of each series for which hydrogen bonding is possible (HF, NH_3, and H_2O) are anomalously high for compounds with such low molecular masses.

Example 4.7.3

Considering CH_3OH, C_2H_6, Xe, and $(CH_3)_3N$, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Given: compounds

Asked for: formation of hydrogen bonds and structure

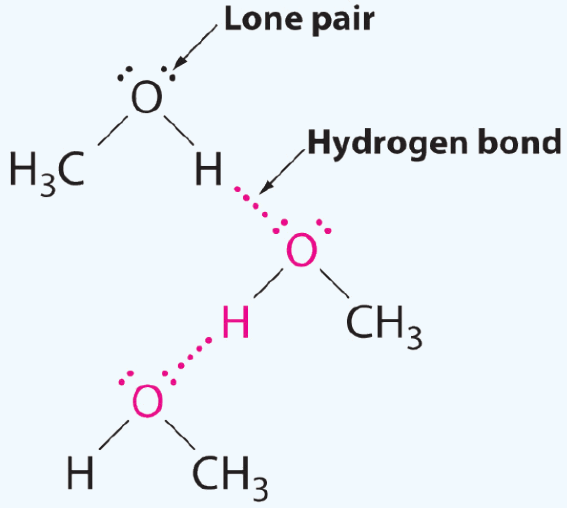
Strategy:

A. Identify the compounds with a hydrogen atom attached to O, N, or F. These are likely to be able to act as hydrogen bond donors.
B. Of the compounds that can act as hydrogen bond donors, identify those that also contain lone pairs of electrons, which allow them to be hydrogen bond acceptors. If a substance is both a hydrogen donor and a hydrogen bond acceptor, draw a structure showing the hydrogen bonding.

Solution:

A Of the species listed, xenon (Xe), ethane (C_2H_6), and trimethylamine [$(CH_3)_3N$] do not contain a hydrogen atom attached to O, N, or F; hence they cannot act as hydrogen bond donors.

B The one compound that can act as a hydrogen bond donor, methanol (CH_3OH), contains both a hydrogen atom attached to O (making it a hydrogen bond donor) and two lone pairs of electrons on O (making it a hydrogen bond acceptor); methanol can thus form hydrogen bonds by acting as either a hydrogen bond donor or a hydrogen bond acceptor. The hydrogen-bonded structure of methanol is as follows:



Hydrogen bonding in methanol

Exercise 4.7.3

Considering CH_3CO_2H, $(CH_3)_3N$, NH_3, and CH_3F, which can form hydrogen bonds with themselves? Draw the hydrogen-bonded structures.

Answer available at the end of this section.

Although hydrogen bonds are significantly weaker than covalent bonds, with typical dissociation energies of only 15–25 kJ/mol, they have a significant influence on the physical properties of a compound. Compounds such as HF can form only two hydrogen bonds at a time as can, on average, pure liquid NH_3. Consequently, even though their molecular masses are similar to that of water, their boiling points are significantly lower than the boiling point of water, which forms *four* hydrogen bonds at a time.

Example 4.7.4: Buckyballs

Arrange C_{60} (buckminsterfullerene, which has a cage structure), NaCl, He, Ar, and N_2O in order of increasing boiling points.

Given: compounds

Asked for: order of increasing boiling points

Strategy:

Identify the intermolecular forces in each compound and then arrange the compounds according to the strength of those forces. The substance with the weakest forces will have the lowest boiling point.

Solution:

Electrostatic interactions are strongest for an ionic compound, so we expect NaCl to have the highest boiling point. To predict the relative boiling points of the other compounds, we must consider their polarity (for dipole–dipole interactions), their ability to form hydrogen bonds, and their molar mass (for London dispersion forces). Helium is nonpolar and by far the lightest, so it should have the lowest boiling point. Argon and N_2O have very similar molar masses (40 and 44 g/mol, respectively), but N_2O is polar while Ar is not. Consequently, N_2O should have a higher boiling point. A C_{60} molecule is nonpolar, but its molar mass is 720 g/mol, much greater than that of Ar or N_2O. Because the boiling points of nonpolar substances increase rapidly with molecular mass, C_{60} should boil at a higher temperature than the other nonionic substances. The predicted order is thus as follows, with actual boiling points in parentheses:

He (−269°C) < Ar (−185.7°C) < N_2O (−88.5°C) < C_{60} (>280°C) < NaCl (1465°C).

Exercise 4.7.4

Arrange 2,4-dimethylheptane, Ne, CS_2, Cl_2, and KBr in order of decreasing boiling points.

Answer available at the end of this section.

Example 4.7.5:

Identify the most significant intermolecular force in each substance.

a. C_3H_8
b. CH_3OH
c. H_2S

Solution

a. Although C–H bonds are polar, they are only minimally polar. The most significant intermolecular force for this substance would be dispersion forces.

b. This molecule has an H atom bonded to an O atom, so it will experience hydrogen bonding.

c. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram and VSEPR indicate that it is bent, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.

Exercise 4.7.5

Identify the most significant intermolecular force in each substance.

a. HF
b. HCl

Answer a available at the end of this section.

Summary

- Molecules in liquids are held to other molecules by intermolecular interactions, which are weaker than the intramolecular interactions that hold the atoms together within molecules and polyatomic ions.
- **Dipole–dipole interactions** arise from the electrostatic interactions of the positive and negative ends of molecules with permanent dipole moments.
- **London dispersion forces** are due to the formation of **instantaneous dipole moments** in polar and nonpolar molecules as a result of short-lived fluctuations of electron charge distribution, which in turn cause the temporary formation of an **induced dipole** in adjacent molecules.
- **Hydrogen bonds** are especially strong dipole–dipole interactions between molecules that have hydrogen bonded to a highly electronegative atom, such as O, N, or F. The resulting partially positively charged H atom on one molecule (the *hydrogen bond donor*) can interact strongly with a lone pair of electrons of a partially negatively charged O, N, or F atom on adjacent molecules (the *hydrogen bond acceptor*).

Contributors and Attributions

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- Marisa Alviar-Agnew (Sacramento City College)

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Learning Objective

- Relate phase to intermolecular forces.

Why does a substance have the phase that it does? The preferred phase of a substance at a given set of conditions is a balance between the energy of the particles and intermolecular forces (or intermolecular interactions) between the particles. If the forces between particles are strong enough, the substance is a liquid or, if stronger, a solid. If the forces between particles are weak and sufficient energy is present, the particles separate from each other, so the gas phase is the preferred phase. The energy of the particles is mostly determined by temperature, so temperature is the main variable that determines what phase is stable at any given point.

What forces define intermolecular interactions? There are several. A force present in all substances with electrons is the **dispersion force** (sometimes called the *London dispersion force*, after the physicist Fritz London, who first described this force in the early 1900s). This interaction is caused by the instantaneous position of an electron in a molecule, which temporarily makes that point of the molecule negatively charged and the rest of the molecule positively charged. In an instant, the electron is now somewhere else, but the fleeting imbalance of electric charge in the molecule allows molecules to interact with each other. As you might expect, the greater the number of electrons in a species, the stronger the dispersion force; this partially explains why smaller molecules are gases and larger molecules are liquids and solids at the same temperature. (Mass is a factor as well.)

Molecules with a permanent dipole moment experience **dipole-dipole interactions,** which are generally stronger than dispersion forces if all other things are equal. The oppositely charged ends of a polar molecule, which have partial charges on them, attract each other (Figure 4.7.1). Thus, a polar molecule such as CH_2Cl_2 has a significantly higher boiling point (313 K, or 40°C) than a nonpolar molecule like CF_4 (145 K, or –128°C), even though it has a lower molar mass (85 g/mol vs. 88 g/mol).



Figure 4.7.1: Dipole-Dipole Interactions. Oppositely charged ends of polar molecules attract each other.
2 ovals with partial positive left & right sides are attracted to each other. The right of oval 1 is attracted to the left of oval 2.

An unusually strong form of dipole-dipole interaction is called **hydrogen bonding**. Hydrogen bonding is found in molecules with an H atom bonded to an N atom, an O atom, or an F atom. Such covalent bonds are very polar, and the dipole-dipole interaction between these bonds in two or more molecules is strong enough to create a new category of intermolecular force. Hydrogen bonding is the reason water has unusual properties. For such a small molecule (its molar mass is only 18 g/mol), H_2O has relatively high melting and boiling points. Its boiling point is 373 K (100°C), while the boiling point of a similar molecule, H_2S, is 233 K (–60°C). This is because H_2O molecules experience hydrogen bonding, while H_2S molecules do not. This strong attraction between H_2O molecules requires additional energy to separate the molecules in the condensed phase, so its boiling point is higher than would be expected. Hydrogen bonding is also responsible for water's ability as a solvent, its high heat capacity, and its ability to expand when freezing; the molecules line up in such a way that there is extra space between the molecules, increasing its volume in the solid state (Figure 4.7.2).

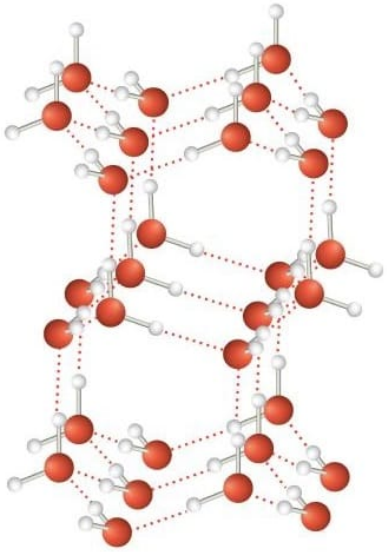


Figure 4.7.2: Hydrogen Bonding. When water solidifies, hydrogen bonding between the molecules forces the molecules to line up in a way that creates empty space between the molecules, increasing the overall volume of the solid. This is why ice is less dense than liquid water.

Example 4.7.1

Identify the most significant intermolecular force in each substance.

a. C_3H_8
b. CH_3OH
c. H_2S

Solution

1. Although C–H bonds are polar, they are only minimally polar. The most significant intermolecular force for this substance would be dispersion forces.
2. This molecule has an H atom bonded to an O atom, so it will experience hydrogen bonding.
3. Although this molecule does not experience hydrogen bonding, the Lewis electron dot diagram and VSEPR indicate that it is bent, so it has a permanent dipole. The most significant force in this substance is dipole-dipole interaction.

Exercise 4.7.1

Identify the most significant intermolecular force in each substance.

a. HF
b. HCl

Answers

a. hydrogen bonding
b. dipole-dipole interactions

The preferred phase a substance adopts can change with temperature. At low temperatures, most substances are solids (only helium is predicted to be a liquid at absolute zero). As the temperature increases, those substances with very weak intermolecular forces become gases directly (in a process called *sublimation*, which will be discussed in Section 10.3). Substances with weak interactions can become liquids as the temperature increases. As the temperature increases even more, the individual particles will have so much energy that the intermolecular forces are overcome, so the particles separate from each other, and the substance becomes a gas (assuming that their chemical bonds are not so weak that the compound decomposes from the high

temperature). Although it is difficult to predict the temperature ranges for which solid, liquid, or gas is the preferred phase for any random substance, all substances progress from solid to liquid to gas (in that order) as temperature increases.

Summary

- All substances experience dispersion forces between their particles.
- Substances that are polar experience dipole-dipole interactions.
- Substances with covalent bonds between an H atom and N, O, or F atoms experience hydrogen bonding.
- The preferred phase of a substance depends on the strength of the intermolecular force and the energy of the particles.

Answers

Exercise 4.7.1

ICl. ICl and Br_2 have similar masses (~160 amu) and therefore experience similar London dispersion forces. ICl is polar and thus also exhibits dipole-dipole attractions; Br_2 is nonpolar and does not. The relatively stronger dipole-dipole attractions require more energy to overcome, so ICl will have the higher boiling point

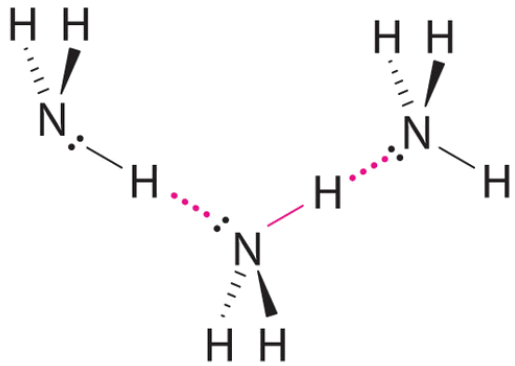
Exercise 4.7.2

All of these compounds are nonpolar and only have London dispersion forces: the larger the molecule, the larger the dispersion forces and the higher the boiling point. The ordering from lowest to highest boiling point is therefore

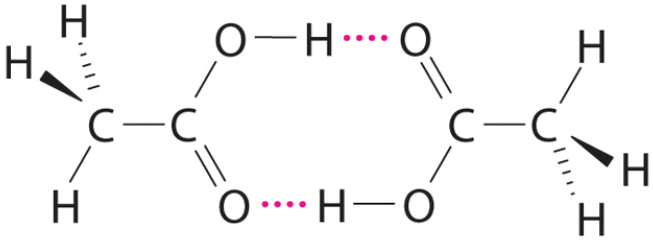
$$C_2H_6 < C_3H_8 < C_4H_{10}.$$

Exercise 4.7.3

CH_3CO_2H and NH_3;



Hydrogen bonding in ammonia



Hydrogen bonding in acetic acid

Exercise 4.7.4

KBr (1435°C) > 2,4-dimethylheptane (132.9°C) > CS_2 (46.6°C) > Cl_2 (–34.6°C) > Ne (–246°C)

Exercise 4.7.5

hydrogen bonding

dipole-dipole interactions

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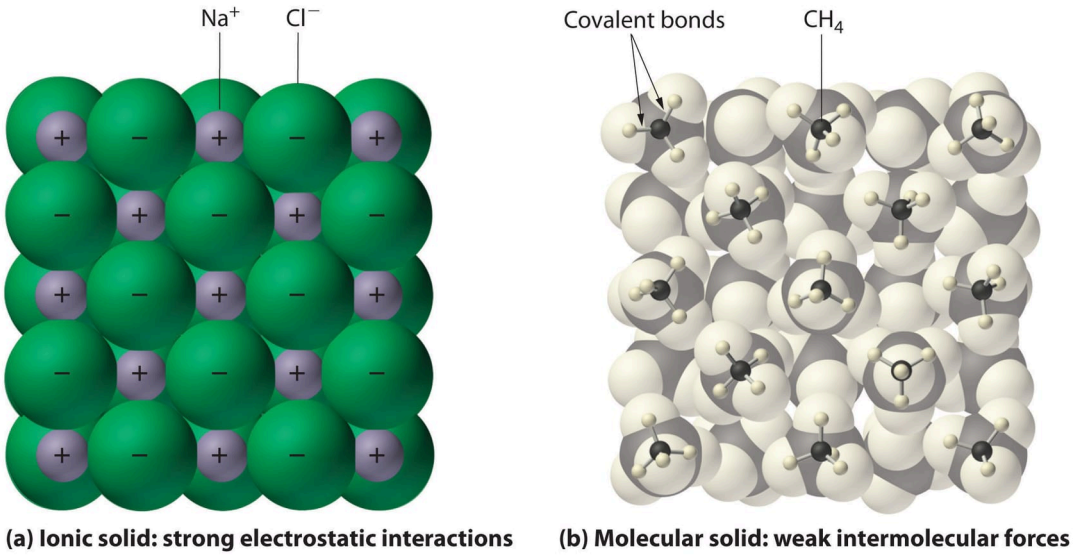
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4.8: Impact of Intermolecular Forces on Physical Properties

Learning Objectives

- Know the physical properties of ionic and molecular substances.

The physical state and properties of a particular compound depend in large part on the type of chemical bonding it displays. **Molecular compounds**, sometimes called



(a) The positively and negatively charged ions in an ionic solid such as sodium chloride (NaCl) are held together by strong electrostatic interactions. (b) In this representation of the packing of methane (CH4) molecules in solid methane, a prototypical molecular solid, the methane molecules are held together in the solid only by relatively weak intermolecular forces, even though the atoms within each methane molecule are held together by strong covalent bonds.

covalent compounds, display a wide range of physical properties due to the different types of intermolecular attractions such as different kinds of polar interactions. The melting and boiling points of molecular compounds are generally quite low compared to those of **ionic compounds**. This is because the energy required to disrupt the **intermolecular forces** between molecules is far less than the energy required to break the ionic bonds in a crystalline ionic compound (Figure 4.8.1) . Ionic solids typically melt at high temperatures and boil at even higher temperatures. For example, sodium chloride melts at 801 °C and boils at 1413 °C. (As a comparison, the molecular compound water melts at 0 °C and boils at 100 °C.). The water solubility of molecular compounds is variable and depends primarily on the type of intermolecular forces involved.

Figure 4.8.1 Interactions in Ionic and Covalent Solids.

Since molecular compounds are composed of neutral molecules, their electrical conductivity is generally quite poor, whether in the solid or liquid state. In solid form, an ionic compound is not electrically conductive because its ions are unable to flow ("electricity" is the flow of charged particles). When molten, however, it can conduct electricity because its ions are able to move freely through the liquid (Figure 4.8.2; Video 4.8.1).

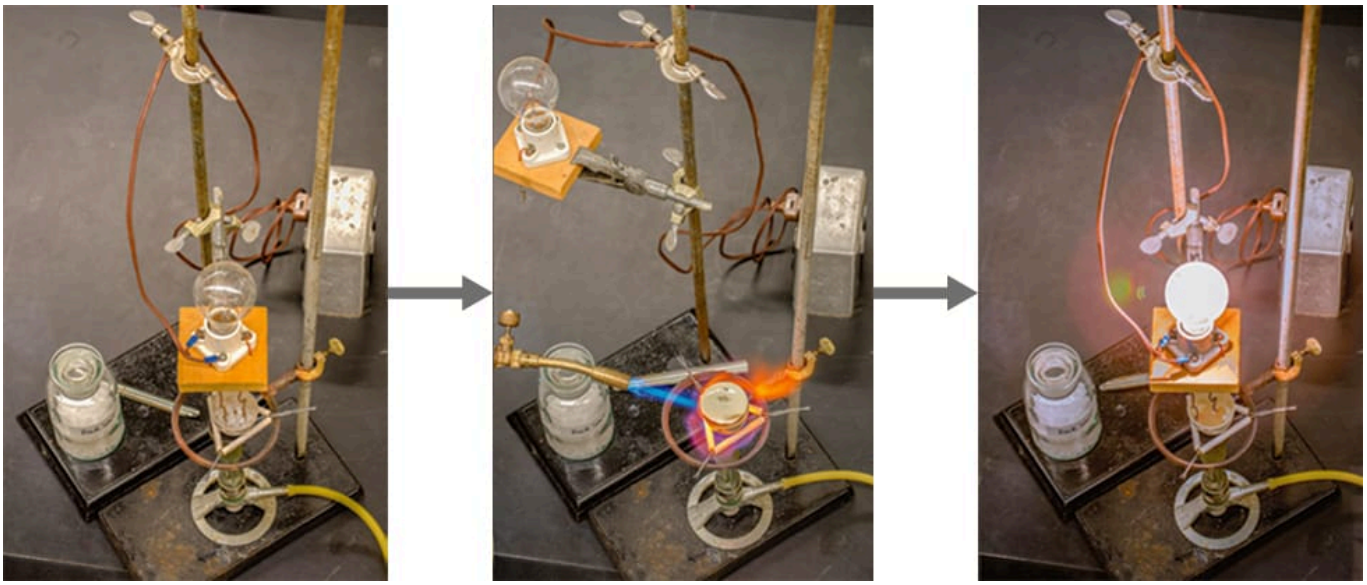


Figure 4.8.2 Sodium chloride melts at 801 °C and conducts electricity when molten. (credit: modification of work by Mark Blaser and Matt Evans)

This figure shows three photos connected by right-facing arrows. The first shows a light bulb as part of a complex lab equipment setup. The light bulb is not lit. The second photo shows a substances being heated or set on fire. The third shows the light bulb again which is lit.

Conductivity of Molten Salt

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video** 4.8.1 **Watch this video to see a mixture of salts melt and conduct electricity.

The table below summarizes some of the differences between ionic and molecular compounds.

Table** 4.8.1 **Comparison of Ionic and Molecular Compounds

Property	**Ionic Compounds**	**Molecular Compounds**
Type of elements	Metal and nonmetal	Nonmetals only
Bonding	Ionic - transfer of electron(s) between atoms	Covalent - sharing of pair(s) of electrons between atoms
Representative unit	Formula unit	Molecule
Physical state at room temperature	Solid	Gas, liquid, or solid
Water solubility	Usually high	Variable
Melting and boiling temperatures	Generally high	Generally low
Electrical conductivity	Good when molten or in solution	Poor

One type of molecular compound behaves quite differently than that described so far. A covalent network solid is a compound in which all of the atoms are connected to one another by covalent bonds. Diamond is composed entirely of carbon atoms, each bonded to four other carbon atoms in a tetrahedral geometry. Melting a covalent network solid is not accomplished by overcoming the relatively weak intermolecular forces. Rather, all of the covalent bonds must be broken, a process that requires extremely high

temperatures. Diamond, in fact, does not melt at all. Instead, it vaporizes to a gas at temperatures above $3500^{\circ}\mathrm{C}$.

Summary

- The physical properties of a material are affected by the intermolecular forces holding the molecules together.
- Ionic compounds usually form hard crystalline solids with high melting points. Covalent molecular compounds, in contrast, consist of discrete molecules held together by weak intermolecular forces and can be gases, liquids, or solids at room temperature and pressure.
- Ionic compounds in molten form or in solution can conduct electricity while molecular compounds do not..

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- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
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Learning Objectives

- Learn some terminology involving solutions.
- Explain the significance of the statement "like dissolves like".
- Explain why certain substances dissolve in other substances.

A **solution** is another name for a homogeneous mixture. A **mixture** as a material composed of two or more substances. In a solution, the combination is so intimate that the



Figure 4.8.1 Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

different substances cannot be differentiated by sight, even with a microscope. Compare, for example, a mixture of salt and pepper and another mixture consisting of salt and water. In the first mixture, we can readily see individual grains of salt and the flecks of pepper. A mixture of salt and pepper is not a solution. However, in the second mixture, no matter how carefully we look, we cannot see two different substances. Salt dissolved in water is a solution.

The major component of a solution, called the **solvent**, is typically the same phase as the solution itself. Each minor component of a solution (and there may be more than one) is called the **solute**. In most of the solutions we will describe in this textbook, there will be no ambiguity about whether a component is the solvent or the solute.) For example, in a solution of salt in water, the solute is salt, and solvent is water Figure 4.8.1.

Solutions come in all phases, and the solvent and the solute do not have to be in the same phase to form a solution (such as salt and water). For example, air is a gaseous solution of about 80% nitrogen and about 20% oxygen, with some other gases present in much smaller amounts. An alloy is a solid solution consisting of a metal (like iron) with some other metals or nonmetals dissolved in it. Steel, an alloy of iron and carbon and small amounts of other metals, is an example of a solid solution. Table 4.8.1 lists some common types of solutions, with examples of each.

Table 4.8.1 Types of Solutions

Solvent Phase	Solute Phase	Example
gas	gas	air
liquid	gas	carbonated beverages
liquid	liquid	ethanol (C_2H_5OH) in H_2O (alcoholic beverages)
liquid	solid	salt water
solid	gas	H_2 gas absorbed by Pd metal
solid	liquid	Hg(ℓ) in dental fillings
solid	solid	steel alloys

Example 4.8.1: Sugar and Water

A solution is made by dissolving 1.00 g of sucrose ($C_{12}H_{22}O_{11}$) in 100.0 g of liquid water. Identify the solvent and solute in the resulting solution.

Solution

Either by mass or by moles, the obvious minor component is **sucrose**, so it is **the solute**. **Water**—the majority component—is **the solvent**. The fact that the resulting solution is the same phase as water also suggests that water is the solvent.

Exercise 4.8.1

A solution is made by dissolving 3.33 g of $HCl(g)$ in 40.0 g of liquid methyl alcohol (CH_3OH). Identify the solvent and solute in the resulting solution.

Answer available at the end of this section.

Like Dissolves Like

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

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

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

Table 4.8.2 Summary of Solubilities

Solute	Solvent	Is Solution Formed?
Polar Covalent	Polar	yes
Non-polar Covalent	Non-polar	yes
Polar Covalent	Non-polar	no
Non-polar Covalent	Polar	no
Ionic	Polar	yes
Ionic	Non-polar	no

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

> When an ionic compound is considered "insoluble", it doesn't necessarily mean the compound is completely untouched by water. All ionic compounds dissolve to some extent. An insoluble compound just doesn't dissolve in any noticeable or appreciable amount.

What is it that makes a solute soluble in some solvents but not others?

The answer is intermolecular interactions. The intermolecular interactions include London dispersion forces, dipole-dipole interactions, and hydrogen bonding (as described in



Figure 4.8.1 Making a saline water solution by dissolving table salt (NaCl) in water. The salt is the solute and the water the solvent. (CC-BY-SA 3.0; Chris 73).

the previous section). From experimental studies, it has been determined that if molecules of a solute experience the same intermolecular forces that the solvent does, the solute will likely dissolve in that solvent. So, NaCl—a very polar substance because it is composed of ions—dissolves in water, which is very polar, but not in oil, which is generally nonpolar. Nonpolar wax dissolves in nonpolar hexane, but not in polar water. *Liquids that dissolve in one another in all proportions are said to be* **miscible**. *Liquids that do not dissolve in one another are called* **immiscible**.

Figure 4.8.2 Water (clear liquid) and oil (yellow) do not form liquid solutions. (CC BY-SA 1.0 Generic; Victor Blacus)

Example 4.8.2: Polar and Nonpolar Solvents

Would I_2 be more soluble in CCl_4 or H_2O? Explain your answer.

Solution

I_2 is nonpolar. Of the two solvents, CCl_4 is nonpolar and H_2O is polar, so I_2 would be expected to be more soluble in CCl_4.

Exercise 4.8.2

Would C_3H_7OH be more soluble in CCl_4 or H_2O? Explain your answer.

Answer available at the end of this section.

Example 4.8.3

Water is considered a polar solvent. Which substances should dissolve in water?

a. methanol (CH_3OH)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Solution

Because water is polar, substances that are polar or ionic will dissolve in it.

a. Because of the OH group in methanol, we expect its molecules to be polar. Thus, we expect it to be soluble in water. As both water and methanol are liquids, the word *miscible* can be used in place of *soluble*.
b. Sodium sulfate is an ionic compound, so we expect it to be soluble in water.
c. Like other hydrocarbons, octane is nonpolar, so we expect that it would not be soluble in water.

Exercise 4.8.3

Toluene ($C_6H_5CH_3$) is widely used in industry as a nonpolar solvent. Which substances should dissolve in toluene?

a. water (H_2O)
b. sodium sulfate (Na_2SO_4)
c. octane (C_8H_{18})

Answer available at the end of this section.

The Dissolving Process

Water typically dissolves most ionic compounds and polar molecules. Nonpolar molecules, such as those found in grease or oil, do not dissolve in water. We will first examine the process that occurs when an ionic compound, such as table salt (sodium chloride), dissolves in water.

Water molecules move about continuously due to their kinetic energy. When a crystal of sodium chloride is placed into water, the water's molecules collide with the crystal lattice. Recall that the crystal lattice is composed of alternating positive and negative

ions. Water is attracted to the sodium chloride crystal because water is polar; it has both a positive and a negative end. The positively charged sodium ions in the crystal attract the oxygen end of the water molecules because they are partially negative. The negatively charged chloride ions in the crystal attract the hydrogen end of the water molecules because they are partially positive. The action of the polar water molecules takes the crystal lattice apart (see figure below).

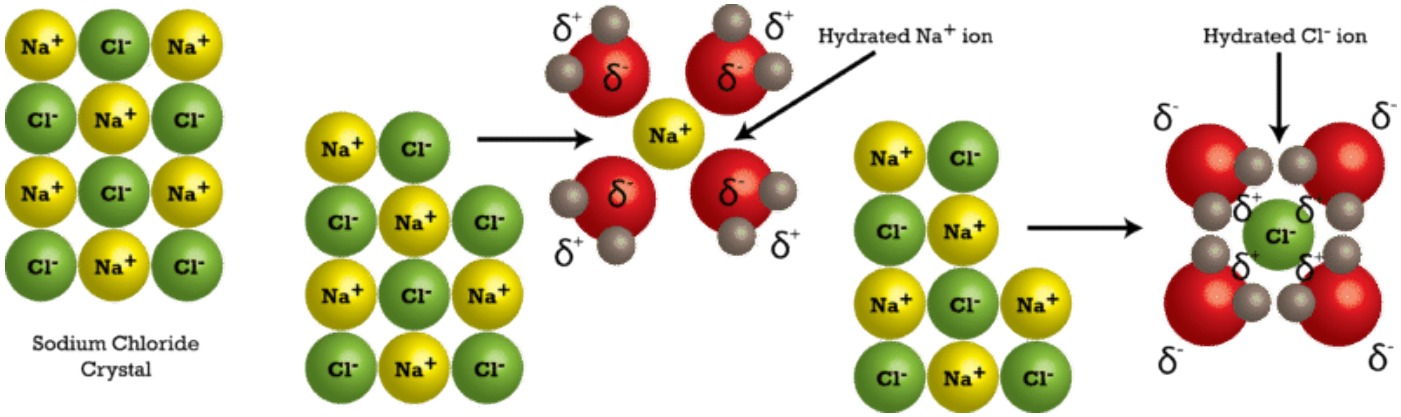


Figure 4.8.3: *The process of an ionic sodium chloride crystal being dissolved and hydrated by water.*

After coming apart from the crystal, *the individual ions are then surrounded by solvent particles in a process called* **solvation**. Note in the figure above that the individual $\mathrm{Na^+}$ ions are surrounded by water molecules with the oxygen atom oriented near the positive ion. Likewise, the chloride ions are surrounded by water molecules with the opposite orientation. Thus, numerous **ion-dipole** interactions are formed. **Hydration** *is the process of solute particles being surrounded by water molecules arranged in a specific manner*. Hydration helps to stabilize aqueous solutions by preventing the positive and negative ions from coming back together and forming a precipitate.

Table sugar is made of the molecular compound sucrose $(\mathrm{C_{12}H_{22}O_{11}})$. Solid sugar consists of individual sugar molecules held together by intermolecular attractive forces. When water dissolves sugar, it separates the individual sugar molecules by disrupting the attractive forces, but it *does not break the covalent bond*s between the carbon, hydrogen, and oxygen atoms. Dissolved sugar molecules are also hydrated. The hydration shell around a molecule of sucrose is arranged so that its partially negative oxygen atoms are near the partially positive hydrogen atoms in the solvent, and vice versa.

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- “Like dissolves like” is a useful rule for deciding if a solute will be soluble in a solvent.
- Liquids that dissolve in one another in all proportions are said to be **miscible**.
- Liquids that do not dissolve in one another are called **immiscible**.
- **Ion-dipole** interactions are formed when ionic compounds dissolve in water.

Contributors and Attributions

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- Libretext: The Basics of GOB Chemistry (Ball et al.)

Learning Objective

- Define the vapor pressure of liquids.
- Explain the origin of both surface tension and capillary action.

There are some properties that all liquids, including water, have. All liquids have a certain portion of particles with enough energy to enter the gas phase, and if these particles are at the surface of the liquid, they do so (Figure 4.8.1). The formation of a gas from a liquid at temperatures below the boiling point is called **evaporation**. At these temperatures, the material in the gas phase is called **vapor**, rather than gas; the term *gas* is reserved for when the gas phase is the stable phase.

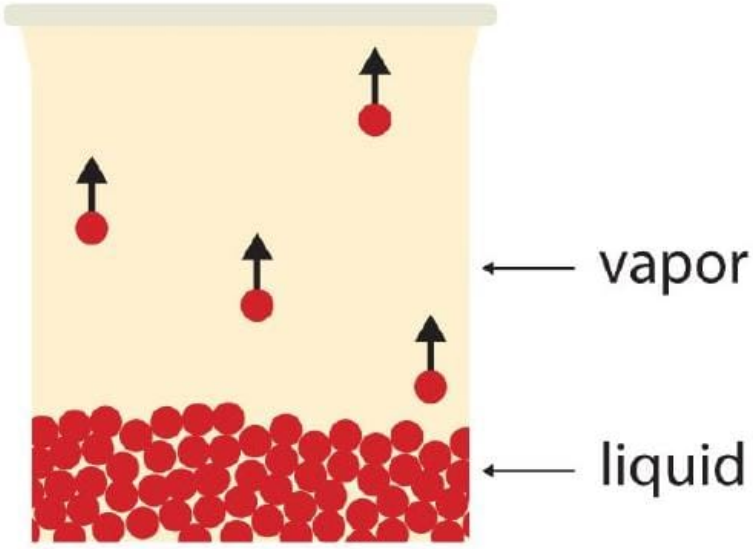


Figure 4.8.1: Evaporation. Some particles of a liquid have enough energy to escape the liquid phase to become a vapor.

If the available volume is large enough, eventually all the liquid will become vapor. But if the available volume is not enough, eventually some of the vapor particles will reenter the liquid phase (Figure 4.8.2 Equilibrium). At some point, the number of particles entering the vapor phase will equal the number of particles leaving the vapor phase, so there is no net change in the amount of vapor in the system. We say that the system is *at equilibrium*. The partial pressure of the vapor at equilibrium is called the *vapor pressure of the liquid*.

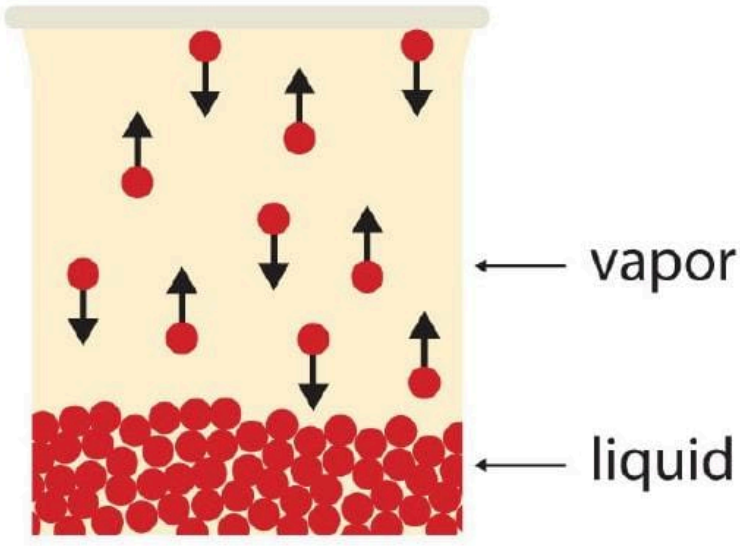


Figure 4.8.2: Equilibrium. At some point, the number of particles entering the vapor phase will be balanced by the number of particles returning to the liquid. This point is called equilibrium.

Understand that the liquid has not stopped evaporating. The reverse process—condensation—is occurring as fast as evaporation, so there is no net change in the amount of vapor in the system. The term **dynamic equilibrium** represents a situation in which a process still occurs, but the opposite process also occurs at the same rate, so that there is no net change in the system.

The vapor pressure for a substance is dependent on the temperature of the substance; as the temperature increases, so does the

vapor pressure. Figure 4.8.3 - Plots of Vapor Pressure versus Temperature for Several Liquids, is a plot of vapor pressure versus temperature for several liquids. Having defined vapor pressure, we can also redefine the *boiling point* of a liquid: the temperature at which the vapor pressure of a liquid equals the surrounding environmental pressure. The normal vapor pressure, then, is the temperature at which the vapor pressure is 760 torr, or exactly 1 atm. Thus boiling points vary with surrounding pressure, a fact that can have large implications on cooking foods at lower- or higher-than-normal elevations. Atmospheric pressure varies significantly with altitude.

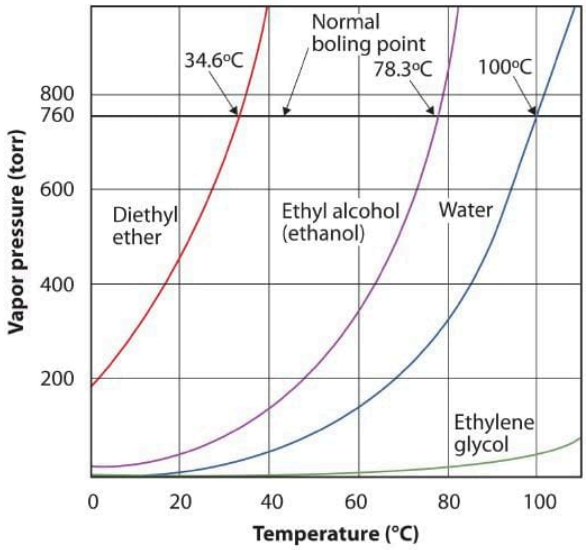


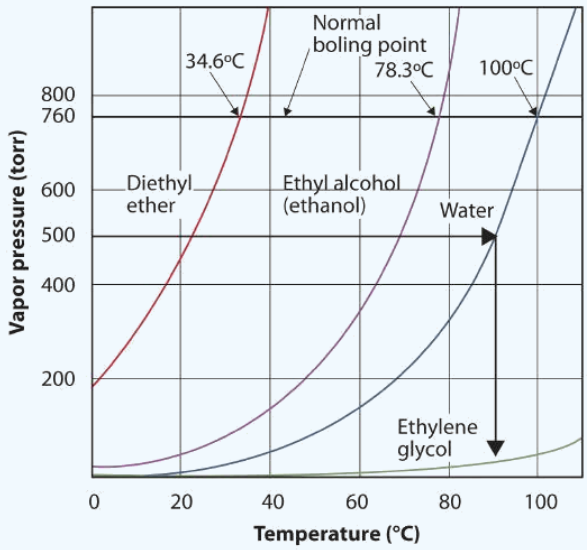
Figure 4.8.3: Plots of Vapor Pressure versus Temperature for Several Liquids. The vapor pressure of a liquid depends on the identity of the liquid and the temperature, as this plot shows.

Example 4.8.1

Use Figure 4.8.3 to estimate the boiling point of water at 500 torr, which is the approximate atmospheric pressure at the top of Mount Everest.

Solution

See the accompanying figure. Five hundred torr is between 400 and 600, so we extend a line from that point on the *y*-axis across to the curve for water and then drop it down to the *x*-axis to read the associated temperature. It looks like the point on the water vapor pressure curve corresponds to a temperature of about 90°C, so we conclude that the boiling point of water at 500 torr is 90°C.



By reading the graph properly, you can estimate the boiling point of a liquid at different temperatures.

Exercise 4.8.1

Use Figure 4.8.3 to estimate the boiling point of ethanol at 400 torr.

Answer available at the end of this section.

The vapor pressure curve for water is not exactly zero at the melting point 0°C. Even ice has a vapor pressure—that is why it sublimes over time. However, the vapor pressures of solids are typically much lower than that of liquids. At –1°C, the vapor pressure of ice is 4.2 torr. At a freezer temperature of 0°F (–17°C), the vapor pressure of ice is only 1.0 torr; so-called deep freezers can get down to –23°C, where the vapor pressure of ice is only 0.6 torr.

All liquids share some other properties as well. **Surface tension** is an effect caused by an imbalance of forces on the atoms at the surface of a liquid, as shown in Figure 4.8.4. The blue particle in the bulk of the liquid experiences intermolecular forces from all around, as illustrated by the arrows. However, the yellow particle on the surface does not experience any forces above it because there are no particles above it. This leads to an imbalance of forces, called surface tension.

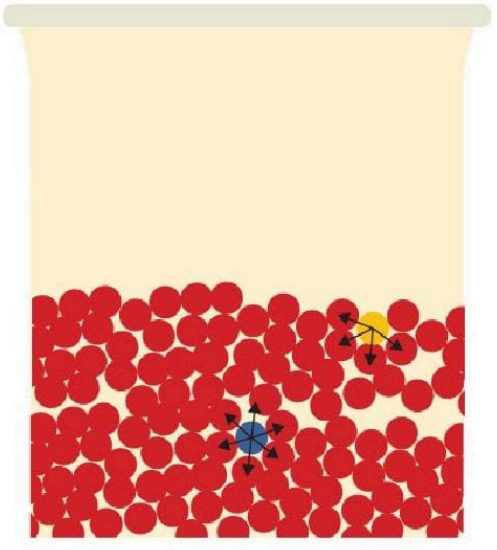


Figure 4.8.4 Surface Tension. Surface tension comes from the fact that particles at the surface of a liquid do not experience interactions from all directions, leading to an imbalance of forces on the surface.

Surface tension is responsible for several well-known behaviors of liquids, including water. Liquids with high surface tension tend to bead up when present in small amounts (Figure 4.8.5).



Figure 4.8.5 Effects of Surface Tension © Thinkstock. Water on the surface of this apple beads up due to the effect of surface tension.

Surface tension causes liquids to form spheres in free fall or zero gravity. Surface tension is also responsible for the fact that small insects can "walk" on water. Because of surface tension, it takes energy to break the surface of a liquid, and if an object (such as an insect) is light enough, there is not enough force due to gravity for the object to break through the surface, so the object stays on top of the water (Figure 4.8.6). Carefully done, this phenomenon can also be illustrated with a thin razor blade or a paper clip.



Figure 4.8.6 Walking on Water. Small insects can actually walk on top of water because of surface tension effects. (Unsplash License; Miguel Alcântara via unsplash)

The fact that small droplets of water bead up on surfaces does not mean that water—or any other liquid—does not interact with other substances. Sometimes the attraction can be very strong. **Adhesion** is the tendency of a substance to interact with other substances because of intermolecular forces, while **cohesion** is the tendency of a substance to interact with itself. If cohesive forces within a liquid are stronger than adhesive forces between a liquid and another substance, then the liquid tends to keep to itself; it will bead up. However, if adhesive forces between a liquid and another substance are stronger than cohesive forces, then the liquid will spread out over the other substance, trying to maximize the interface between the other substance and the liquid. It is said that the liquid *wets* the other substance.

Adhesion and cohesion are important for other phenomena as well. In particular, if adhesive forces are strong, then when a liquid is introduced to a small-diameter tube of another substance, the liquid will move up or down in the tube, as if ignoring gravity. Because tiny tubes are called capillaries, this phenomenon is called **capillary action.** For example, one type of capillary action, *capillary rise*, is seen when water or water-based liquids rise up in thin glass tubes (like the capillaries sometimes used in blood tests), forming an upwardly curved surface called a **meniscus**. Capillary action is also responsible for the "wicking" effect that towels and sponges use to dry wet objects; the matting of fibers forms tiny capillaries that have good adhesion with water. Cotton is a good material for this; polyester and other synthetic fabrics do not display similar capillary action, which is why you seldom find rayon bath towels. A similar effect is observed with liquid fuels or melted wax and their wicks. Capillary action is thought to be at least partially responsible for transporting water from the roots to the tops of trees, even tall ones.

On the other hand, some liquids have stronger cohesive forces than adhesive forces. In this case, in the presence of a capillary, the liquid is forced down from its surface; this is an example of a type of capillary action called *capillary depression*. In this case, the meniscus curves downward. Mercury has very strong cohesive forces; when a capillary is placed in a pool of mercury, the surface of the mercury liquid is depressed (Figure 4.8.7).



Figure 4.8.7: Capillary Action. (a) Capillary rise is seen when adhesion is strong, such as with water in a thin glass tube. (b) Capillary depression is seen when cohesive forces are stronger than adhesive forces, such as with mercury and thin glass tubes.

Chemistry is Everywhere: Waxing a Car

Responsible car owners are encouraged to wax their cars regularly. In addition to making the car look nicer, it also helps protect the surface, especially if the surface is metal. Why?

The answer has to do with cohesion and adhesion (and, to a lesser extent, rust). Water is an important factor in the rusting of iron, sometimes used extensively in outer car bodies. Keeping water away from the metal is one way to minimize rusting. A coat of paint helps with this. However, dirty or scratched paint can attract water, and adhesive forces will allow the water to wet the surface, maximizing its contact with the metal and promoting rust.



Figure 4.8.8: Waxing a car. Droplets of water on a freshly waxed car do not wet the car well because of low adhesion between water and the waxed surface. This helps protect the car from rust. (Unsplash License; Yuvraj Singh via Unspash)

Wax is composed of long hydrocarbon molecules that do not interact well with water. (Hydrocarbons are compounds with C and H atoms; for more information on hydrocarbons, see Chapter 16). That is, a thin layer of wax will not be wetted by water. A freshly waxed car has low adhesive forces with water, so water beads up on the surface, as a consequence of its cohesion and surface tension. This minimizes the contact between water and metal, thus minimizing rust.

Summary

- All liquids evaporate.
- If volume is limited, evaporation eventually reaches a dynamic equilibrium, and a constant vapor pressure is maintained.
- All liquids experience surface tension, an imbalance of forces at the surface of the liquid.
- All liquids experience capillary action, demonstrating either capillary rise or capillary depression in the presence of other substances.

Answers

Exercise 4.8.1

solute: HCl(g); solvent: CH_3OH

Exercise 4.8.2

H_2O because both experience hydrogen bonding

Exercise 4.8.3

octane (C_8H_{18}) will dissolve. It is also non-polar.

Exercise 4.8.1

about 65°C

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4.9: Impact of Solutes on Physical Properties

Learning Objective

- Name the four colligative properties.
- Calculate changes in vapor pressure, melting point, and boiling point of solutions.
- Calculate the osmotic pressure of solutions.

The properties of solutions are very similar to the properties of their respective pure solvents. This makes sense because the majority of the solution *is* the solvent. However, some of the properties of solutions differ from pure solvents in measurable and predictable ways. The differences are proportional to the fraction that the solute particles occupy in the solution. These properties are called **colligative properties**; the word *colligative* comes from the Greek word meaning "related to the number," implying that these properties are related to the number of solute particles, not their identities.

Before we introduce the first colligative property, we need to introduce a new concentration unit. The **mole fraction** of the ith component in a solution, χ_i, is the number of moles of that component divided by the total number of moles in the sample:

$$\chi_i = \frac{\text{moles of } i^{th}\text{component}}{\text{total moles}}$$

(χ is the lowercase Greek letter chi.) The mole fraction is always a number between 0 and 1 (inclusive) and has no units; it is just a number.

Example 4.9.1: Hydrocarbon Solution

A solution is made by mixing 12.0 g of $\mathrm{C_{10}H_8}$ in 45.0 g of $\mathrm{C_6H_6}$. What is the mole fraction of $\mathrm{C_{10}H_8}$ in the solution?

Solution

We need to determine the number of moles of each substance, add them together to get the total number of moles, and then divide to determine the mole fraction of $\mathrm{C_{10}H_8}$. The number of moles of $\mathrm{C_{10}H_8}$ is as follows:

$$12.0g\,C_{10}H_8 \times \frac{1\,mol\,C_{10}H_8}{128.10g\,C_{10}H_8} = 0.0936\,mol\,C_{10}H_8$$

The number of moles of $\mathrm{C_6H_6}$ is as follows:

$$45.0g\,C_6H_6 \times \frac{1\,mol\,C_6H_6}{78.12g\,C_6H_6} = 0.576\,mol\,C_6H_6$$

The total number of moles is

$$0.0936\ \text{mol} + 0.576\,\text{mol} = 0.670\ \text{mol}$$

Now we can calculate the mole fraction of $\mathrm{C_{10}H_8}$:

$$\chi_{C_{10}H_8} = \frac{0.0936mol}{0.670mol} = 0.140$$

The mole fraction is a number between 0 and 1 and is unit-less.

Exercise 4.9.1

A solution is made by mixing 33.8 g of $\mathrm{CH_3OH}$ in 50.0 g of $\mathrm{H_2O}$. What is the mole fraction of $\mathrm{CH_3OH}$ in the solution?

Answer available at the end of this section.

A useful thing to note is that the sum of the mole fractions of all substances in a mixture equals 1. Thus the mole fraction of $\mathrm{C_6H_6}$ in Example 4.9.1 could be calculated by evaluating the definition of mole fraction a second time, or—because there are only two substances in this particular mixture—we can subtract the mole fraction of the $\mathrm{C_{10}H_8}$ from 1 to get the mole fraction of $\mathrm{C_6H_6}$.

Now that this new concentration unit has been introduced, the first colligative property can be considered. As mentioned in Chapter 10, all pure liquids have a characteristic vapor pressure in equilibrium with the liquid phase, the partial pressure of which is dependent on temperature. Solutions, however, have a lower vapor pressure than the pure solvent has, and the amount of lowering is dependent on the fraction of solute particles, as long as the solute itself does not have a significant vapor pressure (the term *nonvolatile* is used to describe such solutes). This colligative property is called **vapor pressure depression** (or *lowering*). The actual vapor pressure of the solution can be calculated as follows:

$$P_{soln} = \chi_{solv} P^*_{solv} \tag{4.9.1}$$

where P_{soln} is the vapor pressure of the solution, χ_{solv} is the mole fraction of the solvent particles, and P^*_{solv} is the vapor pressure of the pure solvent at that temperature (which is data that must be provided). Equation 4.9.1 is known as **Raoult's law** (the approximate pronunciation is *rah-OOLT*). Vapor pressure depression is rationalized by presuming that solute particles take positions at the surface in place of solvent particles, so not as many solvent particles can evaporate.

Example 4.9.2: Vapor Pressure Reduction

A solution is made by mixing 12.0 g of $\mathrm{C_{10}H_8}$ in 45.0 g of $\mathrm{C_6H_6}$. If the vapor pressure of pure $\mathrm{C_6H_6}$ is 95.3 torr, what is the vapor pressure of the solution?

Solution

This is the same solution that was in Example 15, but here we need the mole fraction of $\mathrm{C_6H_6}$. The number of moles of \(\ce{C10H8}\ is as follows:

$$12.0g\,C_{10}H_8 \times \frac{1\,mol\,C_{10}H_8}{128.10g\,C_{10}H_8} = 0.0936\,mol\,C_{10}H_8$$

The number of moles of $\mathrm{C_6H_6}$ is as follows:

$$45.0g\,C_6H_6 \times \frac{1\,mol\,C_6H_6}{78.12g\,C_6H_6} = 0.576\,mol\,C_6H_6$$

So the total number of moles is

$$0.0936\,\mathrm{mol} + 0.576\,\mathrm{mol} = 0.670\ \ \mathrm{mol}$$

Now we can calculate the mole fraction of $\mathrm{C_6H_6}$:

$$\chi_{C_{10}H_8} = \frac{0.576mol}{0.670mol} = 0.860$$

(The mole fraction of $\mathrm{C_{10}H_8}$ calculated in Example 4.9.1 plus the mole fraction of $\mathrm{C_6H_6}$ equals 1, which is mathematically required by the definition of mole fraction.) Now we can use Raoult's law to determine the vapor pressure in equilibrium with the solution:

$$P_{soln} = (0.860)(95.3\,\mathrm{torr}) = 82.0\,\mathrm{torr}$$

The solution has a lower vapor pressure than the pure solvent.

Exercise 4.9.2

A solution is made by mixing 33.8 g of $C_6H_{12}O_6$ in 50.0 g of H_2O. If the vapor pressure of pure water is 25.7 torr, what is the vapor pressure of the solution?

Answer available at the end of this section.

Two colligative properties are related to solution concentration as expressed in molality. As a review, recall the definition of **molality**:

$$\text{molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}}$$

Because the vapor pressure of a solution with a nonvolatile solute is depressed compared to that of the pure solvent, it requires a higher temperature for the solution's vapor pressure to reach 1.00 atm (760 torr). Recall that this is the definition of the normal boiling point: the temperature at which the vapor pressure of the liquid equals 1.00 atm. As such, the normal boiling point of the solution is higher than that of the pure solvent. This property is called **boiling point elevation**.

The change in boiling point (ΔT_b) is easily calculated:

$$\Delta T_b = mK_b$$

where *m* is the molality of the solution and K_b is called the **boiling point elevation constant,** which is a characteristic of the solvent. Several boiling point elevation constants (as well as boiling point temperatures) are listed in Table 4.9.1.

Table 4.9.1: Boiling Point Data for Various Liquids

Liquid	Boiling Point (°C)	K_b (°C/*m*)
$HC_2H_3O_2$	117.90	3.07
C_6H_6	80.10	2.53
CCl_4	76.8	4.95
H_2O	100.00	0.512

Remember that what is initially calculated is the *change* in boiling point temperature, not the new boiling point temperature. Once the change in boiling point temperature is calculated, it must be added to the boiling point of the pure solvent—because boiling points are always elevated—to get the boiling point of the solution.

Example 4.9.3

What is the boiling point of a 2.50 *m* solution of $C_6H_4Cl_2$ in CCl_4? Assume that $C_6H_4Cl_2$ is not volatile.

Solution

Using the equation for the boiling point elevation,

$$\Delta T_b = (2.50\ m)(4.95°C/m) = 12.4°C$$

Note how the molality units have canceled. However, we are not finished. We have calculated the change in the boiling point temperature, not the final boiling point temperature. If the boiling point goes up by 12.4°C, we need to add this to the normal boiling point of CCl_4 to get the new boiling point of the solution:

$$T_{BP} = 76.8°C + 12.4°C = 89.2°C$$

The boiling point of the solution is predicted to be 89.2°C.

Exercise 4.9.3

What is the boiling point of a 6.95 *m* solution of $\mathrm{C_{12}H_{22}O_{11}}$ in $\mathrm{H_2O}$?

Answer available at the end of this section.

The boiling point of a solution is higher than the boiling point of the pure solvent, but the opposite occurs with the freezing point. The freezing point of a solution is lower than the freezing point of the pure solvent. Think of this by assuming that solute particles interfere with solvent particles coming together to make a solid, so it takes a lower temperature to get the solvent particles to solidify. This is called **freezing point depression**.

The equation to calculate the change in the freezing point for a solution is similar to the equation for the boiling point elevation:

$$\Delta T_f = mK_f$$

where *m* is the molality of the solution and K_f is called the **freezing point depression constant,** which is also a characteristic of the solvent only. Several freezing point depression constants (as well as freezing point temperatures) are listed in Table 4.9.2.

Table 4.9.2: Freezing Point Data for Various Liquids

Liquid	Freezing Point (°C)	K_f (°C/*m*)
$HC_2H_3O_2$	16.60	3.90
C_6H_6	5.51	4.90
C_6H_{12}	6.4	20.2
$C_{10}H_8$	80.2	6.8
H_2O	0.00	1.86

Remember that this equation calculates the change in the freezing point, not the new freezing point. What is calculated needs to be subtracted from the normal freezing point of the solvent, because freezing points always decrease.

Example 4.9.4: Freezing Point Reduction

What is the freezing point of a 1.77 *m* solution of $\mathrm{CBr_4}$ in $\mathrm{C_6H_6}$?

Solution

We use the equation to calculate the change in the freezing point and then subtract this number from the normal freezing point of C_6H_6 to get the freezing point of the solution:

$$\Delta T_f = (1.77\ m)(4.90°C/m) = 8.67°C$$

Now we subtract this number from the normal freezing point of C_6H_6, which is 5.51°C:

$$5.51 - 8.67 = -3.16°C$$

The freezing point of the solution is −3.16°C.

Exercise 4.9.4

What is the freezing point of a 3.05 *m* solution of CBr_4 in $C_{10}H_8$?

Answer available at the end of this section.

Freezing point depression is one colligative property that we use in everyday life. Many antifreezes used in automobile radiators use solutions that have a lower freezing point than normal so that automobile engines can operate at subfreezing temperatures. We also take advantage of freezing point depression when we sprinkle various compounds on ice to thaw it in the winter for safety (Figure 4.9.1). The compounds make solutions that have a lower freezing point, so rather than forming slippery ice, any ice is liquefied and runs off, leaving a safer pavement behind.



Figure 4.9.1: Workers spreading salt from a salt truck for deicing the road. Salt or other compounds take advantage of the freezing point depression to minimize the formation of ice on sidewalks and roads, thus increasing safety. (CC BY 2.0; Michael Pereckas via Wikipedia)

Before we introduce the final colligative property, we need to present a new concept. A **semipermeable membrane** is a thin membrane that will pass certain small molecules but not others. A thin sheet of cellophane, for example, acts as a semipermeable membrane. Consider the system in Figure 4.9.2.

a. A semipermeable membrane separates two solutions having the different concentrations marked. Curiously, this situation is not stable; there is a tendency for water molecules to move from the dilute side (on the left) to the concentrated side (on the right) until the concentrations are equalized, as in Figure 4.9.2.
b. This tendency is called **osmosis**. In osmosis, the solute remains in its original side of the system; only solvent molecules move through the semipermeable membrane. In the end, the two sides of the system will have different volumes. Because a column of liquid exerts a pressure, there is a pressure difference Π on the two sides of the system that is proportional to the height of the taller column. This pressure difference is called the **osmotic pressure**, which is a colligative property.

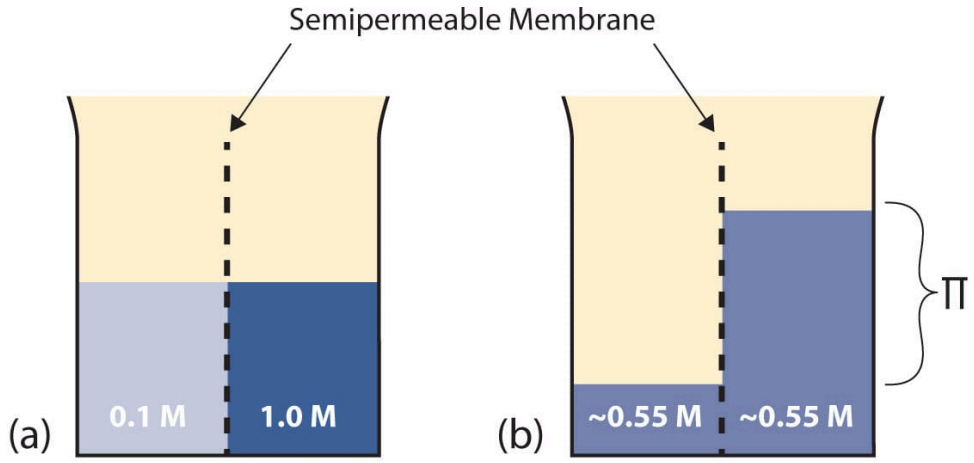


Figure 4.9.2: Osmosis. (a) Two solutions of differing concentrations are placed on either side of a semipermeable membrane. (b) When osmosis occurs, solvent molecules selectively pass through the membrane from the dilute solution to the concentrated solution, diluting it until the two concentrations are the same. The pressure exerted by the different height of the solution on the right is called the osmotic pressure.

The osmotic pressure of a solution is easy to calculate:

$$\Pi = MRT$$

where Π is the osmotic pressure of a solution, M is the molarity of the solution, R is the ideal gas law constant, and T is the absolute temperature. This equation is reminiscent of the ideal gas law discussed previously.

Example 4.9.5: Osmotic Pressure

What is the osmotic pressure of a 0.333 M solution of $\mathrm{C_6H_{12}O_6}$ at 25°C?

Solution

First we need to convert our temperature to kelvin:

$$\mathrm{T = 25 + 273 = 298\,K}$$

Now we can substitute into the equation for osmotic pressure, recalling the value for R:

$$\prod = (0.333M)\left(0.08205\frac{L.atm}{mol.K}\right)(298K)$$

The units may not make sense until we realize that molarity is defined as moles per liter:

$$\prod = \left(0.333\frac{mol}{L}\right)\left(0.08205\frac{L.atm}{mol.K}\right)(298K)$$

Now we see that the moles, liters, and kelvins cancel, leaving atmospheres, which is a unit of pressure. Solving,

$$\mathrm{\Pi = 8.14\,atm}$$

This is a substantial pressure! It is the equivalent of a column of water 84 m tall.

Exercise 4.9.5

What is the osmotic pressure of a 0.0522 M solution of $C_{12}H_{22}O_{11}$ at 55°C?

Answer available at the end of this section.

Osmotic pressure is important in biological systems because cell walls are semipermeable membranes. In particular, when a person is receiving intravenous (IV) fluids, the osmotic pressure of the fluid needs to be approximately the same as blood serum; otherwise there may be negative consequences. Figure 4.9.3 shows three red blood cells:

- a healthy red blood cell
- a red blood cell that has been exposed to a lower concentration than normal blood serum (a so-called *hypotonic* solution); the cell has plumped up as solvent moves into the cell to dilute the solutes inside.
- a red blood cell exposed to a higher concentration than normal blood serum (*hypertonic*); water leaves the red blood cell, so it collapses onto itself. Only when the solutions inside and outside the cell are the same (*isotonic*) will the red blood cell be able to do its job.

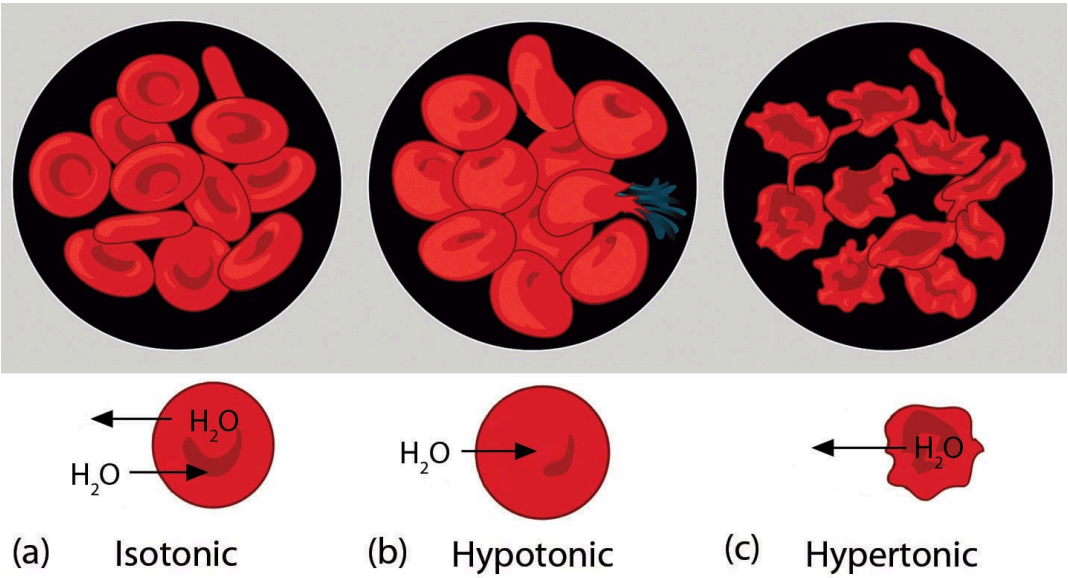


Figure 4.9.3: Osmotic Pressure and Red Blood Cells. (a) This is what a normal red blood cell looks like. (b) When a red blood cell is exposed to a hypotonic solution, solvent goes through the cell membrane and dilutes the inside of the cell. (c) When a red blood cell is exposed to a hypertonic solution, solvent goes from the cell to the surrounding solution, diluting the hypertonic solution and collapsing the cell. Neither of these last two cases is desirable, so IV solutions must be isotonic with blood serum to not cause deleterious effects. (CC BY-SA-NC; anonymous)

Osmotic pressure is also the reason you should not drink seawater if you're stranded in a lifeboat on an ocean; seawater has a higher osmotic pressure than most of the fluids in your body. You *can* drink the water, but ingesting it will pull water out of your cells as osmosis works to dilute the seawater. Ironically, your cells will die of thirst, and you will also die. (It is OK to drink the water if you are stranded on a body of freshwater, at least from an osmotic pressure perspective.) Osmotic pressure is also thought to be important in getting water to the tops of tall trees, in addition to capillary action.

Summary

- Colligative properties depend only on the number of dissolved particles (that is—the concentration), not their identity.
- Raoult's law is concerned with the vapor pressure depression of solutions.
- The boiling points of solutions are always higher, and the freezing points of solutions are always lower, than those of the pure solvent.
- Osmotic pressure is caused by concentration differences between solutions separated by a semipermeable membrane and is an important biological issue.

Learning Objective

- Determine the colligative properties of solutions of ionic solutes.

Previously, we considered the colligative properties of solutions with molecular solutes. What about solutions with ionic solutes? Do they exhibit colligative properties?

There is a complicating factor: ionic solutes separate into ions when they dissolve. This increases the total number of particles dissolved in solution and *increases the impact on the resulting colligative property*. Historically, this greater-than-expected impact on colligative properties was one main piece of evidence for ionic compounds separating into ions (increased electrical conductivity was another piece of evidence).

For example, when NaCl dissolves, it separates into two ions:

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

This means that a 1 M solution of NaCl actually has a net particle concentration of 2 M. The observed colligative property will then be twice as large as expected for a 1 M solution.

It is easy to incorporate this concept into our equations to calculate the respective colligative property. We define the **van 't Hoff factor** (i) as the number of particles each solute formula unit breaks apart into when it dissolves. Previously, we have always tacitly assumed that the van 't Hoff factor is simply 1. But for some ionic compounds, i is not 1, as shown in Table $4.9.1$.

Table $4.9.1$ *Ideal van 't Hoff Factors for Ionic Compounds*

Compound	*i*
NaCl	2
KBr	2
LiNO3	2
CaCl2	3
Mg(C2H3O2)2	3
FeCl3	4
Al2(SO4)3	5

The ideal van 't Hoff factor is equal to the number of ions that form when an ionic compound dissolves.

Example $4.9.1$

Predict the van 't Hoff factor for Sr(OH)2.

Solution

When Sr(OH)2 dissolves, it separates into one $\mathrm{Sr^{2+}}$ ion and two $\mathrm{OH^-}$ ions:

$$\mathrm{Sr(OH)_2 \rightarrow Sr^{2+}(aq) + 2\,OH^-(aq)}$$

Because it breaks up into three ions, its van 't Hoff factor is 3.

Exercise 4.9.1

What is the van 't Hoff factor for $Fe(NO_3)_3$?

Answer available at the end of this section.

It is the "ideal" van 't Hoff factor because this is what we expect from the ionic formula. However, this factor is usually correct only for dilute solutions (solutions less than 0.001 M). At concentrations greater than 0.001 M, there are enough interactions between ions of opposite charge that the net concentration of the ions is less than expected—sometimes significantly. The actual van 't Hoff factor is thus less than the ideal one. Here, we will use ideal van 't Hoff factors.

Revised equations to calculate the effect of ionization are then easily produced:

$$\Delta T_b = imK_b$$

$$\Delta T_f = imK_g$$

$$\Pi = iMRT$$

where all variables have been previously defined. To calculate vapor pressure depression according to Raoult's law, the mole fraction of solvent particles must be recalculated to take into account the increased number of particles formed on ionization.

Example 4.9.2

Determine the freezing point of a 1.77 *m* solution of NaCl in H_2O.

Solution

For NaCl, we need to remember to include the van 't Hoff factor, which is 2. Otherwise, the calculation of the freezing point is straightforward:

$$\Delta T_f = (2)(1.77\,m)(1.86\,°C/m) = 6.58°C$$

This represents the change in the freezing point, which is decreasing. So we have to subtract this change from the normal freezing point of water, 0.00°C:

0.00 – 6.58 = –6.58°C

Exercise 4.9.2

Determine the boiling point of a 0.887 m solution of $CaCl_2$ in H_2O.

Answer available at the end of this section.

Food and Drink App: Salting Pasta Cooking Water

When cooking dried pasta, many recipes call for salting the water before cooking the pasta. Some argue—with colligative properties on their side—that adding salt to the water raises the boiling point, thus cooking the pasta faster. Is there any truth to this?



Figure 4.9.1: Cooking dried pasta. Why do so many recipes call for adding salt to water when boiling pasta? Is it to raise the boiling temperature of the water? (Public Domain; Jan Vašek via pixabay)

To judge the veracity of this claim, we can calculate how much salt should be added to the water to raise the boiling temperature by 1.0°C, with the presumption that dried pasta cooks noticeably faster at 101°C than at 100°C (although a 1° difference may make only a negligible change in cooking times). We can calculate the molality that the water should have:

$$1.0°C = m(0.512°C/m)m = 1.95$$

We have ignored the van 't Hoff factor in our estimation because this obviously is not a dilute solution. Let us further assume that we are using 4 L of water (which is very close to 4 qt, which in turn equals 1 gal). Because 4 L of water is about 4 kg (it is actually slightly less at 100°C), we can determine how much salt (NaCl) to add:

$$4\ \cancel{kg\ H_2O} \times \frac{1.95\ \cancel{mol\ NaCl}}{\cancel{kg\ H_2O}} \times \frac{58.5g\ NaCl}{1\ \cancel{mol\ NaCl}} = 456.3g\ NaCl$$

This is just over 1 lb of salt and is equivalent to nearly 1 cup in the kitchen. In your experience, do you add almost a cup of salt to a pot of water to make pasta? Certainly not! A few pinches, perhaps one-fourth of a teaspoon, but not almost a cup! It is obvious that the little amount of salt that most people add to their pasta water is not going to significantly raise the boiling point of the water.

So why do people add some salt to boiling water? There are several possible reasons, the most obvious of which is taste: adding salt adds a little bit of salt flavor to the pasta. It cannot be much because most of the salt remains in the water, not in the cooked pasta. However, it may be enough to detect with our taste buds. The other obvious reason is habit; recipes tell us to add salt, so we do, even if there is little scientific or culinary reason to do so.

Summary

For ionic solutes, the calculation of colligative properties must include the fact that the solutes separate into multiple particles when they dissolve. The equations for calculating colligative properties of solutions of ionic solvents include the van 't Hoff factor, *i*.

Answers

Exercise 4.9.1

0.275

Exercise 4.9.2

24.1 torr

Exercise 4.9.3

103.6°C

Exercise 4.9.4

59.5°C

Exercise 4.9.5

1.40 atm

Exercise 4.9.1

4

Exercise 4.9.2

101.36 °C

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5: MEASUREMENTS AND CONVERSIONS IN CHEMISTRY (UNDER CONSTRUCTION)

5.1: Common Measurements in Chemistry

5.2: Systems of Units

5.3: Metric Conversions

5.4: Moles

5.5: Molar Mass

5.6: Mole Conversions

5.7: Chemical Formula Conversions

5.8: Concentrations

5.9: Concentration Conversions

5.10: Density Conversions

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5.1: Common Measurements in Chemistry

Mass and Weight

Mass is a measure of the amount of matter that an object contains. The mass of an object is made in comparison to the standard mass of 1 kilogram. The kilogram was originally defined as the mass of $1\ \mathrm{L}$ of liquid water at $4^{\circ}\mathrm{C}$ (volume of a liquid changes slightly with temperature). In the laboratory, mass is measured with a balance (see below), which must be calibrated with a standard mass so that its measurements are accurate.



Figure $5.1.2$ (From left to right:) Triple beam balance (source(opens in new window)), digital top-loading balance (source(opens in new window)), and an analytical balance (source).

Other common units of mass are the gram and the milligram. A gram is 1/1000th of a kilogram, meaning that there are $1000\ \mathrm{g}$ in $1\ \mathrm{kg}$. A milligram is 1/1000th of a gram, so there are $1000\ \mathrm{mg}$ in $1\ \mathrm{g}$.

Mass is often confused with the term weight. **Weight** is a measure of force that is equal to the gravitational pull on an object. The weight of an object is dependent on its location. On the moon, the force due to **gravity** is about one sixth that of the gravitational force on Earth. Therefore, a given object will weigh six times more on Earth than it does on the moon. Since mass is dependent only on the amount of matter present in an object, mass does not change with location. Weight measurements are often made with a spring scale, by reading the distance that a certain object pulls down and stretches a spring.

Summary

- Mass is a measure of the amount of matter that an object contains.
- Weight is a measure of force that is equal to the gravitational pull on an object.
- Mass is independent of location, while weight depends on location.

Contributors and Attributions

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Length and Volume

Length is the measurement of the extent of something along its greatest dimension. The SI basic unit of length, or linear measure, is the **meter** (m). All measurements of length may be made in meters, though the prefixes listed in various tables will often be more convenient. The width of a room may be expressed as about 5 meters (m), whereas a large distance, such as the distance between New York City and Chicago, is better expressed as 1150 kilometers (km). Very small distances can be expressed in units such as the millimeter or the micrometer. The width of a typical human hair is about 10 micrometers $(\mu\mathrm{m})$.

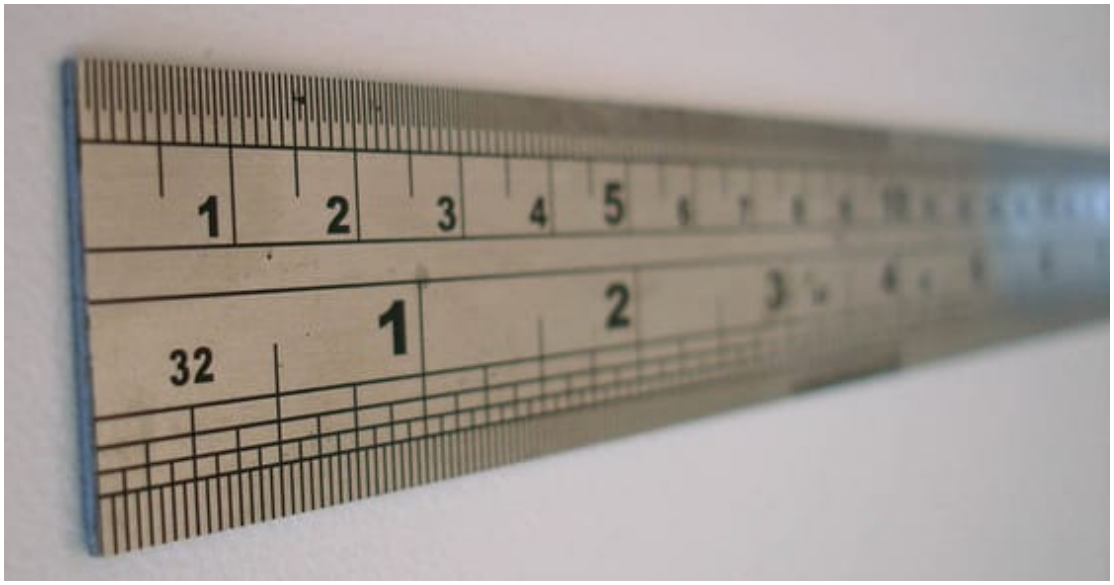


Figure 5.1.3: Metric and English measuring devices. (CC BY-SA 4.0; Ejay)

Volume is the amount of space occupied by a sample of matter. The volume of a regular object can be calculated by multiplying its length by its width and its height. Since each of those is a linear measurement, we say that units of volume are derived from units of length. The SI unit of volume is the cubic meter (m^3), which is the volume occupied by a cube that measures $1\ \mathrm{m}$ on each side. This very large volume is not very convenient for typical use in a chemistry laboratory. A liter (L) is the volume of a cube that measures $10\ \mathrm{cm}\ (1\ \mathrm{dm})$ on each side. A liter is thus equal to both $1000\ \mathrm{cm}^3\ (10\ \mathrm{cm} \times 10\ \mathrm{cm} \times 10\ \mathrm{cm})$ and to $1\ \mathrm{dm}^3$.

$$1\mathrm{L} = 1000\ \mathrm{cm}^3\ (10\ \mathrm{cm} \times 10\ \mathrm{cm} \times 10\ \mathrm{cm}) = 1\ \mathrm{dm}^3$$

A smaller unit of volume that is commonly used is the milliliter (mL). A milliliter is the volume of a cube that measures $1\ \mathrm{cm}$ on each side. Therefore, a milliliter is equal to a cubic centimeter (cm^3).

$$1\ \mathrm{mL}\ (\mathrm{cm}^3)$$

There are $1000\ \mathrm{mL}$ in $1\ \mathrm{L}$, which is the same as saying that there are $1000\ \mathrm{cm}^3$ in $1\ \mathrm{dm}^3$.



Figure 5.1.4 A typical water bottle is 1 liter in volume.

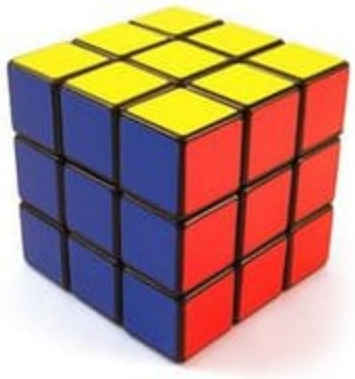


Figure 5.1.5 This Rubik's cube is $5.7\ \text{cm}$ on each side and has a volume of $185.2\ \text{cm}^3$ or $185.2\ \text{mL}$.

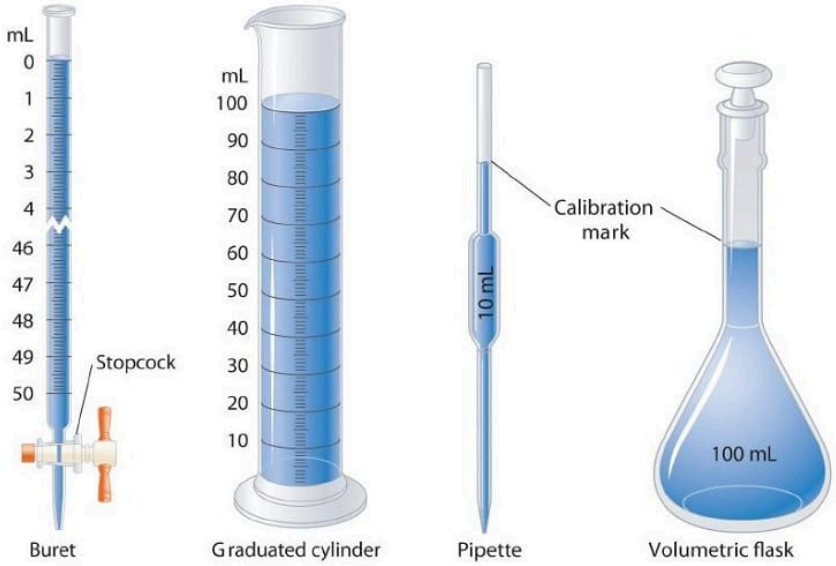


Figure 5.1.6: Buret & graduated cylinder: graduated glassware is used to deliver variable volumes of liquid. Pipette & volumetric flask: volumetric glassware is used to deliver (pipette) or contain (volumetric flask) a single volume accurately when filled to the mark. (CC BY-SA-NC; anonymous).

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

Summary

- Length is the measurement of the extent of something along its greatest dimension.
- Volume is the amount of space occupied by a sample of matter.
- Volume can be determined by knowing the length of each side of the item.

Contributors and Attributions

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Learning Objectives

- Define density.
- Determine the density of various substances.

Density (ρ) is a physical property found by dividing the mass of an object by its volume. Regardless of the sample size, density is always constant. For example, the density of a pure sample of tungsten is always 19.25 grams per cubic centimeter. This means that whether you have one gram or one kilogram of the sample, the density will never vary. The equation, as we already know, is as follows:

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

or just

$$\rho = \frac{m}{V}$$

Table 5.1.1 *Densities of Common Substances*

Substance	Density at 25°C (g/cm3)
blood	1.035
body fat	0.918
whole milk	1.030
corn oil	0.922
mayonnaise	0.910
honey	1.420

Notice in Table 5.1.1 that corn oil has a lower mass to volume ratio than water. This means that when added to water, corn oil will "float."

Density can be measured for all substances, solids, liquids and gases. For solids and liquids, density is often reported using the units of g/cm^3. Densities of gases, which are significantly lower than the densities of solids and liquids, are often given using units of g/L.

Example 5.1.1: Ethyl Alcohol

Calculate the density of a 30.2 mL sample of ethyl alcohol with a mass of 23.71002 g

Solution

$$\rho = \frac{23.71002\,g}{30.2\,mL} = 0.785\,g/mL$$

Exercise 5.1.1

a. Find the density (in kg/L) of a sample that has a volume of 36.5 L and a mass of 10.0 kg.
b. If you have a 2.130 mL sample of acetic acid with mass 2.234 g, what is the density in g/mL?

Answer a available at the end of this section.

Summary

- Density is defined as the mass of an object divided by its volume.

Contributors and Attributions

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Learning Objective

- Learn about the various temperature scales that are commonly used in chemistry.
- Define density and use it as a conversion factor.

There are other units in chemistry that are important, and we will cover others over the course of the entire book. One of the fundamental quantities in science is temperature. Temperature is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called degrees, and there are several temperature scales in use. In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as "degrees Fahrenheit"). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32°F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212°F.

Science also uses other scales to express temperature. The Celsius scale (symbolized by °C and spoken as "degrees Celsius") is a temperature scale where 0°C is the freezing point of water and 100°C is the boiling point of water; the scale is divided into 100 divisions between these two landmarks and extended higher and lower. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

$$^{\circ}\text{C} = (^{\circ}\text{F-32}) \times \tfrac{5}{9} \tag{5.1.1}$$

$$^{\circ}\text{F} = \left(^{\circ}\text{C} \times \tfrac{9}{5}\right) + 32 \tag{5.1.2}$$

Example 5.1.1: Conversions

a. What is 98.6 °F in degrees Celsius?
b. What is 25.0 °C in degrees Fahrenheit?

Solution

a. Using Equation 5.1.1, we have

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

b. Using Equation 5.1.2, we have

$$\begin{aligned} ^{\circ}F &= \left(25.0 \times \frac{9}{5}\right) + 32 \\ &= 45.0 + 32 \\ &= 77.0^{\circ}F \end{aligned}$$

Exercise 5.1.1

a. Convert 0 °F to degrees Celsius.
b. Convert 212 °C to degrees Fahrenheit.

Answer a available at the end of this section.

Answer b available at the end of this section.

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

$$K = {}^{\circ}C + 273.15$$

For most purposes, it is acceptable to use 273 instead of 273.15. Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as "two hundred ninety-five kelvins" and not "two hundred ninety-five degrees Kelvin."

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

Example 5.1.2: Room Temperature

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

Solution

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

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

Then we use the appropriate formula above to determine the temperature in the Kelvin scale:

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

So, room temperature is about 295 K.

Exercise 5.1.2

What is 98.6°F on the Kelvin scale?

Answer available at the end of this section.

Figure 5.1.1 compares the three temperature scales. Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. The other two countries are Liberia and Myanmar (formerly Burma). People driving near the borders of Canada or Mexico may pick up local radio stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.

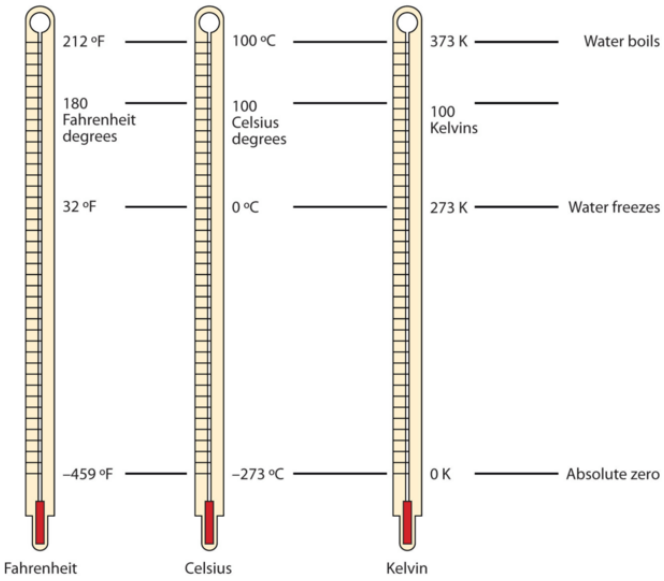


Figure 5.1.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales.

Density

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

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

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has derived units, just like velocity. Common units for density include g/mL, g/cm^3, g/L, kg/L, and even kg/m^3. Densities for some common substances are listed in Table 5.1.1.

Table 5.1.1: Densities of Some Common Substances

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

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (from Table 5.1.1), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

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

where we have limited our answer to two significant figures.

Example 5.1.3: Mercury

What is the mass of 44.6 mL of mercury?

Solution

Use the density from Table 5.1.1 "Densities of Some Common Substances" as a conversion factor to go from volume to mass:

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

The mass of the mercury is 607 g.

Exercise 5.1.3

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

Answer available at the end of this section.

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

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

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

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

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

$$\frac{1}{d} = \frac{1\,mL}{19.3\,g}$$

This reciprocal density is still a useful conversion factor, but now the mass unit will cancel and the volume unit will be introduced. Thus, if we want to know the volume of 45.9 g of gold, we would set up the conversion as follows:

$$45.9 \cancel{g} \times \frac{1\,mL}{19.3\,\cancel{g}} = 2.38\,mL$$

Note how the mass units cancel, leaving the volume unit, which is what we are looking for.

Example 5.1.4: Calculating Volume from Density

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/cm^3, what is the volume of the cork?

Solution

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

$$\frac{1}{d} = \frac{1\,cm^3}{0.22\,g}$$

We can use this expression as the conversion factor. So

$$3.78\,\cancel{g} \times \frac{1\,cm^3}{0.22\,\cancel{g}} = 17.2\,cm^3$$

Exercise 5.1.4

What is the volume of 3.78 g of gold?

Answer available at the end of this section.

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

Food and Drink Application: Cooking Temperatures

Because degrees Fahrenheit is the common temperature scale in the United States, kitchen appliances, such as ovens, are calibrated in that scale. A cool oven may be only 150°F, while a cake may be baked at 350°F and a chicken roasted at 400°F. The broil setting on many ovens is 500°F, which is typically the highest temperature setting on a household oven.

People who live at high altitudes, typically 2,000 ft above sea level or higher, are sometimes urged to use slightly different cooking instructions on some products, such as cakes and bread, because water boils at a lower temperature the higher in altitude you go, meaning that foods cook slower. For example, in Cleveland water typically boils at 212°F (100°C), but in Denver, the Mile-High City, water boils at about 200°F (93.3°C), which can significantly lengthen cooking times. Good cooks need to be aware of this.



A meat thermometer with a dial. Notice the markings for Fahrenheit (outer scale) and Celsius (inner scale) temperatures. Recipes for cooking food in an oven can use very different numbers, depending on the country you're in. (CC BY2.0 Bev Sykes)

At the other end is pressure cooking. A pressure cooker is a closed vessel that allows steam to build up additional pressure, which increases the temperature at which water boils. A good pressure cooker can get to temperatures as high as 252°F (122°C); at these temperatures, food cooks much faster than it normally would. Great care must be used with pressure cookers because of the high pressure and high temperature. (When a pressure cooker is used to sterilize medical instruments, it is called an *autoclave*.)

Other countries use the Celsius scale for everyday purposes. Therefore, oven dials in their kitchens are marked in degrees Celsius. It can be confusing for US cooks to use ovens abroad—a 425°F oven in the United States is equivalent to a 220°C oven in other countries. These days, many oven thermometers are marked with both temperature scales.

Key Takeaways

- Chemistry uses the Celsius and Kelvin scales to express temperatures.
- A temperature on the Kelvin scale is the Celsius temperature plus 273.15.
- The minimum possible temperature is absolute zero and is assigned 0 K on the Kelvin scale.
- Density relates the mass and volume of a substance.
- Density can be used to calculate volume from a given mass or mass from a given volume.

Learning Objectives

- Identify the difference between temperature and heat.
- Recognize the different scales used to measure temperature.
- Perform temperature conversions from one unit to another.

The concept of temperature may seem familiar to you, but many people confuse temperature with heat. **Temperature** is a measure of how hot or cold an object is relative to another object (its thermal energy content), whereas **heat** is the flow of thermal energy between objects with different temperatures. Temperature is a measure of the average kinetic energy of the particles in matter. In everyday usage, temperature indicates a measure of how hot or cold an object is. Temperature is an important parameter in chemistry. When a substance changes from solid to liquid, it is because there was in increase in the temperature of the material. Chemical reactions usually proceed faster if the temperature is increased. Many unstable materials (such as enzymes) will be viable longer at lower temperatures.



Figure 5.1.1: The glowing charcoal on the left represents high kinetic energy, while the snow and ice on the right are of much lower kinetic energy.

Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K). Thermometers measure temperature by using materials that expand or contract when heated or cooled. Mercury or alcohol thermometers, for example, have a reservoir of liquid that expands when heated and contracts when cooled, so the liquid column lengthens or shortens as the temperature of the liquid changes.



Figure 5.1.2 Daniel Gabriel Fahrenheit (top), Anders Celsius (center), and Lord Kelvin (bottom).

The Fahrenheit Scale

The first thermometers were glass and contained alcohol, which expanded and contracted as the temperature changed. The German scientist Daniel Gabriel Fahrenheit used mercury in the tube, an idea put forth by Ismael Boulliau. The Fahrenheit scale was first developed in 1724, and tinkered with for some time after that. The main problem with this scale is the arbitrary definitions of temperature. The freezing point of water was defined as $32^{\circ}\mathrm{F}$ and the boiling point as $212^{\circ}\mathrm{F}$. The Fahrenheit scale is typically not used for scientific purposes.

The Celsius Scale

The Celsius scale of the metric system is named after Swedish astronomer Anders Celsius (1701-1744). The Celsius scale sets the freezing point and boiling point of water at $0^{\circ}\mathrm{C}$ and $100^{\circ}\mathrm{C}$, respectively. The distance between those two points is divided into 100 equal intervals, each of which is one degree. Another term sometimes used for the Celsius scale is "centigrade" because there are 100 degrees between the freezing and boiling points of water on this scale. However, the preferred term is "Celsius".

The Kelvin Scale

The Kelvin temperature scale is named after Scottish physicist and mathematician Lord Kelvin (1824–1907). It is based on molecular motion, with the temperature of $0\ \mathrm{K}$, also known as absolute zero, being the point where all molecular motion ceases. The freezing point of water on the Kelvin scale is $273.15\ \mathrm{K}$, while the boiling point is $373.15\ \mathrm{K}$. Notice that there is no "degree" used in the temperature designation. Unlike the Fahrenheit and Celsius scales where temperatures are referred to as "degrees F" or "degrees C", we simply designated temperatures in the Kelvin scale as kelvins.

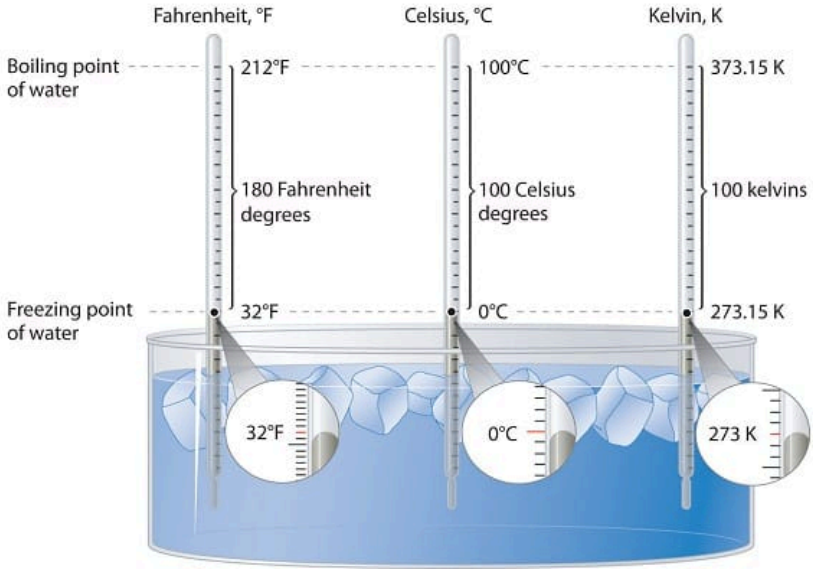


Figure 5.1.1: A Comparison of the Fahrenheit, Celsius, and Kelvin Temperature Scales. Because the difference between the freezing point of water and the boiling point of water is 100° on both the Celsius and Kelvin scales, the size of a degree Celsius (°C) and a kelvin (K) are precisely the same. In contrast, both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit (°F).

Converting between Scales

The Kelvin is the same size as the Celsius degree, so measurements are easily converted from one to the other. The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K. The Kelvin and Celsius scales are related as follows:

$$T\,(\text{in °C}) + 273.15 = T\,(\text{in K}) \tag{5.1.3}$$

$$T\ \ (\text{in K}) - 273.15 = T\ (\text{in °C}) \tag{5.1.4}$$

Degrees on the Fahrenheit scale, however, are based on an English tradition of using 12 divisions, just as 1 ft = 12 in. The relationship between degrees Fahrenheit and degrees Celsius is as follows:

$$°C = \frac{5}{9} \times (°F - 32) \tag{5.1.5}$$

$$°F = \frac{9}{5} \times (°C) + 32 \tag{5.1.6}$$

There is only one temperature for which the numerical value is the same on both the Fahrenheit and Celsius scales: –40°C = –40°F.

Example 5.1.1: Temperature Conversions

A student is ill with a temperature of 103.5°F. What is her temperature in °C and K?

Solution

Converting from Fahrenheit to Celsius requires the use of Equation 5.1.5:

$$\begin{aligned} °C &= \frac{5}{9} \times (103.5°F - 32) \\ &= 39.7\,°C \end{aligned}$$

Converting from Celsius to Kelvin requires the use of Equation 5.1.3:

$$\begin{aligned} K &= 39.7\,°C + 273.15 \\ &= 312.9\,K \end{aligned}$$

Exercise 5.1.1

Convert each temperature to °C and °F.

a. the temperature of the surface of the sun (5800 K)
b. the boiling point of gold (3080 K)
c. the boiling point of liquid nitrogen (77.36 K)

Answer (a) available at the end of this section.

Answer (b) available at the end of this section.

Answer (c) available at the end of this section.

Summary

- Three different scales are commonly used to measure temperature: Fahrenheit (expressed as °F), Celsius (°C), and Kelvin (K).
- The freezing point of water is 0°C = 273.15 K; the boiling point of water is 100°C = 373.15 K.
- Both a degree Celsius and a kelvin are 9/5 the size of a degree Fahrenheit.

Contributors and Attributions

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Learning Objective

- Learn about the various temperature scales that are commonly used in chemistry.
- Define density and use it as a conversion factor.

There are other units in chemistry that are important, and we will cover others over the course of the entire book. One of the fundamental quantities in science is temperature. Temperature is a measure of the average amount of energy of motion, or *kinetic energy*, a system contains. Temperatures are expressed using scales that use units called degrees, and there are several temperature scales in use. In the United States, the commonly used temperature scale is the *Fahrenheit scale* (symbolized by °F and spoken as "degrees Fahrenheit"). On this scale, the freezing point of liquid water (the temperature at which liquid water turns to solid ice) is 32°F, and the boiling point of water (the temperature at which liquid water turns to steam) is 212°F.

Science also uses other scales to express temperature. The Celsius scale (symbolized by °C and spoken as "degrees Celsius") is a temperature scale where 0°C is the freezing point of water and 100°C is the boiling point of water; the scale is divided into 100 divisions between these two landmarks and extended higher and lower. By comparing the Fahrenheit and Celsius scales, a conversion between the two scales can be determined:

Label 'eq1' multiply defined

Example 5.1.1: Conversions

a. What is 98.6 °F in degrees Celsius?
b. What is 25.0 °C in degrees Fahrenheit?

Solution

a. Using Equation 5.1.1, we have

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

b. Using Equation 5.1.2, we have

$$\begin{aligned} {}^{\circ}F &= \left(25.0 \times \frac{9}{5}\right) + 32 \\ &= 45.0 + 32 \\ &= 77.0^{\circ}F \end{aligned}$$

Exercise 5.1.1

a. Convert 0 °F to degrees Celsius.
b. Convert 212 °C to degrees Fahrenheit.

Answer a available at the end of this section.

Answer b available at the end of this section.

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

$$K = {}^{\circ}C + 273.15$$

For most purposes, it is acceptable to use 273 instead of 273.15. Note that the Kelvin scale does not use the word *degrees*; a temperature of 295 K is spoken of as "two hundred ninety-five kelvins" and not "two hundred ninety-five degrees Kelvin."

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

Example 5.1.2: Room Temperature

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

Solution

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

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

Then we use the appropriate formula above to determine the temperature in the Kelvin scale:

$$K = 22.2^{\circ}C + 273.15$$
$$= 295.4K$$

So, room temperature is about 295 K.

Exercise 5.1.2

What is 98.6°F on the Kelvin scale?

Answer available at the end of this section.

Figure 5.1.1 compares the three temperature scales. Note that science uses the Celsius and Kelvin scales almost exclusively; virtually no practicing chemist expresses laboratory-measured temperatures with the Fahrenheit scale. In fact, the United States is one of the few countries in the world that still uses the Fahrenheit scale on a daily basis. The other two countries are Liberia and Myanmar (formerly Burma). People driving near the borders of Canada or Mexico may pick up local radio stations on the other side of the border that express the daily weather in degrees Celsius, so do not get confused by their weather reports.

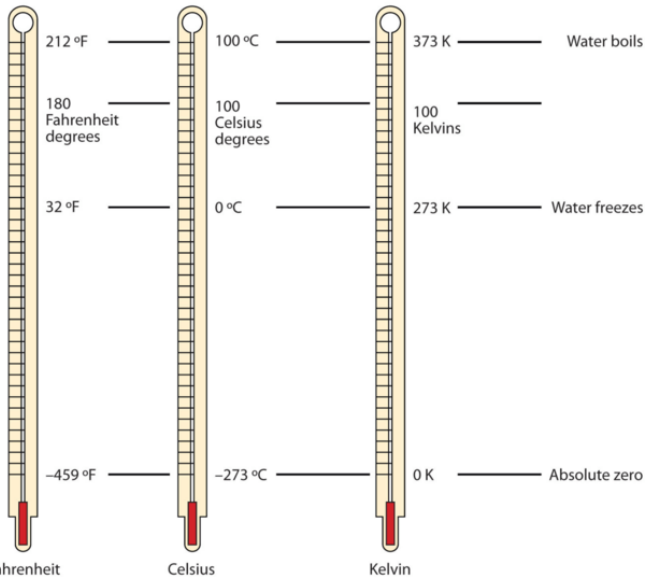


Figure 5.1.1: Fahrenheit, Celsius, and Kelvin Temperatures. A comparison of the three temperature scales.

Density

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

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

Density is usually a measured property of a substance, so its numerical value affects the significant figures in a calculation. Notice that density is defined in terms of two dissimilar units, mass and volume. That means that density overall has derived units, just like velocity. Common units for density include g/mL, g/cm^3, g/L, kg/L, and even kg/m^3. Densities for some common substances

are listed in Table 5.1.1.

Table 5.1.1: Densities of Some Common Substances

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

Because of how it is defined, density can act as a conversion factor for switching between units of mass and volume. For example, suppose you have a sample of aluminum that has a volume of 7.88 cm^3. How can you determine what mass of aluminum you have without measuring it? You can use the volume to calculate it. If you multiply the given volume by the known density (from Table 5.1.1), the volume units will cancel and leave you with mass units, telling you the mass of the sample:

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

where we have limited our answer to two significant figures.

Example 5.1.3: Mercury

What is the mass of 44.6 mL of mercury?

Solution

Use the density from Table 5.1.1 "Densities of Some Common Substances" as a conversion factor to go from volume to mass:

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

The mass of the mercury is 607 g.

Exercise 5.1.3

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

Answer available at the end of this section.

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

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

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

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

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

$$\frac{1}{d} = \frac{1\,mL}{19.3\,g}$$

This reciprocal density is still a useful conversion factor, but now the mass unit will cancel and the volume unit will be introduced. Thus, if we want to know the volume of 45.9 g of gold, we would set up the conversion as follows:

$$45.9\,\cancel{g} \times \frac{1\,mL}{19.3\,\cancel{g}} = 2.38\,mL$$

Note how the mass units cancel, leaving the volume unit, which is what we are looking for.

Example 5.1.4: Calculating Volume from Density

A cork stopper from a bottle of wine has a mass of 3.78 g. If the density of cork is 0.22 g/cm^3, what is the volume of the cork?

Solution

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

$$\frac{1}{d} = \frac{1\,cm^3}{0.22\,g}$$

We can use this expression as the conversion factor. So

$$3.78\,\cancel{g} \times \frac{1\,cm^3}{0.22\,\cancel{g}} = 17.2\,cm^3$$

Exercise 5.1.4

What is the volume of 3.78 g of gold?

Answer available at the end of this section.

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

Food and Drink Application: Cooking Temperatures

Because degrees Fahrenheit is the common temperature scale in the United States, kitchen appliances, such as ovens, are calibrated in that scale. A cool oven may be only 150°F, while a cake may be baked at 350°F and a chicken roasted at 400°F. The broil setting on many ovens is 500°F, which is typically the highest temperature setting on a household oven.

People who live at high altitudes, typically 2,000 ft above sea level or higher, are sometimes urged to use slightly different cooking instructions on some products, such as cakes and bread, because water boils at a lower temperature the higher in altitude you go, meaning that foods cook slower. For example, in Cleveland water typically boils at 212°F (100°C), but in Denver, the Mile-High City, water boils at about 200°F (93.3°C), which can significantly lengthen cooking times. Good cooks need to be aware of this.



A meat thermometer with a dial. Notice the markings for Fahrenheit (outer scale) and Celsius (inner scale) temperatures. Recipes for cooking food in an oven can use very different numbers, depending on the country you're in. (CC BY2.0 Bev Sykes)

At the other end is pressure cooking. A pressure cooker is a closed vessel that allows steam to build up additional pressure, which increases the temperature at which water boils. A good pressure cooker can get to temperatures as high as 252°F (122°C); at these temperatures, food cooks much faster than it normally would. Great care must be used with pressure cookers because of the high pressure and high temperature. (When a pressure cooker is used to sterilize medical instruments, it is called an *autoclave*.)

Other countries use the Celsius scale for everyday purposes. Therefore, oven dials in their kitchens are marked in degrees Celsius. It can be confusing for US cooks to use ovens abroad—a 425°F oven in the United States is equivalent to a 220°C oven in other countries. These days, many oven thermometers are marked with both temperature scales.

Key Takeaways

- Chemistry uses the Celsius and Kelvin scales to express temperatures.
- A temperature on the Kelvin scale is the Celsius temperature plus 273.15.
- The minimum possible temperature is absolute zero and is assigned 0 K on the Kelvin scale.
- Density relates the mass and volume of a substance.
- Density can be used to calculate volume from a given mass or mass from a given volume.

Learning Objectives

- Define *pressure*.
- Learn the units of pressure and how to convert between them.

The kinetic theory of gases indicates that gas particles are always in motion and are colliding with other particles and the walls of the container holding them. Although collisions with container walls are elastic (i.e., there is no net energy gain or loss because of the collision), a gas particle does exert a force on the wall during the collision. The accumulation of all these forces, distributed over the area of the walls of the container, causes pressure. **Pressure** (P) is defined as the force of all the gas particle/wall collisions divided by the area of the wall:

$$\text{pressure} = \frac{\text{force}}{\text{area}}$$

All gases exert pressure; it is one of the fundamental measurable quantities of this phase of matter. Even our atmosphere exerts pressure—in this case, the gas is being "held in" by the earth's gravity, rather than the gas being in a container. The pressure of the atmosphere is about 14.7 pounds of force for every square inch of surface area: 14.7 lb/in^2.

Pressure has a variety of units. The formal, SI-approved unit of pressure is the *pascal* (Pa), which is defined as 1 N/m^2 (one newton of force over an area of one square meter). However, this is usually too small in magnitude to be useful. A common unit of pressure is the **atmosphere** (atm), which was originally defined as the average atmospheric pressure at sea level.

However, "average atmospheric pressure at sea level" is difficult to pinpoint because of atmospheric pressure variations. A more reliable and common unit is **millimeters of mercury** (mmHg), which is the amount of pressure exerted by a column of mercury exactly 1 mm high. An equivalent unit is the **torr**, which equals 1 mmHg. (The torr is named after Evangelista Torricelli, a seventeenth-century Italian scientist who invented the mercury barometer.) With these definitions of pressure, the atmosphere unit is redefined: 1 atm is defined as exactly 760 mmHg, or 760 torr. We thus have the following equivalencies:

$$1\text{ atm} = 760\text{ mmHg} = 760\text{ torr}$$

We can use these equivalencies as with any equivalencies—to perform conversions from one unit to another. Relating these to the formal SI unit of pressure, 1 atm = 101,325 Pa.

Example 5.1.1: Pressure Conversion

How many atmospheres are there in 595 torr?

Solution

Using the pressure equivalencies, we construct a conversion factor between torr and atmospheres:

$$\frac{1\,atm}{760\,torr}$$

$$595\ \cancel{torr} \times \frac{1\,atm}{760\ \cancel{torr}} = 0.783\,atm$$

Because the numbers in the conversion factor are exact, the number of significant figures in the final answer is determined by the initial value of pressure.

Exercise 5.1.1

How many atmospheres are there in 1,022 torr?

Answer available at the end of this section.

Example 5.1.2: Mars

The atmosphere on Mars is largely CO_2 at a pressure of 6.01 mmHg. What is this pressure in atmospheres?

Solution

Use the pressure equivalencies to construct the proper conversion factor between millimeters of mercury and atmospheres.

$$6.01\ \cancel{mmHg} \times \frac{1\,atm}{760\ \cancel{mmHg}} = 0.00791\,atm = 7.91 \times 10^{-3} atm$$

At the end, we expressed the answer in scientific notation.

Exercise 5.1.2

Atmospheric pressure is low in the eye of a hurricane. In a 1979 hurricane in the Pacific Ocean, a pressure of 0.859 atm was reported inside the eye. What is this pressure in torr?

Answer available at the end of this section.

Summary

- Pressure is a force exerted over an area.
- Pressure has several common units that can be converted.

Learning Objectives

- Define *energy*.
- Know the units of energy.
- Understand the law of conservation of energy.

Energy is the ability to do work. Think about it: when you have a lot of energy, you can do a lot of work; but if you're low on energy, you do not want to do much work. Work (w) itself is defined as a force (F) operating over a distance (Δx):

$$w = F \times \Delta x$$

In SI, force has units of newtons (N), while distance has units of meters. Therefore, work has units of N·m. This compound unit is redefined as a **Joule (J)**.

$$1 \text{ joule} = 1 \text{ newton·meter} \quad 1 \text{ J} = 1 \text{ N·m}$$

Because energy is the ability to do work, energy is also measured in joules. This is the primary unit of energy we will use here.

How much is 1 J? It is enough to warm up about one-fourth of a gram of water by 1°C. It takes about 12,000 J to warm a cup of coffee from room temperature to 50°C. So a joule is not a lot of energy. It will not be uncommon to measure energies in thousands of joules, so the kilojoule (kJ) is a common unit of energy, with 1 kJ equal to 1,000 J.

An older—but still common—unit of energy is the *calorie*. The calorie (cal) was originally defined in terms of warming up a given quantity of water. The modern definition of calorie equates it to joules:

$$1 \text{ cal} = 4.184 \text{ J}$$

The calorie is used when considering nutrition. Energy contents of foods are often expressed in calories. However, the calorie unit used for foods is actually the kilocalorie (kcal). Most foods indicate this by spelling the word with a capital C—Calorie. Figure 5.1.1 - Calories on Food Labels, shows one example.



Nutrition Facts

Serving Size 3 oz. (85g)

Amount Per Serving	As Served
Calories 38	**Calories from Fat** 0
	% Daily Value
Total Fat 0g	0%
Saturated Fat 0g	0%
Cholesterol 0g	0%
Sodium 0g	2%
Total Carbohydrate 0g	3%
Dietary Fiber 0g	8%
Sugars 0g	
Protein 0g	

Vitamin A 270%	•	**Vitamin C** 10%	
Calcium 2%	•	**Iron** 0%	

Percent Daily Values are based on a 2,000 calorie diet. Your daily values may be higher or lower depending on your calorie needs:

	Calories	2,000	2,500
Total Fat	Less than	65g	80g
Sat Fat	Less than	20g	80g
Cholesterol	Less than	300mg	300mg
Sodium	Less than	2,400mg	2,400mg
Total Carbohydrate		300g	375g
Dietary Fiber		25g	30g

Figure 5.1.1 Calories on Food Labels © Thinkstock. *This label expresses the energy content of the food, but in Calories (which are actually kilocalories).*

Example 5.1.1

The label in Figure 5.1.1 states that the serving has 38 Cal. How many joules is this?

Solution

We recognize that with a capital C, the Calories unit is actually kilocalories. To determine the number of joules, we convert first from kilocalories to calories (using the definition of the *kilo-* prefix) and then from calories to joules (using the relationship between calories and joules). So

$$38\ \cancel{kcal} \times \frac{1000\ \cancel{cal}}{1\ \cancel{kcal}} \times \frac{4.184\ J}{1\ \cancel{cal}} = 160,000\ J$$

Exercise 5.1.1

A serving of breakfast cereal usually has 110 Cal. How many joules of energy is this?

Answer available at the end of this section.

In the study of energy, we use the term **system** to describe the part of the universe under study: a beaker, a flask, or a container whose contents are being observed and measured. An **isolated system** is a system that does not allow a transfer of energy or matter into or out of the system. A good approximation of an isolated system is a closed, insulated thermos-type bottle. The fact that the thermos-type bottle is closed keeps matter from moving in or out, and the fact that it is insulated keeps energy from moving in or out.

One of the fundamental ideas about the total energy of an isolated system is that it does not increase or decrease. When this happens to a quantity, we say that the quantity is *conserved*. The **law of conservation of energy** states that the total energy of an isolated system does not change. As a scientific law, this concept occupies the highest level of understanding we have about the natural universe.

Key Takeaways

- Energy is the ability to do work and uses the unit joule.

- The law of conservation of energy states that the total energy of an isolated system does not increase or decrease.

Answers

Exercise 5.1.1

$0.274\ kg/L$

$1.049\ g/mL$

Exercise 5.1.1

–17.8 °C

Exercise 5.1.1

414 °F

Exercise 5.1.2

310.2 K

Exercise 5.1.3

197 g

Exercise 5.1.4

0.196 cm^3

Exercise 5.1.1

5527 °C, 9980 °F

Exercise 5.1.1

2807 °C, 5084 °F

Exercise 5.1.1

-195.8 °C, -320.4 °F

Exercise 5.1.1

–17.8 °C

Exercise 5.1.1

414 °F

Exercise 5.1.2

310.2 K

Exercise 5.1.3

197 g

Exercise 5.1.4

0.196 cm^3

Exercise 5.1.1

1.345 atm

Exercise 5.1.2

652 torr

Exercise 5.1.1

460,000 J

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5.2: Systems of Units

Learning Objectives

- Express quantities properly, using a number and a unit.
- State the different measurement systems used in chemistry.
- Express a large number or a small number in scientific notation.
- Learn how to use SI prefixes.
- Perform unit conversions using conversion factors.

A coffee maker's instructions tell you to fill the coffeepot with 4 cups of water and use 3 scoops of coffee. When you follow these instructions, you are measuring. When you visit a doctor's office, a nurse checks your temperature, height, weight, and perhaps blood pressure (Figure 5.2.1); the nurse is also measuring.



Figure 5.2.1: A nurse or a doctor measuring a patient's blood pressure is taking a measurement. Figure used with permission (GFDL; Pia von Lützau).

Chemists measure the properties of matter and express these measurements as quantities. A quantity is an amount of something, and consists of a number and a unit. The number tells us how many (or how much), and the unit tells us what the scale of measurement is. For example, when a distance is reported as "5 kilometers," we know that the quantity has been expressed in units of kilometers and that the number of kilometers is 5. If you ask a friend how far he or she walks from home to school, and the friend answers "12" without specifying a unit, you do not know whether your friend walks 12 miles, 12 kilometers, 12 furlongs, or 12 yards...etc. *Both a number and a unit must be included to express a quantity properly.*

To understand chemistry, we need a clear understanding of the units chemists work with and the rules they follow for expressing numbers. The next two sections examine the rules for expressing numbers.

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 the seven base units shown in Table 5.2.1.

Table 5.2.1 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.

Exponential Numbers: Powers of Ten

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. Furthermore, use of prefixes is another way to express measurements involving large and small numbers.

Scientific Notation

In scientific notation, numbers are expressed in the form

$$N \times 10^n$$

where N is greater than or equal to 1 and less than 10 ($1 \le N < 10$), 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 5.2.1.

Example 5.2.1: Expressing Numbers in Scientific Notation

Convert each number to scientific notation.

a. 637.8
b. 0.0479
c. 12,378
d. 0.00032

Solution

Example 5.2.1 *Solution and Explanation*

	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	Because the decimal point was moved **four places to the left, n = 4**.	1.2378×10^{4}
d	Because the decimal point was moved **four places to the right, n = –4**.	3.2×10^{-4}

Exercise 5.2.1

Convert each ordinary number to scientific notation, or vice versa.

a. 67,000,000,000
b. 1,689
c. 12.6

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 5.2.2

Convert each ordinary number to scientific notation, or vice versa.

a. 0.000006567
b. 6.22×10^{-2}
c. 9.9×10^{-9}

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Metric Prefixes

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 $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great". Table 5.2.2 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 5.2.2 SI Prefixes

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

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, 3.5 milliliters is written as $3.5\ \mathrm{mL}$.

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 $6.5\ \mathrm{kg}$, this is easier than saying it weighs $6500\ \mathrm{g}$ or $0.65\ \mathrm{dag}$. 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 5.2.2: Unit Abbreviations

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

a. kiloliter
b. microsecond
c. decimeter
d. nanogram

Solutions

Example 5.2.2 Solution and Explanation

	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 5.2.3

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

a. kilometer
b. milligram
c. nanosecond
d. centiliter

Answer a: available at the end of this section.

Answer b: available at the end of this section.

Answer c: available at the end of this section.

Answer d: available at the end of this section.

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.
- Mass is a measure of the amount of matter that an object contains.
- Weight is a measure of force that is equal to the gravitational pull on an object.
- Mass is independent of location, while weight depends on location.
- Length is the measurement of the extent of something along its greatest dimension.
- Volume is the amount of space occupied by a sample of matter.
- Volume can be determined by knowing the length of each side of the item.
- Conversion factors are used to convert one unit of measurement into another.
- Dimensional analysis (unit conversions) involves the use of conversion factors that will cancel unwanted units and produce desired units.

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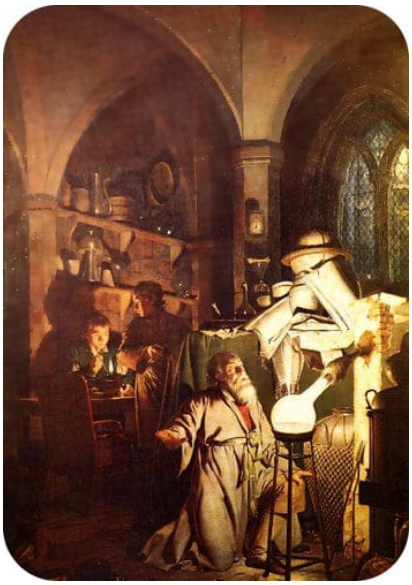


Figure 5.2.1 (Public Domain; Joseph Wright of Derby via Wikipedia)

What do Latin and Greek have to do with modern science?

Isn't it hard enough to learn English terms? For hundreds of years, the languages of the educated class were Latin and Greek. In part, because the literature of philosophy was Latin and Greek. Even the medieval Bibles were written in those two languages – the first English translation was in the late 1380s. Using Latin and Greek allowed scholars from different countries to communicate more easily with one another. Today we still see many Latin phrases in legal communications ("pro bono" meaning to do something "for the good" and not charge legal fees), scientific naming of biological species, and Latin is used for the annual student speech at Harvard University graduations. Not bad for a "dead" language.

Metric Prefixes

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 $\mu\varepsilon\gamma\alpha\varsigma$, meaning "great".

Table 5.2.1 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 5.2.1: SI Prefixes

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kilo	k	1,000	1 kilometer $(\mathrm{km}) = 1,000\ \mathrm{m}$
hecto	h	100	1 hectometer $(\mathrm{hm}) = 100\ \mathrm{m}$
deka	da	10	1 dekameter $(\mathrm{dam}) = 10\ \mathrm{m}$
		1	**1 meter** (m)
deci	d	1/10	1 decimeter $(\mathrm{dm}) = 0.1\ \mathrm{m}$
centi	c	1/100	1 centimeter $(\mathrm{cm}) = 0.01\ \mathrm{m}$
milli	m	1/1,000	1 millimeter $(\mathrm{mm}) = 0.001\ \mathrm{m}$

Prefix	Unit Abbreviation	Meaning	Example
micro	μ	1/1,000,000	1 micrometer $(\mu\text{m}) = 10^{-6}\ \text{m}$
nano	n	1/1,000,000,000	1 nanometer $(\text{nm}) = 10^{-9}\ \text{m}$
pico	p	1/1,000,000,000,000	1 picometer $(\text{pm}) = 10^{-12}\ \text{m}$

There are more prefixes - some of them rarely used. Have you ever heard of a zeptometer?

There are a couple of odd little practices with the use of metric abbreviations. Most abbreviations are lowercase. The lowercase "m" is used for meter, instead of "M". However, when it comes to volume, the base unit "liter" is abbreviated as "L" and not "l". So, 3.5 milliliters is written as $3.5\ \text{mL}$.

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 $6.5\ \text{kg}$, this is easier than saying that it weighs $6500\ \text{g}$ or $0.65\ \text{dag}$. 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.

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Simulations

Practice using metric prefixes with this simulation.

Summary

- Metric prefixes derive from Latin or Greek terms.
- Metric prefixes are used to make the units manageable.

Review

1. What is the prefix for “thousand”?
2. What is the prefix for 0.01?
3. How would you write 500 milliliters?
4. How many decimeters in one meter?
5. You have a mass that weighs 1.2 hectograms. How many grams does it weigh?

Answers

Exercise 5.2.1

6.7×10^{10}

Exercise 5.2.1

1.689×10^{3}

Exercise 5.2.1

1.26×10^{1}

Exercise 5.2.2

6.567×10^{-6}

Exercise 5.2.2

0.0622

Exercise 5.2.2

0.0000000099

Exercise 5.2.3

km

Exercise 5.2.3

mg

Exercise 5.2.3

ns

Exercise 5.2.3

cL

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5.3: Metric Conversions

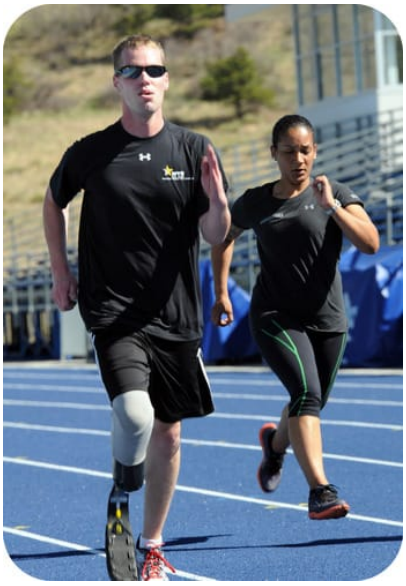


Figure 5.3.1 (Public Domain; the US Army via Wikipedia)

How can a number of track laps be converted to a distance in meters?

You are training for a 10-kilometer run by doing laps on a 400-meter track. You ask yourself how many times you need to run around the track in order to cover ten kilometers (...more than you realize). By using dimensional analysis, you can easily determine the number of laps needed to cover the $10\ \text{km}$ distance.

Metric Unit Conversions

The metric system's many prefixes allow quantities to be expressed in many different units. Dimensional analysis is useful to convert from one metric system unit to another.

Example 5.3.1

A particular experiment for science class requires $120\ \text{mL}$ of a solution. The science teacher knows that he will need to make enough solution for 40 experiments to be performed throughout the day. How many liters of solution should he prepare?

Solution

Step 1: List the known quantities and plan the problem.

Known

- 1 experiment requires $120\ \text{mL}$
- $1\ \text{L} = 1000\ \text{mL}$

Unknown

Since each experiment requires $120\ \text{mL}$ of solution and the teacher needs to prepare enough for 40 experiments, multiply 120 by 40 to get $4800\ \text{mL}$ of solution needed. Now you must convert mL to L by using a conversion factor.

Step 2: Calculate.

$$4800\ \text{mL} \times \frac{1\ \text{L}}{1000\ \text{mL}} = 4.8\ \text{L}$$

Note that the conversion factor is arranged so that the mL unit is in the denominator and thus cancels out, leaving L as the remaining unit in the answer.

Step 3:Think about your result.

A liter is much larger than a milliliter, so it makes sense that the number of liters required is less than the number of milliliters.

Two-Step Metric Unit Conversions

Some metric conversion problems are most easily solved by breaking them down into more than one step. When both the given unit and the desired unit have prefixes, one can first convert to the simple (un-prefixed) unit, followed by a conversion to the desired unit. The following example will illustrate this method.

Example 5.3.2

Convert $4.3\ \text{cm}$ to μm.

Solution

Step 1: List the known quantities and plan the problem.

Known

- $1\ \text{m} = 100\ \text{cm}$
- $1\ \text{m} = 10^6\ \mu\text{m}$

Unknown

You may need to consult a table for the multiplication factor represented by each metric prefix. First convert cm to m, then convert m to μm.

Step 2: Calculate.

$$4.3\ \text{cm} \times \frac{1\ \text{m}}{100\ \text{cm}} \times \frac{10^6\ \mu\text{m}}{1\ \text{m}} = 43,000\ \mu\text{m}$$

Each conversion factor is written so that the unit of the denominator cancels with the unit of the numerator of the previous factor.

Step 3: Think about your result.

A micrometer is a smaller unit of length than a centimeter, so the answer in micrometers is larger than the number of centimeters given.

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Summary

- Dimensional analysis can be used to carry out metric unit conversions.

Review

1. Perform the following conversions.
   a. 0.074 km to m
   b. 24,600 µg to g
   c. 4.9×10^{7} µg to kg
   d. 84 dm to mm

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5.4: Moles

The Mole and Avogadro's Number

It certainly is easy to count bananas or to count elephants (as long as you stay out of their way). However, you would be counting grains of sugar from your sugar canister for a long, long time. Atoms and molecules are extremely small—far, far smaller than a grain of sugar. Counting atoms or molecules is not only unwise, it is absolutely impossible. One drop of water contains about 10^{22} molecules of water. If you counted 10 molecules every second for 50 years without stopping, you would have counted only 1.6×10^{10} molecules. Put another way, at that counting rate, it would take you over 30 trillion years to count the water molecules in one tiny drop.

Chemists of the past needed a name that stood for a very large number of items. Amadeo Avogadro (1776-1856), an Italian scientist, provided such a number. He is responsible for the counting unit of measure called the mole. A **mole** (mol) is the amount of a substance that contains 6.02×10^{23} representative particles of that substance. The mole is the SI unit for amount of a substance. Just like the dozen and the gross, it is a name that stands for a number. There are therefore 6.02×10^{23} water molecules in a mole of water molecules. There also would be 6.02×10^{23} bananas in a mole of bananas—if such a huge number of bananas ever existed.



Figure 5.4.1: Italian scientist Amadeo Avogadro, whose work led to the concept of the mole as a counting unit in chemistry.

The number 6.02×10^{23} is called **Avogadro's number**, the number of representative particles in a mole. It is an experimentally determined number. A **representative particle** is the smallest unit in which a substance naturally exists. For the majority of elements, the representative particle is the atom. Iron, carbon, and helium consist of iron atoms, carbon atoms, and helium atoms, respectively. Seven elements exist in nature as diatomic molecules and they are $\mathrm{H_2}$, $\mathrm{N_2}$, $\mathrm{O_2}$, $\mathrm{F_2}$, $\mathrm{Cl_2}$, $\mathrm{Br_2}$, and $\mathrm{I_2}$. The representative particle for these elements is the molecule. Likewise, all molecular compounds such as $\mathrm{H_2O}$ and $\mathrm{CO_2}$ exist as molecules, and so the molecule is their representative particle. For ionic compounds such as NaCl and $\mathrm{Ca(NO_3)_2}$, the representative particle is the formula unit. A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.



Figure 5.4.2: The animal mole is very different than the counting unit of the mole. Chemists nonetheless have adopted the mole as their unofficial mascot. National Mole Day is a celebration of chemistry that occurs on October 23 (10/23) of each year.

Summary

- A mole of any substance contains Avogadro's number $\left(6.02 \times 10^{23}\right)$ of representative particles.

Contributors and Attributions

-

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Learning Objectives

- Use Avogadro's number to convert to moles and vice versa given the number of particles of an element.
- Know the definition of the mole.
- Determine the formula mass of an ionic or molecular compound.
- Determine the percent composition of each element in a compound from the chemical formula.

When objects are very small, it is often inconvenient or inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross".

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number:Counting Atoms

Owing to their tiny size, atoms and molecules cannot be counted by direct observation. But much as we do when "counting" beans in a jar, we can estimate the number of particles in a sample of an element or compound if we have some idea of the volume occupied by each particle and the volume of the container. Once this has been done, we know the number of formula units (to use the most general term for any combination of atoms we wish to define) in any arbitrary weight of the substance. The number will of course depend both on the formula of the substance and on the weight of the sample. However, if we consider a weight of substance that is the same as its formula (molecular) weight expressed in grams, we have only one number to know: **Avogadro's number**.

Avogadro's number

Avogadro's number is known to ten significant digits:

$$N_A = 6.022141527 \times 10^{23}.$$

However, you only need to know it to three significant figures:

$$N_A \approx 6.02 \times 10^{23}. \tag{5.4.1}$$

So 6.02×10^{23} of *what*? Well, of anything you like: apples, stars in the sky, burritos. However, the only *practical* use for N_A is to have a more convenient way of expressing the huge numbers of the tiny particles such as atoms or molecules that we deal with in chemistry. Avogadro's number is a *collective number*, just like a dozen. Students can think of 6.02×10^{23} as the "chemist's dozen".

Before getting into the use of Avogadro's number in problems, take a moment to convince yourself of the reasoning embodied in the following examples.

Things to understand about Avogadro's number

- It is a *number*, just as is "dozen", and thus is *dimensionless*.
- It is a *huge* number, far greater in magnitude than we can visualize
- Its practical use is limited to counting tiny things like atoms, molecules, "formula units", electrons, or photons.
- The value of N_A can be known only to the precision that the number of atoms in a measurable weight of a substance can be estimated. Because large numbers of atoms cannot be counted directly, a variety of ingenious indirect measurements have been made involving such things as Brownian motion and X-ray scattering.

The Mole: "A Dozen Eggs and a Mole of Sugar, Please"

The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly 6.02214076×10^{23} particles, e.g., atoms, molecules, ions or electrons. The current definition was adopted in November 2018, revising its old definition based on the number of atoms in 12 grams of carbon-12 (^{12}C) (the isotope of carbon with relative atomic mass 12 Daltons by definition).

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.02 × 10^{23}.

The Mole

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Video $5.4.1$ *How big is a mole?*

Converting Between Number of Atoms to Moles and Vice Versa

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of

atoms of an element X, we can convert it into moles of by using the relationship

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

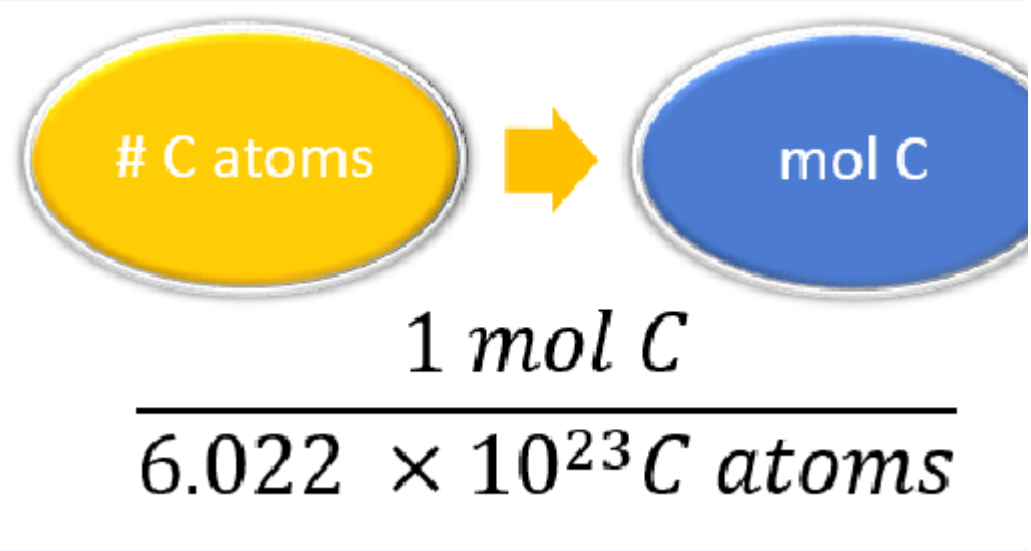
An example on the use of Avogadro's number as a conversion factor is given below for carbon.

Example 5.4.1: Moles of Carbon

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

Example 5.4.1*:Steps for Problem Solving how many moles of carbon atoms is* 4.72×10^{24} *atoms of carbon?*

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: 4.72×10^{24} C atoms Find: mol C
List other known quantities	$1\ mol = 6.022 \times 10^{23}$ C atoms
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$4.72 \times 10^{24}\ \cancel{\text{C atoms}} \times \frac{1 \text{ mol C}}{6.02 \times 10^{23}\ \cancel{\text{C atoms}}} = 7.84 \text{ mol C}$
Think about your result.	The given number of carbon atoms was greater than Avogadro's number,so the number of moles of C atoms mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calcul to three significant figures

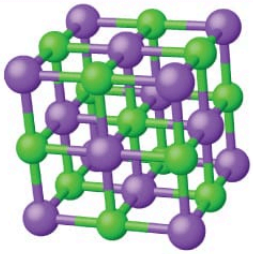
Formula Mass

One skill needed in future chapters is the ability to determine the mass of the formula of various chemical substances. This quantity is called the formula mass. The **formula mass** is obtained by adding the masses of each individual atom in the formula of the substance. Because a proper formula is electrically neutral (with no net electrons gained or lost), the ions can be considered

atoms for the purpose of calculating the formula mass.

Let us start by calculating the formula mass of sodium chloride (NaCl). This formula mass is the sum of the atomic masses of one sodium atom and one chlorine atom, which we find from the periodic table; here, we use the masses to two decimal places:

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

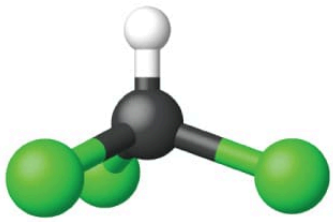


A table and diagram are shown. The table is made up of six columns and four rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space and “Subtotal (a m u).” The first column contains the symbols “N a”, “C l,” and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers “1” and “1” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “22.99” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “22.99,” “35.45,” and “58.44.” There is a thick black line below the number “35.45.” The merged cell under the first five columns reads “Formula mass.” To the left of the table is a diagram of a chemical structure. The diagram shows green and purple spheres placed in an alternating pattern, making up the corners of eight stacked cubes to form one larger cube. The green spheres are slightly smaller than the purple spheres.

To two decimal places, the formula mass of NaCl is 58.44 amu.

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a **molecular mass**. Consider chloroform ($CHCl_3$), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the “anti-stick” polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37



A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space, and “Subtotal (a m u).” The first column contains the symbols “C,” “H,” “C l” and a blank, merged cell that runs the width of the first five columns. The second column contains the numbers “1,” “1,” and “3” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “12.01,” “1.008,” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “12.01,” “1.008,” “106.35,” and “119.37.” There is a thick black line below the number 106.35. The merged cell under the first five columns reads “Molecular mass.” To the left of the table is a diagram of a molecule. Three green spheres are attached to a slightly smaller black sphere, which is also attached to a smaller white sphere. The green spheres lie beneath and to the sides of the black sphere while the white sphere is located straight up from the black sphere.

To two decimal places, the formula mass of $CHCl_3$ is 119.37 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. as shown in the example below for aluminum sulfate, $Al_2(SO_4)_3$.

Example 5.4.2 Formula Mass for an Ionic Compound

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

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The formula mass for $Al_2(SO_4)_3$, is 342.14 amu.

Exercise 5.4.1

Use the atomic masses (rounded to two decimal places) to determine the formula mass for each ionic compound.

a. TiO_2
b. AgBr
c. $Au(NO_3)_3$
d. $Fe_3(PO_4)_2$

Answer

a. 79.87 amu

b. 187.77 amu

c. 383.0 amu

Percent Composition of a Compound from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example 5.4.3

Dichlorine heptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorine heptoxide.

Solution

Example 5.4.3:*Steps for Problem Solving calculate the percent composition of dichlorine heptoxide* (Cl_2O_7).

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given : Cl_2O_7 Find: % Composition (% Cl and %O)
List other known quantities	Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g Molar mass of Cl_2O_7 = 182.90 g/mol
Cancel units and calculate.	$\%\mathrm{Cl} = \frac{70.90\ \mathrm{g\ Cl}}{182.90\ \mathrm{g}} \times 100\% = 38.76\%\ \mathrm{Cl}$ $\%\ \mathrm{O} = \frac{112.00\ \mathrm{g\ O}}{182.90\ \mathrm{g}} \times 100\% = 61.24\%\ \mathrm{O}$ Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the co molar mass of the compound and multiplying by 100%.
Think about your result.	The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\%\ \mathrm{Cl}$ and $61.24\%\ \mathrm{O}$. Suppose that you needed to know the masses of chlorine and oxygen present in a $12.50\ \mathrm{g}$ sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element.

$$12.50\ \mathrm{g\ Cl_2O_7} \times \frac{38.76\ \mathrm{g\ Cl}}{100\ \mathrm{g\ Cl_2O_7}} = 4.845\ \mathrm{g\ Cl}$$

$$12.50\ \mathrm{g\ Cl_2O_7} \times \frac{61.24\ \mathrm{g\ O}}{100\ \mathrm{g\ Cl_2O_7}} = 7.655\ \mathrm{g\ O}$$

The sum of the two masses is $12.50\ \mathrm{g}$, the mass of the sample size.

Exercise 5.4.2

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer a: available at the end of this section.

Summary

- The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly 6.02214076×10^{23} particles, e.g., atoms, molecules, ions or electrons.
- Avogadro's number is related to moles of any substance X as follows:

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

- Formula masses of ionic and molecular compounds can be determined from the masses of the atoms in their formulas.
- Processes are described for calculating the percent composition of a compound based on the chemical formula.

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- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Wikipedia

Learning Objectives

- Describe the unit *mole*.
- Relate the mole quantity of substance to its mass.

So far, we have been talking about chemical substances in terms of individual atoms and molecules. Yet we do not typically deal with substances an atom or a molecule at a time; we work with millions, billions, and trillions of atoms and molecules at a time. We need a way to deal with macroscopic, rather than microscopic, amounts of matter. We need a unit of amount that relates quantities of substances on a scale that we can interact with.

Chemistry uses a unit called mole. A **mole** (mol) is a number of things equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \times 10^{23} \text{ things}$$

Understand that a mole means a specific number of things, just like a dozen means a certain number of things—twelve, in the case of a dozen. But a mole is a much larger number of things. These things can be atoms, or molecules, or eggs; however, in chemistry, we usually use the mole to refer to the amounts of atoms or molecules. Although the number of things in a mole is known to eight decimal places, it is usually fine to use only two or three decimal places in calculations. The numerical value of things in a mole is often called *Avogadro's number* (N_A). *Avogadro's number* is also known as the *Avogadro constant*, after Amadeo Avogadro, an Italian chemist who first proposed its importance.

Example 5.4.1

How many molecules are present in 2.76 mol of H_2O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O, we can also determine the number of atoms in the sample.

$$2.76 \ \cancel{mol\ H_2O} \times \frac{6.022 \times 10^{23} molecules\ H_2O}{\cancel{mol\ H_2O}} = 1.66 \times 10^{24} molecules\ H_2O$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24}\ \cancel{molecules\ H_2O} \times \frac{3\ atoms}{1\ molecule} = 4.99 \times 10^{24}\ atoms$$

Exercise 5.4.1

How many molecules are present in 4.61×10^{-2} mol of O_2?

Answer available at the end of this section.

How big is a mole? It is very large. Suppose you had a mole of dollar bills that need to be counted. If everyone on earth (about 6 billion people) counted one bill per second, it would take about 3.2 million years to count all the bills. A mole of sand would fill a cube about 32 km on a side. A mole of pennies stacked on top of each other would have about the same diameter as our galaxy, the Milky Way. A mole is a lot of things—but atoms and molecules are very tiny. One mole of carbon atoms would make a cube that is 1.74 cm on a side, small enough to carry in your pocket.

Why is the mole unit so important? It represents the link between the microscopic and the macroscopic, especially in terms of mass. *A mole of a substance has the same mass in grams as one unit (atom or molecules) has in atomic mass units*. The mole unit allows us to express amounts of atoms and molecules in visible amounts that we can understand.

For example, we already know that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022×10^{23} atoms:

$$12\ g\ C = 6.022 \times 10^{23}\ atoms\ C$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50×10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25}\ \cancel{atoms\ C} \times \frac{12.0000\ g\ C}{6.022 \times 10^{23}\ \cancel{atoms\ C}} = 299\ g\ C$$

But it also goes beyond carbon. Previously we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the number of grams in 1 mol of atoms of that element, while the molecular mass of a compound is the number of grams in 1 mol of molecules of that compound. Sometimes these masses are called **molar masses** to emphasize the fact that they are the mass for 1 mol of things. (The term *molar* is the adjective form of mole and has nothing to do with teeth.)

Here are some examples. The mass of a hydrogen atom is 1.0079 u; the mass of 1 mol of hydrogen atoms is 1.0079 g. Elemental hydrogen exists as a diatomic molecule, H_2. One molecule has a mass of 1.0079 + 1.0079 = 2.0158 u, while 1 mol H_2 has a mass of 2.0158 g. A molecule of H_2O has a mass of about 18.01 u; 1 mol H_2O has a mass of 18.01 g. A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g. In each of these moles of substances, there are 6.022×10^{23} units: 6.022×10^{23} atoms of H, 6.022×10^{23} molecules of H_2 and H_2O, 6.022×10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

Example 5.4.2: Sugar

What is the molar mass of sugar ($C_6H_{12}O_6$)?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula; but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table.

Solutions to Example 5.3.2

6 C = 6 × 12.011	= 72.066
12 H = 12 × 1.0079	= 12.0948
6 O = 6 × 15.999	= 95.994
TOTAL	= 180.155 g/mol

Per convention, the unit *grams per mole* is written as a fraction.

Exercise 5.4.2

What is the molar mass of $\mathrm{AgNO_3}$?

Answer available at the end of this section.

Knowing the molar mass of a substance, we can calculate the number of moles in a certain mass of a substance and vice versa, as these examples illustrate. The molar mass is used as the conversion factor.

Example 5.4.3

What is the mass of 3.56 mol of HgCl2? The molar mass of HgCl2 is 271.49 g/mol.

Solution

Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$3.56\ \cancel{mol\ HgCl_2} \times \frac{271.49\ g\ HgCl_2}{\cancel{mol\ HgCl_2}} = 967\ g\ HgCl_2$$

Exercise 5.4.3

What is the mass of 33.7 mol of $\mathrm{H_2O}$?

Answer available at the end of this section.

Example 5.4.4

How many moles of H2O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H2O as a conversion factor from mass to moles. The molar mass of water is

(1.0079 + 1.0079 + 15.999) = 18.015 g/mol.

However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.015\ \cancel{g\ H_2O}} = 13.32\ mol\ H_2O$$

Exercise 5.4.4

How many moles are present in 35.6 g of H_2SO_4 (molar mass = 98.08 g/mol)?

Answer available at the end of this section.

Other conversion factors can be combined with the definition of mole—density, for example.

Example 5.4.5

The density of ethanol is 0.789 g/mL. How many moles are in 100.0 mL of ethanol? The molar mass of ethanol is 46.08 g/mol.

Solution

Here, we use density to convert from volume to mass and then use the molar mass to determine the number of moles.

$$100 \cancel{ml}\ ethanol \times \frac{0.789\,g}{\cancel{ml}} \times \frac{1\,mol}{46.08\,\cancel{g}} = 1.71\,mol\,ethanol$$

Exercise 5.4.5

If the density of benzene, C_6H_6, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer available at the end of this section.

Summary

The mole is a key unit in chemistry. The molar mass of a substance, in grams, is numerically equal to one atom's or molecule's mass in atomic mass units.

Answers

Exercise 5.4.2
78.32% Ba and 21.67% F

Exercise 5.4.1
2.78×10^{22} molecules

Exercise 5.4.2
169.87 g/mol

Exercise 5.4.3
607 g

Exercise 5.4.4
0.363 mol

Exercise 5.4.5
0.201 mol

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5.5: Molar Mass

Molar Mass

The molar mass of a substance is defined as the mass in grams of 1 mole of that substance. One mole of isotopically pure carbon-12 has a mass of 12 g. For an element, the molar mass is the mass of 1 mol of atoms of that element; for a covalent molecular compound, it is the mass of 1 mol of molecules of that compound; for an ionic compound, it is the mass of 1 mol of formula units. That is, the molar mass of a substance is the mass (in grams per mole) of 6.022×10^{23} atoms, molecules, or formula units of that substance. In each case, the number of grams in 1 mol is the same as the number of atomic mass units that describe the atomic mass, the molecular mass, or the formula mass, respectively.

> The molar mass of any substance is its atomic mass, molecular mass, or formula mass in grams per mole.

The periodic table lists the atomic mass of carbon as 12.011 amu; the average molar mass of carbon—the mass of 6.022×10^{23} carbon atoms—is therefore 12.011 g/mol:

Table 5.5.1 *Molar Mass of Select Substances*

Substance (formula)	Basic Unit	Atomic, Molecular, or Formula Mass (amu)	Molar Mass (g/mol)
carbon (C)	atom	12.011 (atomic mass)	12.011
ethanol (C_2H_5OH)	molecule	46.069 (molecular mass)	46.069
calcium phosphate [$Ca_3(PO_4)_2$]	formula unit	310.177 (formula mass)	310.177

Summary

- The molar mass of a substance is defined as the mass in grams of 1 mole of that substance.

Contributors and Attributions

-
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

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5.6: Mole Conversions

Converting Between Number of Atoms to Moles and Vice Versa

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of atoms of an element X, we can convert it into moles by using the relationship

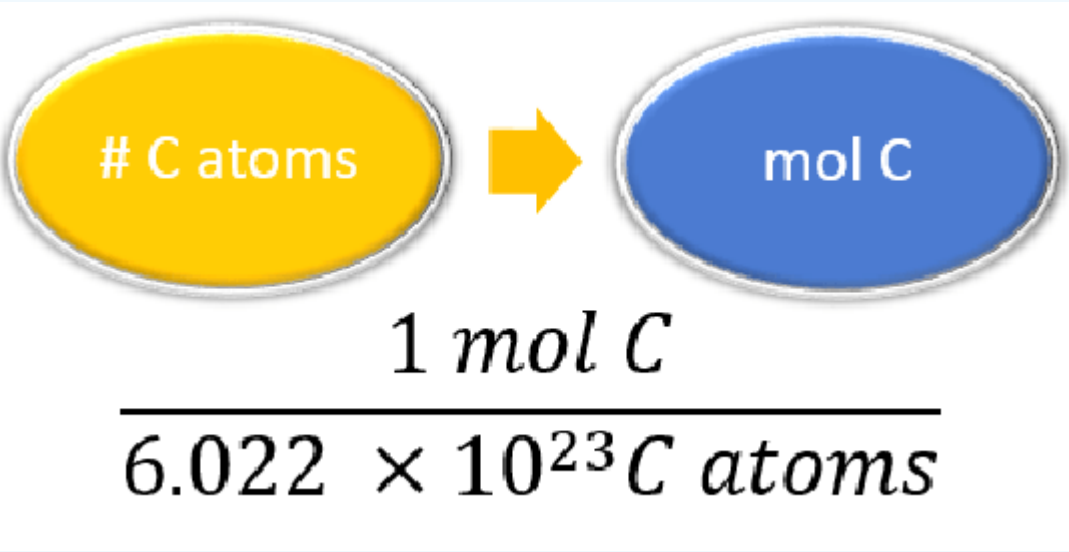
$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms}$$

Example 5.6.1: Moles of Carbon

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

Example 5.6.1: Steps for Problem Solving; Molar Conversion of Carbon

Steps for Problem Solving	The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atom $\times 10^{24}$ atoms of carbon?
Identify the "given" information and what the problem is asking you to "find."	Given: 4.72×10^{24} C atoms Find: mol C
List other known quantities.	$1\ mol = 6.022 \times 10^{23}$ C atoms
Prepare a concept map and use the proper conversion factor.	

Steps for Problem Solving	The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atom $\times 10^{24}$ atoms of carbon?
Cancel units and calculate.	$4.72 \times 10^{24}\ \cancel{\text{C atoms}} \times \dfrac{1\ \text{mol C}}{6.02 \times 10^{23}\ \cancel{\text{C atoms}}} = 7.84\ \text{mol C}$
Think about your result.	The given number of carbon atoms was greater than Avogadro's number, so the number of moles of C atoms is gr 1 mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calcula rounded to three significant figures.

Converting Between Grams and Moles

The molar mass of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. Suppose that for a certain experiment you need 3.00 moles of calcium chloride (CaCl_2). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. Dimensional analysis will allow you to calculate the mass of CaCl_2 that you should measure, as shown in Example 5.6.2.

Example 5.6.2: Calcium Chloride

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

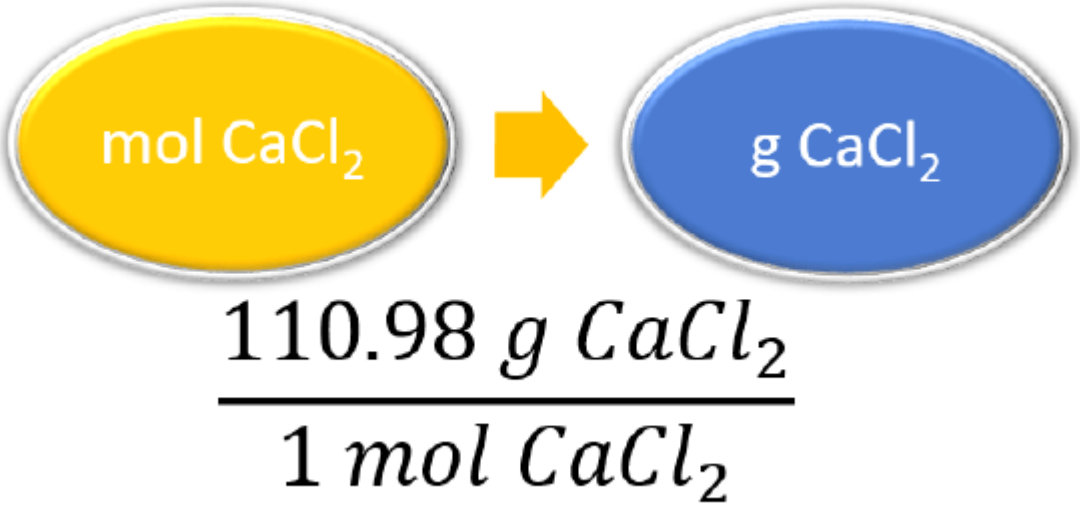


Figure 5.6.3: Calcium chloride is used as a drying agent and as a road deicer.

Solution

Example 5.6.2:Steps for Problem Solving; Calculate the mass of 3.00 moles of calcium chloride ($CaCl_2$).

Steps for Problem Solving	Calculate the mass of 3.00 moles of calcium chloride ($CaCl_2$).
Identify the "given" information and what the problem is asking you to "find."	Given: 3.00 moles of $CaCl_2$ Find: g $CaCl_2$
List other known quantities.	1 ml $CaCl_2$ = 110.98 g $CaCl_2$

Steps for Problem Solving	Calculate the mass of 3.00 moles of calcium chloride ($CaCl_2$).
Prepare a concept map and use the proper conversion factor.	 $\frac{110.98\ g\ CaCl_2}{1\ mol\ CaCl_2}$
Cancel units and calculate.	$3.00\ \cancel{\text{mol CaCl}_2} \times \frac{110.98\ \text{g CaCl}_2}{1\ \cancel{\text{mol CaCl}_2}} = 333\ \text{g CaCl}_2$
Think about your result.	

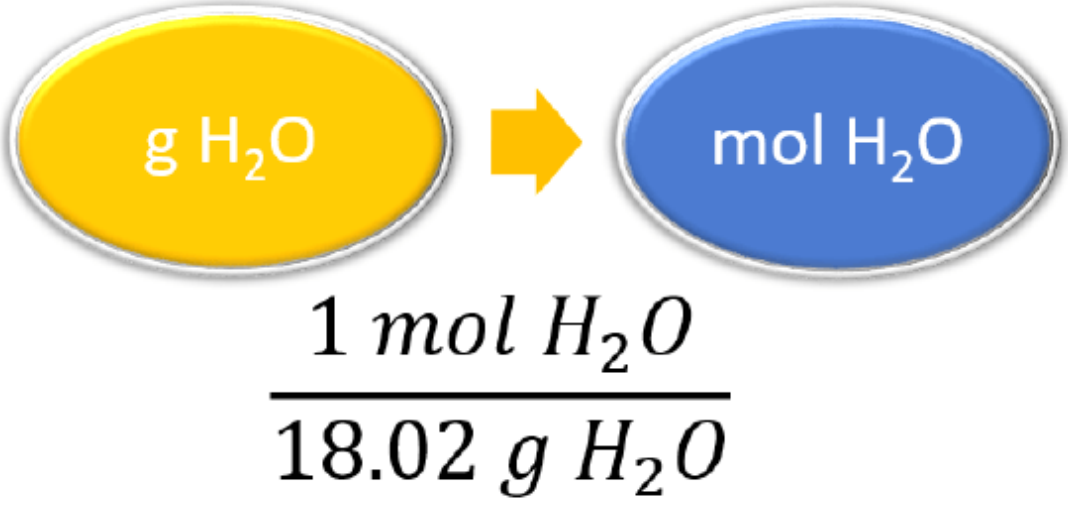
Example 5.6.3: Water

How many moles are present in 108 grams of water?

Solution

Example 5.6.3: Steps for Problem Solving; how many moles are present in 108 grams of water

Steps for Problem Solving	How many moles are present in 108 grams of water?
Identify the "given" information and what the problem is asking you to "find."	Given: 108 g H_2O Find: mol H_2O
List other known quantities.	$1\ \text{mol}\ H_2O = 18.02\ \text{g}\ H_2O$

Steps for Problem Solving	How many moles are present in 108 grams of water?
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ H_2O}{18.02\ g\ H_2O}$
Cancel units and calculate.	$108\ \cancel{\text{g H}_2\text{O}} \times \frac{1\ \text{mol H}_2\text{O}}{18.02\ \cancel{\text{g H}_2\text{O}}} = 5.99\ \text{mol H}_2\text{O}$
Think about your result.	

Exercise 5.6.1: Nitrogen Gas

What is the mass of $7.50\ \text{mol}$ of Nitrogen gas N_2?

Answer: available at the end of this section.

Conversions Between Mass and Number of Particles

In "Conversions Between Moles and Mass", you learned how to convert back and forth between moles and the number of representative particles. Now you have seen how to convert back and forth between moles and mass of a substance in grams. We can combine the two types of problems into one. Mass and number of particles are both related to moles. To convert from mass to number of particles or vice-versa, it will first require a conversion to moles, as shown in Figure 5.6.1 and Example 5.6.4.



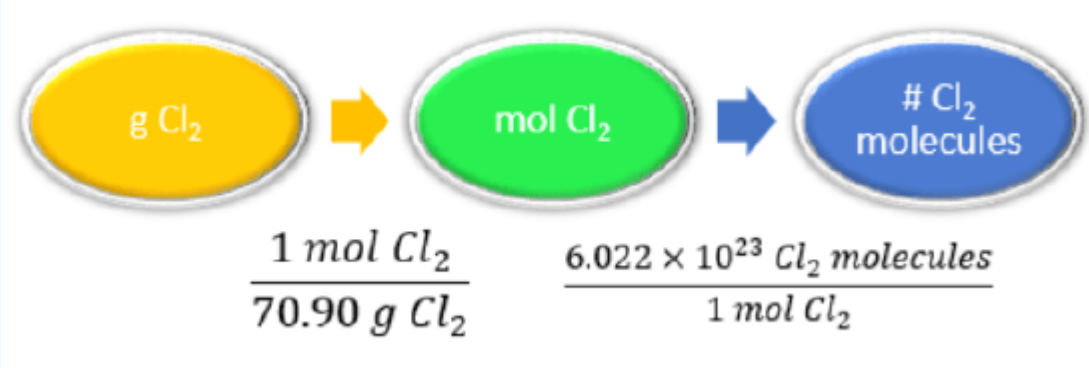
Figure 5.6.4: Conversion from number of particles to mass, or from mass to number of particles, requires two steps. To convert from number of particles to moles, use mol/Avogrado's #, and to convert from moles to mass, use g/mol.

Example 5.6.4: Chlorine

How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Solution

Example 5.6.4:Steps for Problem Solving how many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?

Steps for Problem Solving	How many molecules is $20.0\ \text{g}$ of chlorine gas, Cl_2?
Identify the "given" information and what the problem is asking you to "find."	Given: 20.0 g Cl_2 Find: # Cl_2 molecules
List other known quantities.	1 mol Cl_2 = 70.90 g Cl_2, 1mol Cl_2 = 6.022 x 10^{23} Cl_2 molecules
Prepare a concept map and use the proper conversion factor.	 The conversion factors are 1 mole Cl2 over 70.90 grams Cl2, and 6.022 times 10^23 Cl2 molecules over 1 mole Cl2.
Cancel units and calculate.	$20.0\ \cancel{\text{g Cl}_2} \times \frac{1\ \cancel{\text{mol Cl}_2}}{70.90\ \cancel{\text{g Cl}_2}} \times \frac{6.02 \times 10^{23}\ \text{molecules Cl}_2}{1\ \cancel{\text{mol Cl}_2}} = 1.70 \times 10^{23}\ \text{molecules Cl}_2$
Think about your result.	Since the given mass is less than half of the molar mass of chlorine, the resulting number of molecules is less than half of Avogadro's number.

Exercise 5.6.2: Calcium Chloride

How many formula units are in 25.0 g of $CaCl_2$?

Answer: available at the end of this section.

Summary

- Calculations involving conversions between moles of a material and the mass of that material are described.
- Calculations are illustrated for conversions between mass and number of particles.

Contributors and Attributions

-

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Learning Objectives

- Use Avogadro's number to convert to moles and vice versa given the number of particles of an element.
- Know the definition of the mole.
- Determine the formula mass of an ionic or molecular compound.
- Determine the percent composition of each element in a compound from the chemical formula.

When objects are very small, it is often inconvenient or inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross".

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number:Counting Atoms

Owing to their tiny size, atoms and molecules cannot be counted by direct observation. But much as we do when "counting" beans in a jar, we can estimate the number of particles in a sample of an element or compound if we have some idea of the volume occupied by each particle and the volume of the container. Once this has been done, we know the number of formula units (to use the most general term for any combination of atoms we wish to define) in any arbitrary weight of the substance. The number will of course depend both on the formula of the substance and on the weight of the sample. However, if we consider a weight of substance that is the same as its formula (molecular) weight expressed in grams, we have only one number to know: **Avogadro's number**.

Avogadro's number

Avogadro's number is known to ten significant digits:

$$N_A = 6.022141527 \times 10^{23}.$$

However, you only need to know it to three significant figures:

$$N_A \approx 6.02 \times 10^{23}. \tag{5.6.1}$$

So 6.02×10^{23} of *what*? Well, of anything you like: apples, stars in the sky, burritos. However, the only *practical* use for N_A is to have a more convenient way of expressing the huge numbers of the tiny particles such as atoms or molecules that we deal with in chemistry. Avogadro's number is a *collective number*, just like a dozen. Students can think of 6.02×10^{23} as the "chemist's dozen".

Before getting into the use of Avogadro's number in problems, take a moment to convince yourself of the reasoning embodied in the following examples.

Things to understand about Avogadro's number

- It is a *number*, just as is "dozen", and thus is *dimensionless*.
- It is a *huge* number, far greater in magnitude than we can visualize
- Its practical use is limited to counting tiny things like atoms, molecules, "formula units", electrons, or photons.

- The value of N_A can be known only to the precision that the number of atoms in a measurable weight of a substance can be estimated. Because large numbers of atoms cannot be counted directly, a variety of ingenious indirect measurements have been made involving such things as Brownian motion and X-ray scattering.

The Mole: "A Dozen Eggs and a Mole of Sugar, Please"

The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly $6.02214076 \times 10^{23}$ particles, e.g., atoms, molecules, ions or electrons. The current definition was adopted in November 2018, revising its old definition based on the number of atoms in 12 grams of carbon-12 (^{12}C) (the isotope of carbon with relative atomic mass 12 Daltons by definition).

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.02×10^{23}.

The Mole

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video $5.6.1$ *How big is a mole?*

Converting Between Number of Atoms to Moles and Vice Versa

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of atoms of an element X, we can convert it into moles of by using the relationship

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

An example on the use of Avogadro's number as a conversion factor is given below for carbon.

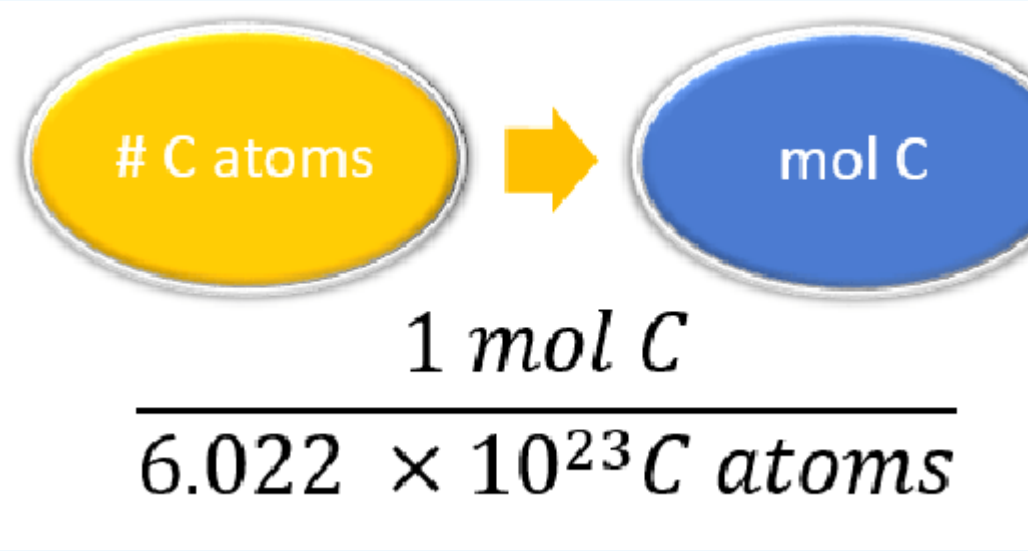
Example $5.6.1$: Moles of Carbon

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

Example $5.6.1$*:Steps for Problem Solving how many moles of carbon atoms is* 4.72×10^{24} *atoms of carbon?*

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: 4.72×10^{24} C atoms Find: mol C

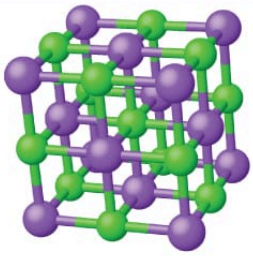
Steps for Problem Solving	Example
List other known quantities	$1\ mol = 6.022 \times 10^{23}$ C atoms
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ C}{6.022 \times 10^{23} C\ atoms}$
Cancel units and calculate.	$4.72 \times 10^{24}\ \cancel{\text{C atoms}} \times \frac{1\ \text{mol C}}{6.02 \times 10^{23}\ \cancel{\text{C atoms}}} = 7.84\ \text{mol C}$
Think about your result.	The given number of carbon atoms was greater than Avogadro's number,so the number of moles of C atoms mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calcul to three significant figures

Formula Mass

One skill needed in future chapters is the ability to determine the mass of the formula of various chemical substances. This quantity is called the formula mass. The **formula mass** is obtained by adding the masses of each individual atom in the formula of the substance. Because a proper formula is electrically neutral (with no net electrons gained or lost), the ions can be considered atoms for the purpose of calculating the formula mass.

Let us start by calculating the formula mass of sodium chloride (NaCl). This formula mass is the sum of the atomic masses of one sodium atom and one chlorine atom, which we find from the periodic table; here, we use the masses to two decimal places:

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

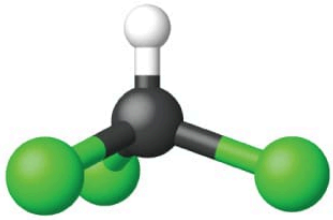


A table and diagram are shown. The table is made up of six columns and four rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space and “Subtotal (a m u).” The first column contains the symbols “N a”, “C l,” and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers “1” and “1” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “22.99” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “22.99,” “35.45,” and “58.44.” There is a thick black line below the number “35.45.” The merged cell under the first five columns reads “Formula mass.” To the left of the table is a diagram of a chemical structure. The diagram shows green and purple spheres placed in an alternating pattern, making up the corners of eight stacked cubes to form one larger cube. The green spheres are slightly smaller than the purple spheres.

To two decimal places, the formula mass of NaCl is 58.44 amu.

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a **molecular mass**. Consider chloroform ($CHCl_3$), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the “anti-stick” polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37



A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space, and “Subtotal (a m u).” The first column contains the symbols “C,” “H,” “C l” and a blank, merged cell that runs the width of the first five columns. The second column contains the numbers “1,” “1,” and “3” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “12.01,” “1.008,” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “12.01,” “1.008,” “106.35,” and “119.37.” There is a thick black line below the number 106.35. The merged cell under the first five columns reads “Molecular mass.” To the left of the table is a diagram of a molecule. Three green spheres are attached to a slightly smaller black sphere, which is also attached to a smaller white sphere. The green spheres lie beneath and to the sides of the black sphere while the white sphere is located straight up from the black sphere.

To two decimal places, the formula mass of $CHCl_3$ is 119.37 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. as shown in the example below for aluminum sulfate, $Al_2(SO_4)_3$.

Example 5.6.2 Formula Mass for an Ionic Compound

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

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The formula mass for $Al_2(SO_4)_3$, is 342.14 amu.

Exercise 5.6.1

Use the atomic masses (rounded to two decimal places) to determine the formula mass for each ionic compound.

a. TiO_2
b. AgBr
c. $Au(NO_3)_3$
d. $Fe_3(PO_4)_2$

Answer

a. 79.87 amu

b. 187.77 amu

c. 383.00 amu

Percent Composition of a Compound from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example 5.6.3

Dichlorine heptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorine heptoxide.

Solution

Example 5.6.3:*Steps for Problem Solving calculate the percent composition of dichlorine heptoxide* (Cl_2O_7).

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given : Cl_2O_7 Find: % Composition (% Cl and %O)
List other known quantities	Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g Molar mass of Cl_2O_7 = 182.90 g/mol
Cancel units and calculate.	$\%\mathrm{Cl} = \frac{70.90\ \mathrm{g\ Cl}}{182.90\ \mathrm{g}} \times 100\% = 38.76\%\ \mathrm{Cl}$ $\%\ \mathrm{O} = \frac{112.00\ \mathrm{g\ O}}{182.90\ \mathrm{g}} \times 100\% = 61.24\%\ \mathrm{O}$ Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the co molar mass of the compound and multiplying by 100%.
Think about your result.	The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\%\ \mathrm{Cl}$ and $61.24\%\ \mathrm{O}$. Suppose that you needed to know the masses of chlorine and oxygen present in a $12.50\ \mathrm{g}$ sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element.

$$12.50\ \mathrm{g\ Cl_2O_7} \times \frac{38.76\ \mathrm{g\ Cl}}{100\ \mathrm{g\ Cl_2O_7}} = 4.845\ \mathrm{g\ Cl}$$

$$12.50\ \mathrm{g\ Cl_2O_7} \times \frac{61.24\ \mathrm{g\ O}}{100\ \mathrm{g\ Cl_2O_7}} = 7.655\ \mathrm{g\ O}$$

The sum of the two masses is $12.50\ \mathrm{g}$, the mass of the sample size.

Exercise 5.6.2

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer a: available at the end of this section.

Summary

- The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly 6.02214076×10^{23} particles, e.g., atoms, molecules, ions or electrons.
- Avogadro's number is related to moles of any substance X as follows:

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

- Formula masses of ionic and molecular compounds can be determined from the masses of the atoms in their formulas.
- Processes are described for calculating the percent composition of a compound based on the chemical formula.

Contributors and Attributions

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- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).
- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Wikipedia

Learning Objectives

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

Molar Mass

The **molar mass** of any substance is the mass in grams of one mole of representative particles of that substance. The representative particles can be atoms, molecules, or formula units of ionic compounds. This relationship is frequently used in the laboratory. The simplest type of manipulation using molar mass as a conversion factor is a mole-mass conversion (or its reverse, a mass-mole conversion). In such a conversion, we use the molar mass of a substance as a conversion factor to convert mole units into mass units (or, conversely, mass units into mole units).

We also established that 1 mol of Al has a mass of 26.98 g (Example). Stated mathematically,

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

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

$$\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \text{ and } \frac{26.98 \text{ g Al}}{1 \text{ mol Al}} \tag{5.6.2}$$

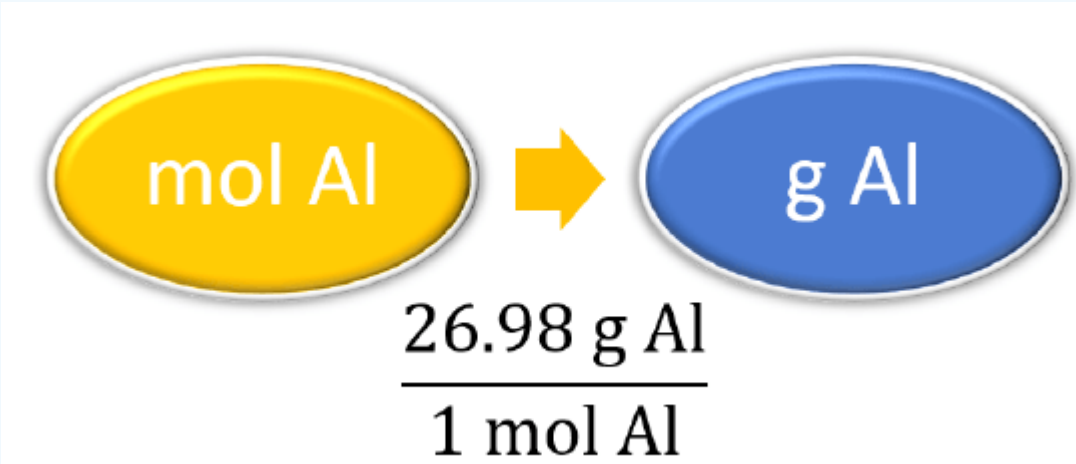
The first conversion factor can be used to convert from mass to moles, and the second converts from moles to mass. Both can be used to solve problems that would be hard to do “by eye.”

Example 5.6.1

What is the mass of 3.987 mol of Al?

Solution

Example 5.6.1: Steps for Problem Solving the mass of 3.987 moles of aluminum (Al).

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

Exercise 5.6.1

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

Answer

3.71 g Al

Conversions like this are possible for any substance, as long as the proper atomic mass, formula mass, or molar mass is known (or can be determined) and expressed in grams per mole. Figure 5.6.1 is a chart for determining what conversion factor is needed, and Figure 5.6.2 is a flow diagram for the steps needed to perform a conversion.

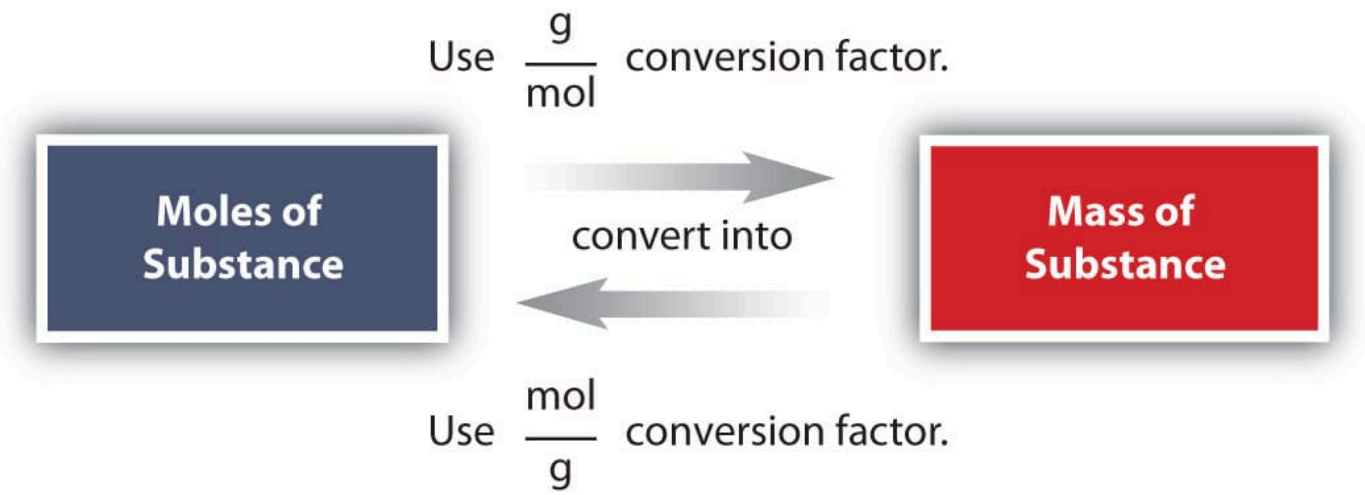


Figure 5.6.1 A Simple Flowchart for Converting between Mass and Moles of a Substance. It takes one mathematical step to convert from moles to mass or from mass to moles.

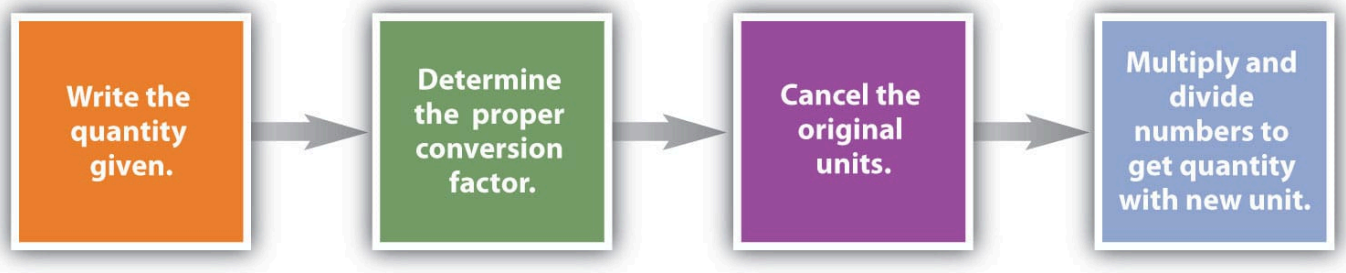


Figure 5.6.2 A Flowchart Illustrating the Steps in Performing a Unit Conversion. When performing many unit conversions, the same logical steps can be taken.

Suppose that for a certain experiment you need 3.00 moles of calcium chloride ($CaCl_2$). Since calcium chloride is a solid, it would be convenient to use a balance to measure the mass that is needed. Dimensional analysis will allow you to calculate the mass of $CaCl_2$ that you should measure as show in Example 5.6.3.

Example 5.6.2: Calcium Chloride

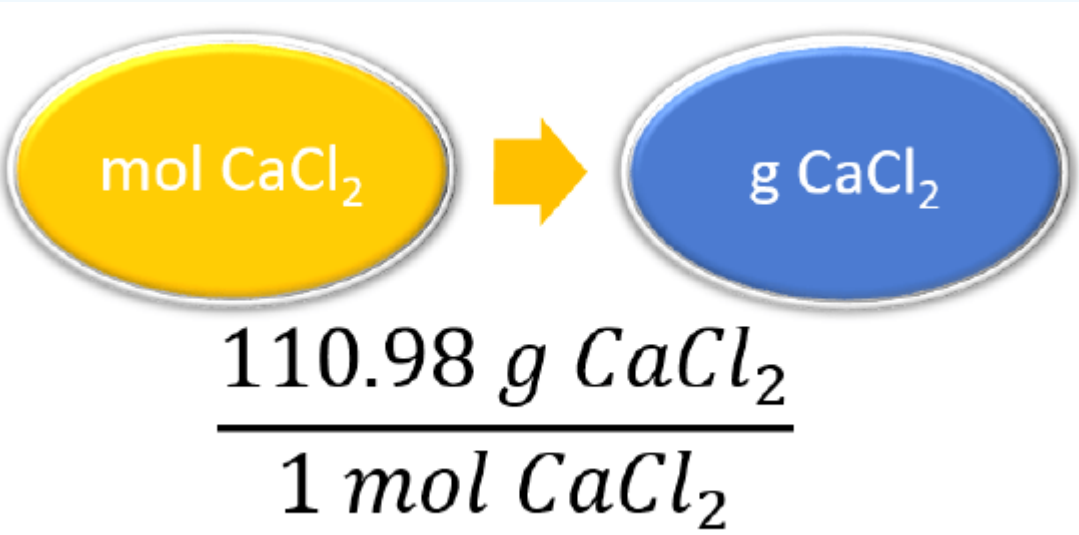
Calculate the mass of 3.00 moles of calcium chloride ($CaCl_2$).



Figure 5.6.1: Calcium chloride is used as a drying agent and as a road deicer.

Solution

Example 5.6.2: Steps for Problem Solving the mass of 3.00 moles of calcium chloride (CaCl2).

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

Exercise 5.6.2: Calcium Oxide

What is the mass of $7.50\ \text{mol}$ of (calcium oxide) CaO?

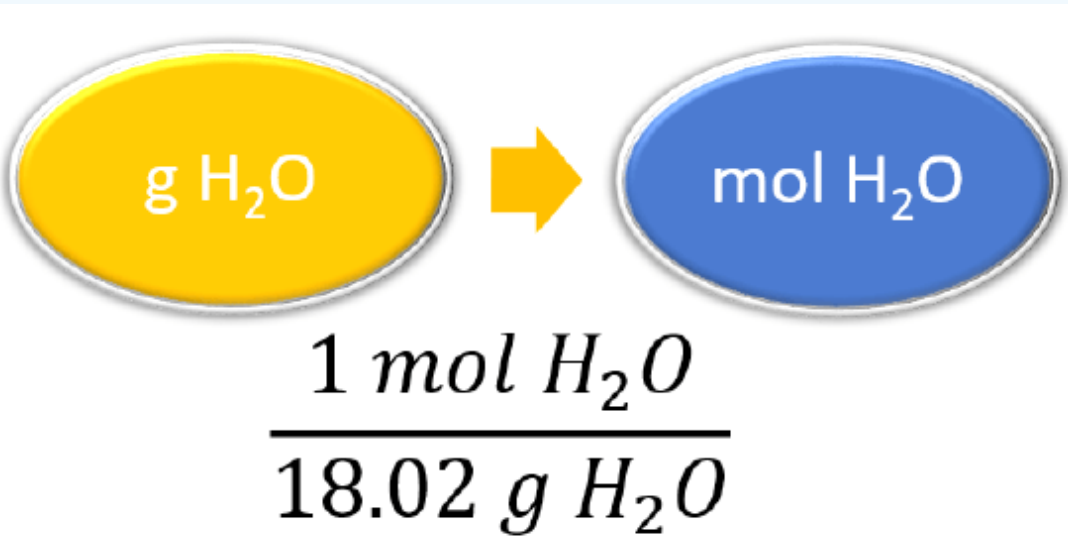
Answer: available at the end of this section.

Example 5.6.3: Water

How many moles are present in 108 grams of water?

Solution

Example 5.6.3: Steps for Problem Solving moles in 108 grams of water?

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: 108 g H_2O Find: mol H_2O
List other known quantities	$1\ \text{mol}\ H_2O = 18.02\ \text{g}\ H_2O$
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$108\ \cancel{\text{g}\ H_2O} \times \frac{1\ \text{mol}\ H_2O}{18.02\ \cancel{\text{g}\ H_2O}} = 5.99\ \text{mol}\ H_2O$
Think about your result.	

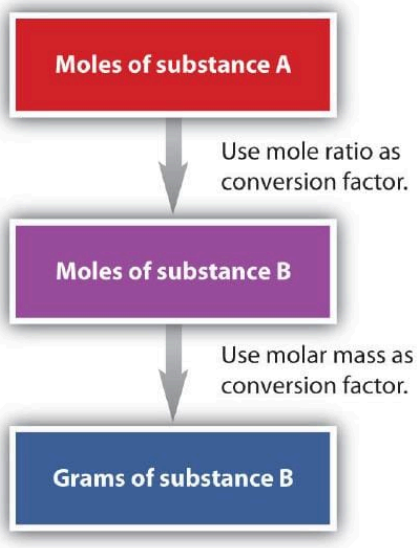
Exercise 5.6.3: Nitrogen Gas

What is the mass of $7.50\ \text{mol}$ of Nitrogen gas N_2?

Answer: available at the end of this section.

Mole and Mass Relationships in Chemical Equations

We have established that a balanced chemical equation is balanced in terms of moles as well as atoms or molecules. We have used balanced equations to set up ratios, now in terms of moles of materials, that we can use as conversion factors to answer stoichiometric questions, such as how many moles of substance A react with so many moles of reactant B. We can extend this technique even further. Recall that we can relate a molar amount to a mass amount using molar mass. We can use that ability to answer stoichiometry questions in terms of the masses of a particular substance, in addition to moles. We do this using the following sequence:



Collectively, these conversions are called mole-mass calculations.

As an example, consider the balanced chemical equation

$$Fe_2O_3 + 3SO_3 \rightarrow Fe_2(SO_4)_3$$

If we have 3.59 mol of Fe_2O_3, how many grams of SO_3 can react with it? Using the mole-mass calculation sequence, we can determine the required mass of SO_3 in two steps. First, we construct the appropriate molar ratio, determined from the balanced chemical equation, to calculate the number of moles of SO_3 needed. Then using the molar mass of SO_3 as a conversion factor, we determine the mass that this number of moles of SO_3 has.

As usual, we start with the quantity we were given:

$$3.59 \ \cancel{\text{mol Fe}_2\text{O}_3} \times \left(\frac{3 \text{ mol SO}_3}{1 \ \cancel{\text{mol Fe}_2\text{O}_3}} \right) = 10.77 \text{ mol SO}_3 \tag{5.6.3}$$

The mol Fe_2O_3 units cancel, leaving mol SO_3 unit. Now, we take this answer and convert it to grams of SO_3, using the molar mass of SO_3 as the conversion factor:

$$10.77 \ \cancel{\text{mol SO}_3} \times \left(\frac{80.06 \text{ g SO}_3}{1 \ \cancel{\text{mol SO}_3}} \right) = 862 \text{ g SO}_3 \tag{5.6.4}$$

Our final answer is expressed to three significant figures. Thus, in a two-step process, we find that 862 g of SO_3 will react with 3.59 mol of Fe_2O_3. Many problems of this type can be answered in this manner.

The same two-step problem can also be worked out in a single line, rather than as two separate steps, as follows:

$$3.59 \ \cancel{mol\ Fe_2O_3} \times \underbrace{\left(\frac{3 \ \cancel{mol\ SO_3}}{1 \ \cancel{mol\ Fe_2O_3}} \right)}_{\text{converts to moles of SO}_3} \times \underbrace{\left(\frac{80.06\ g\ SO_3}{1 \ \cancel{mol\ SO_3}} \right)}_{\text{converts to grams of SO}_3} = 862\ g\ SO_3$$

We get exactly the same answer when combining all the math steps together as we do when we calculate one step at a time.

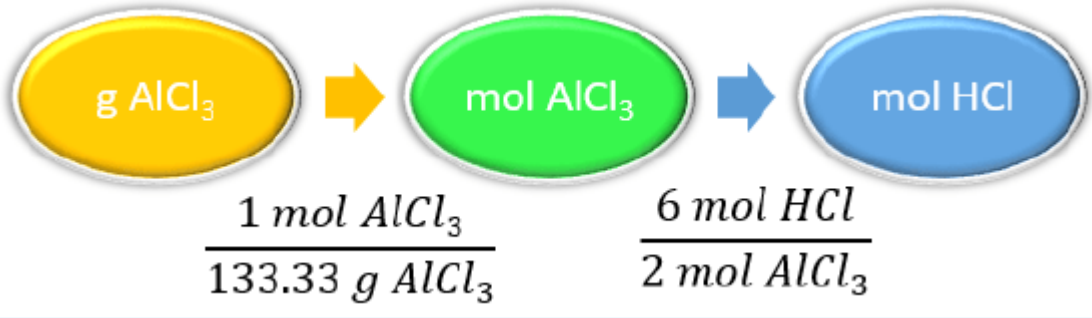
Example 5.6.4: Generation of Aluminum Oxide

How many moles of HCl will be produced when 249 g of AlCl3 are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

Example 5.6.4: Steps for Problem Solving how many moles are in 249 g of AlCl3 ?

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: 249 g AlCl3 Find: moles HCl
List other known quantities	1 mol AlCl3 = 133.33 g/mol 6 mol of HCl to 2 mol $AlCl_3$
Prepare a concept map and use the proper conversion factor.	
Cancel units and calculate.	$249\ \cancel{g\ AlCl_3} \times \frac{1\ \cancel{mol\ AlCl_3}}{133.33\ \cancel{g\ AlCl_3}} \times \frac{6\ mol\ HCl}{2\ \cancel{mol\ AlCl_3}} = 5.60\ mol\ HCl$
Think about your result.	Since 249 g of $AlCl_3$ is less than 266.66 g, the mass for 2 moles of $AlCl_3$ and the relationship is 6 mol of HCl to 2 mol $AlCl_3$, the answer should be less than 6 moles of HCl.

Exercise 5.6.4: Generation of Aluminum Oxide

How many moles of Al2O3 will be produced when 23.9 g of H2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer available at the end of this section.

Molar Relationships in Chemical Equations

Previously, you learned to balance chemical equations by comparing the numbers of each type of atom in the reactants and products. The coefficients in front of the chemical formulas represent the numbers of molecules or formula units (depending on the type of substance). Here, we will extend the meaning of the coefficients in a chemical equation.

Consider the simple chemical equation

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

The convention for writing balanced chemical equations is to use the lowest whole-number ratio for the coefficients. However, the equation is balanced as long as the coefficients are in a 2:1:2 ratio. For example, this equation is also balanced if we write it as

$$4H_2 + 2O_2 \rightarrow 4H_2O$$

The ratio of the coefficients is 4:2:4, which reduces to 2:1:2. The equation is also balanced if we were to write it as

$$22H_2 + 11O_2 \rightarrow 22H_2O$$

because 22:11:22 also reduces to 2:1:2.

Suppose we want to use larger numbers. Consider the following coefficients:

$$12.044 \times 10^{23}\ H_2 + 6.022 \times 10^{23}\ O_2 \rightarrow 12.044 \times 10^{23}\ H_2O$$

These coefficients also have the ratio 2:1:2 (check it and see), so this equation is balanced. But 6.022×10^{23} is 1 mol, while 12.044×10^{23} is 2 mol (and the number is written that way to make this more obvious), so we can simplify this version of the equation by writing it as

$$2\ mol\ H_2 + 1\ mol\ O_2 \rightarrow 2\ mol\ H_2O$$

We can leave out the word *mol* and not write the 1 coefficient (as is our habit), so the final form of the equation, still balanced, is

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

Now we interpret the coefficients as referring to molar amounts, not individual molecules. The lesson? *Balanced chemical equations are balanced not only at the molecular level but also in terms of molar amounts of reactants and products*. Thus, we can read this reaction as "two moles of hydrogen react with one mole of oxygen to produce two moles of water."

By the same token, the ratios we constructed to describe molecules reaction can also be constructed in terms of moles rather than molecules. For the reaction in which hydrogen and oxygen combine to make water, for example, we can construct the following ratios:

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

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

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

We can use these ratios to determine what amount of a substance, in moles, will react with or produce a given number of moles of a different substance. The study of the numerical relationships between the reactants and the products in balanced chemical reactions is called *stoichiometry*.

Example 5.6.5

How many moles of oxygen react with hydrogen to produce 27.6 mol of H_2O?

Solution

Example 5.6.5: Steps for Problem Solving how many moles of oxygen react with hydrogen to produce 27.6 mol of H2O?

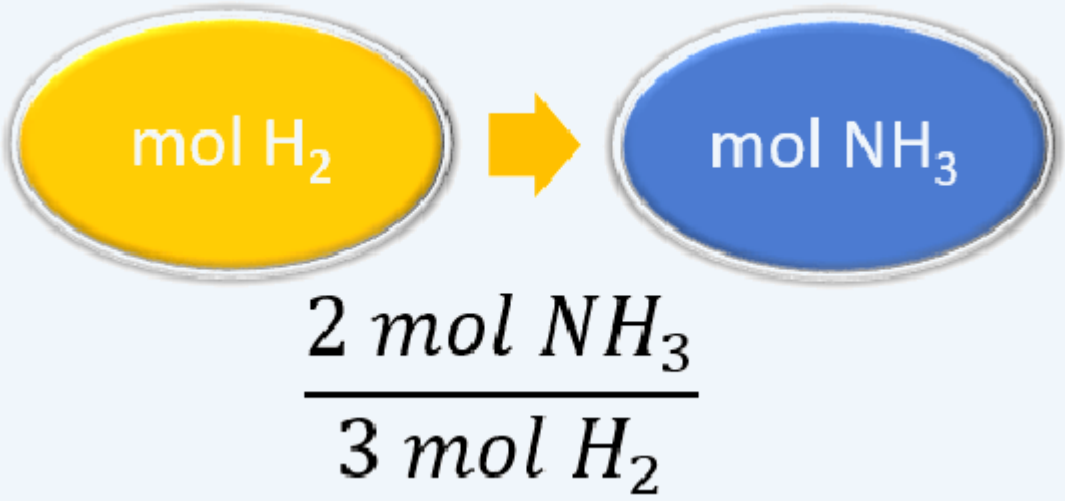
Steps for Problem Solving	Example
Find a balanced equation that describes the reaction	Unbalanced: $H_2 + O_2 \rightarrow H_2O$ Balanced: $\underline{\mathbf{2}}H_2 + O_2 \rightarrow \underline{\mathbf{2}}H_2O$
Identify the "given"information and what the problem is asking you to "find."	Given: moles H_2O Find: moles oxygen
List other known quantities	1 mol O_2 = 2 mol H_2O
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ O_2}{2\ mol\ H_2O}$
Cancel units and calculate.	$\cancel{27.6\ \text{mol}\ H_2O} \times \frac{1\ \text{mol}\ O_2}{\cancel{2\ \text{mol}\ H_2O}} = 13.8\ \text{mol}\ O_2$ To produce 27.6 mol of H_2O, 13.8 mol of O_2 react.
Think about your result.	Since each mole of oxygen produces twice as many moles of water, it makes sense that the produced amount is greater than the reactant amount

Example 5.6.6

How many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen?

Solution

Example 5.6.6: Steps for Problem Solving how many moles of ammonia are produced if 4.20 moles of hydrogen are reacted with an excess of nitrogen

Steps for Problem Solving	Example
Find a balanced equation that describes the reaction	Unbalanced: $N_2 + H_2 \rightarrow NH_3$ Balanced: $N_2 + \underline{3}H_2 \rightarrow \underline{2}NH_3$
Identify the "given"information and what the problem is asking you to "find."	Given: $\mathrm{H_2} = 4.20\ \mathrm{mol}$ Find: mol of $\mathrm{NH_3}$
List other known quantities	3 mol H_2 = 2 mol NH_3
Prepare a concept map and use the proper conversion factor.	 $\frac{2\ mol\ NH_3}{3\ mol\ H_2}$
Cancel units and calculate.	$\cancel{4.20\ \mathrm{mol}\ H_2} \times \frac{2\ \mathrm{mol}\ NH_3}{\cancel{3\ \mathrm{mol}\ H_2}} = 2.80\ \mathrm{mol}\ NH_3$ The reaction of $4.20\ \mathrm{mol}$ of hydrogen with excess nitrogen produces $2.80\ \mathrm{mol}$ of ammonia
Think about your result.	The result corresponds to the 3:2 ratio of hydrogen to ammonia from the balanced equation.

Exercise 5.6.5

a. Given the following balanced chemical equation,

$$\mathrm{C_5H_{12} + 8\,O_2 \longrightarrow 5\,CO_2 + 6\,H_2O}$$

how many moles of H_2O can be formed if 0.0652 mol of C_5H_{12} were to react?

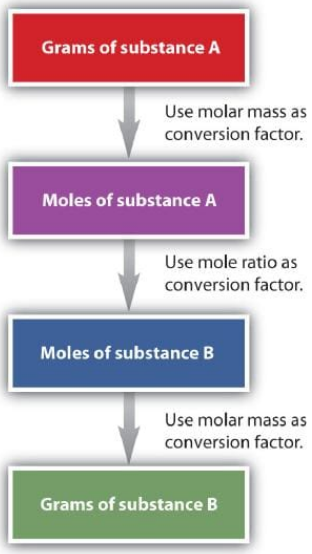
b. Balance the following unbalanced equation and determine how many moles of H_2O are produced when 1.65 mol of NH_3 react.

$$\mathrm{NH_3 + O_2 \longrightarrow N_2 + H_2O}$$

Answer a: available at the end of this section.

Mass Relationships in Chemical Equations

It is a small step from mole-mass calculations to mass-mass calculations. If we start with a known mass of one substance in a chemical reaction (instead of a known number of moles), we can calculate the corresponding masses of other substances in the reaction. The first step in this case is to convert the known mass into moles, using the substance's molar mass as the conversion factor. Then—and only then—we use the balanced chemical equation to construct a conversion factor to convert that quantity to moles of another substance, which in turn can be converted to a corresponding mass. Sequentially, the process is as follows:



This three-part process can be carried out in three discrete steps or combined into a single calculation that contains three conversion factors. The following example illustrates both techniques.

Example 5.6.7: Decomposition of Ammonium Nitrate

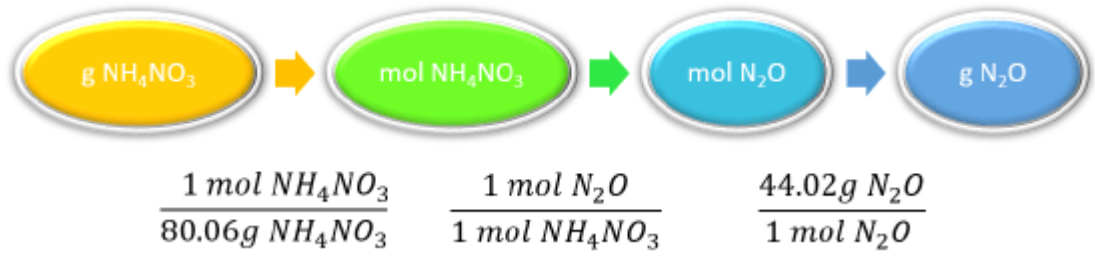
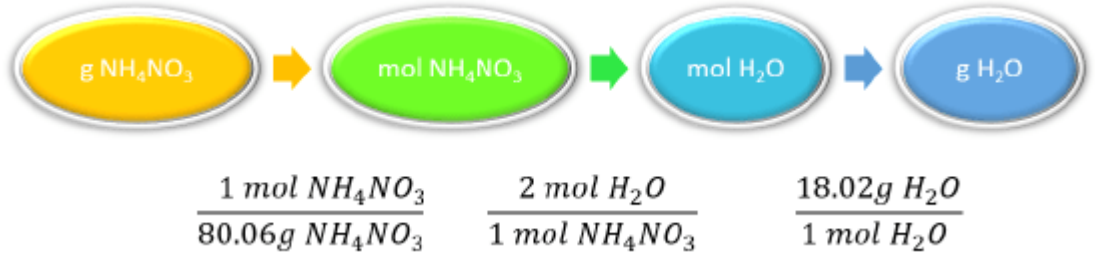
Ammonium nitrate decomposes to dinitrogen monoxide and water according to the following equation.

$$\mathrm{NH_4NO_3}\,(s) \rightarrow \mathrm{N_2O}\,(g) + 2\mathrm{H_2O}\,(l)$$

In a certain experiment, $45.7\ \mathrm{g}$ of ammonium nitrate is decomposed. Find the mass of each of the products formed.

Example 5.6.7: Steps for Problem Solving the Decomposition of Ammonium Nitrate

Steps for Problem Solving	Example 5.6.2
Identify the "given"information and what the problem is asking you to "find."	Given: $45.7\ \mathrm{g}\ \mathrm{NH_4NO_3}$ Find: Mass $\mathrm{N_2O} = ?\ \mathrm{g}$ Mass $\mathrm{H_2O} = ?\ \mathrm{g}$

Steps for Problem Solving	Example 5.6.2
List other known quantities	1 mol $NH_4NO_3 = 80.06\ g/mol$ 1 mol $N_2O = 44.02\ g/mol$ 1 mol $H_2O = 18.02\ g/mol$ 1 mol NH_4NO_3 to 1 mol N_2O to 2 mol H_2O
Prepare two concept maps and use the proper conversion factor.	 $\frac{1\ mol\ NH_4NO_3}{80.06g\ NH_4NO_3}$ $\frac{1\ mol\ N_2O}{1\ mol\ NH_4NO_3}$ $\frac{44.02g\ N_2O}{1\ mol\ N_2O}$  $\frac{1\ mol\ NH_4NO_3}{80.06g\ NH_4NO_3}$ $\frac{2\ mol\ H_2O}{1\ mol\ NH_4NO_3}$ $\frac{18.02g\ H_2O}{1\ mol\ H_2O}$
Cancel units and calculate.	$45.7\ \text{g}\ NH_4NO_3 \times \frac{1\ \text{mol}\ NH_4NO_3}{80.06\ \text{g}\ NH_4NO_3} \times \frac{1\ \text{mol}\ N_2O}{1\ \text{mol}\ NH_4NO_3} \times \frac{44.02\ \text{g}\ N_2O}{1\ \text{mol}\ N_2O} = 25.1\ \text{g}\ N_2O$ $45.7\ \text{g}\ NH_4NO_3 \times \frac{1\ \text{mol}\ NH_4NO_3}{80.06\ \text{g}\ NH_4NO_3} \times \frac{2\ H_2O}{1\ \text{mol}\ NH_4NO_3} \times \frac{18.02\ \text{g}\ H_2O}{1\ \text{mol}\ H_2O} = 20.6\ \text{g}\ H_2O$
Think about your result.	The total mass of the two products is equal to the mass of ammonium nitrate which decomposed, demonstrating the law of conservation of mass. Each answer has three significant figures.

Exercise 5.6.6: Carbon Tetrachloride

Methane can react with elemental chlorine to make carbon tetrachloride (CCl_4). The balanced chemical equation is as follows.

$$CH_4\,(g) + 4\,Cl_2\,(g) \longrightarrow CCl_2\,(l) + 4\,HCl\,(l)$$

How many grams of HCl are produced by the reaction of 100.0g of CH_4?

Answer available at the end of this section.

Summary

- Calculations involving conversions between moles of a substance and the mass of that substance are described.

- The balanced chemical reaction can be used to determine molar and mass relationships between substances.

Contributors and Attributions

- TextMap: The Basics of GOB Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
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Learning Objectives

- Describe the unit *mole*.
- Relate the mole quantity of substance to its mass.

So far, we have been talking about chemical substances in terms of individual atoms and molecules. Yet we do not typically deal with substances an atom or a molecule at a time; we work with millions, billions, and trillions of atoms and molecules at a time. We need a way to deal with macroscopic, rather than microscopic, amounts of matter. We need a unit of amount that relates quantities of substances on a scale that we can interact with.

Chemistry uses a unit called mole. A **mole** (mol) is a number of things equal to the number of atoms in exactly 12 g of carbon-12. Experimental measurements have determined that this number is very large:

$$1 \text{ mol} = 6.02214179 \times 10^{23} \text{ things}$$

Understand that a mole means a specific number of things, just like a dozen means a certain number of things—twelve, in the case of a dozen. But a mole is a much larger number of things. These things can be atoms, or molecules, or eggs; however, in chemistry, we usually use the mole to refer to the amounts of atoms or molecules. Although the number of things in a mole is known to eight decimal places, it is usually fine to use only two or three decimal places in calculations. The numerical value of things in a mole is often called *Avogadro's number* (N_A). *Avogadro's number* is also known as the *Avogadro constant*, after Amadeo Avogadro, an Italian chemist who first proposed its importance.

Example 5.6.1

How many molecules are present in 2.76 mol of H_2O? How many atoms is this?

Solution

The definition of a mole is an equality that can be used to construct a conversion factor. Also, because we know that there are three atoms in each molecule of H_2O, we can also determine the number of atoms in the sample.

$$2.76 \ \cancel{mol\ H_2O} \times \frac{6.022 \times 10^{23} molecules\ H_2O}{\cancel{mol\ H_2O}} = 1.66 \times 10^{24} molecules\ H_2O$$

To determine the total number of atoms, we have

$$1.66 \times 10^{24} \ \cancel{molecules\ H_2O} \times \frac{3\ atoms}{1\ molecule} = 4.99 \times 10^{24}\ atoms$$

Exercise 5.6.1

How many molecules are present in 4.61×10^{-2} mol of O_2?

Answer available at the end of this section.

How big is a mole? It is very large. Suppose you had a mole of dollar bills that need to be counted. If everyone on earth (about 6 billion people) counted one bill per second, it would take about 3.2 million years to count all the bills. A mole of sand would fill a cube about 32 km on a side. A mole of pennies stacked on top of each other would have about the same diameter as our galaxy, the Milky Way. A mole is a lot of things—but atoms and molecules are very tiny. One mole of carbon atoms would make a cube that is 1.74 cm on a side, small enough to carry in your pocket.

Why is the mole unit so important? It represents the link between the microscopic and the macroscopic, especially in terms of mass. *A mole of a substance has the same mass in grams as one unit (atom or molecules) has in atomic mass units*. The mole unit allows us to express amounts of atoms and molecules in visible amounts that we can understand.

For example, we already know that, by definition, a mole of carbon has a mass of exactly 12 g. This means that exactly 12 g of C has 6.022×10^{23} atoms:

$$12 \text{ g C} = 6.022 \times 10^{23} \text{ atoms C}$$

We can use this equality as a conversion factor between the number of atoms of carbon and the number of grams of carbon. How many grams are there, say, in 1.50×10^{25} atoms of carbon? This is a one-step conversion:

$$1.50 \times 10^{25} \cancel{atoms\,C} \times \frac{12.0000\,g\,C}{6.022 \times 10^{23} \cancel{atoms\,C}} = 299\,g\,C$$

But it also goes beyond carbon. Previously we defined atomic and molecular masses as the number of atomic mass units per atom or molecule. Now we can do so in terms of grams. The atomic mass of an element is the number of grams in 1 mol of atoms of that element, while the molecular mass of a compound is the number of grams in 1 mol of molecules of that compound. Sometimes these masses are called **molar masses** to emphasize the fact that they are the mass for 1 mol of things. (The term *molar* is the adjective form of mole and has nothing to do with teeth.)

Here are some examples. The mass of a hydrogen atom is 1.0079 u; the mass of 1 mol of hydrogen atoms is 1.0079 g. Elemental hydrogen exists as a diatomic molecule, H_2. One molecule has a mass of 1.0079 + 1.0079 = 2.0158 u, while 1 mol H_2 has a mass of 2.0158 g. A molecule of H_2O has a mass of about 18.01 u; 1 mol H_2O has a mass of 18.01 g. A single unit of NaCl has a mass of 58.45 u; NaCl has a molar mass of 58.45 g. In each of these moles of substances, there are 6.022×10^{23} units: 6.022×10^{23} atoms of H, 6.022×10^{23} molecules of H_2 and H_2O, 6.022×10^{23} units of NaCl ions. These relationships give us plenty of opportunities to construct conversion factors for simple calculations.

Example 5.6.2: Sugar

What is the molar mass of sugar ($C_6H_{12}O_6$)?

Solution

To determine the molar mass, we simply add the atomic masses of the atoms in the molecular formula; but express the total in grams per mole, not atomic mass units. The masses of the atoms can be taken from the periodic table.

Solutions to Example 5.3.2

6 C = 6 × 12.011	= 72.066
12 H = 12 × 1.0079	= 12.0948
6 O = 6 × 15.999	= 95.994
TOTAL	= 180.155 g/mol

Per convention, the unit *grams per mole* is written as a fraction.

Exercise 5.6.2

What is the molar mass of $\mathrm{AgNO_3}$?

Answer available at the end of this section.

Knowing the molar mass of a substance, we can calculate the number of moles in a certain mass of a substance and vice versa, as these examples illustrate. The molar mass is used as the conversion factor.

Example 5.6.3

What is the mass of 3.56 mol of HgCl2? The molar mass of HgCl2 is 271.49 g/mol.

Solution

Use the molar mass as a conversion factor between moles and grams. Because we want to cancel the mole unit and introduce the gram unit, we can use the molar mass as given:

$$3.56\ \cancel{mol\ HgCl_2} \times \frac{271.49\ g\ HgCl_2}{\cancel{mol\ HgCl_2}} = 967\ g\ HgCl_2$$

Exercise 5.6.3

What is the mass of 33.7 mol of $\mathrm{H_2O}$?

Answer available at the end of this section.

Example 5.6.4

How many moles of H2O are present in 240.0 g of water (about the mass of a cup of water)?

Solution

Use the molar mass of H2O as a conversion factor from mass to moles. The molar mass of water is

(1.0079 + 1.0079 + 15.999) = 18.015 g/mol.

However, because we want to cancel the gram unit and introduce moles, we need to take the reciprocal of this quantity, or 1 mol/18.015 g:

$$240.0\ \cancel{g\ H_2O} \times \frac{1\ mol\ H_2O}{18.015\ \cancel{g\ H_2O}} = 13.32\ mol\ H_2O$$

Exercise 5.6.4

How many moles are present in 35.6 g of H2SO4 (molar mass = 98.08 g/mol)?

Answer available at the end of this section.

Other conversion factors can be combined with the definition of mole—density, for example.

Example 5.6.5

The density of ethanol is 0.789 g/mL. How many moles are in 100.0 mL of ethanol? The molar mass of ethanol is 46.08 g/mol.

Solution

Here, we use density to convert from volume to mass and then use the molar mass to determine the number of moles.

$$100 \cancel{ml}\ ethanol \times \frac{0.789\,g}{\cancel{ml}} \times \frac{1\,mol}{46.08\,\cancel{g}} = 1.71\,mol\,ethanol$$

Exercise 5.6.5

If the density of benzene, C_6H_6, is 0.879 g/mL, how many moles are present in 17.9 mL of benzene?

Answer available at the end of this section.

Summary

The mole is a key unit in chemistry. The molar mass of a substance, in grams, is numerically equal to one atom's or molecule's mass in atomic mass units.

Answers

Exercise 5.6.1: Nitrogen Gas
210 g

Exercise 5.6.2: Calcium Chloride
1.36 x 10^{23} $CaCl_2$ formula units

Exercise 5.6.2
78.32% Ba and 21.67% F

Exercise 5.6.2: Calcium Oxide
420.60 g

Exercise 5.6.3: Nitrogen Gas
210 g

Exercise 5.6.4: Generation of Aluminum Oxide
0.442 mol Al2O3

Exercise 5.6.5
3.14 mol H_2O
$4NH_3 + 3O_2 \rightarrow 2N_2 + 6H_2O$; 2.48 mol H_2O

Exercise 5.6.6: Carbon Tetrachloride
908.7g HCl

Exercise 5.6.1
2.78 × 10^{22} molecules

Exercise 5.6.2
169.87 g/mol

Exercise 5.6.3
607 g

Exercise 5.6.4
0.363 mol

Exercise 5.6.5
0.201 mol

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5.7: Chemical Formula Conversions

Learning Objectives

- Use Avogadro's number to convert to moles and vice versa given the number of particles of an element.
- Know the definition of the mole.
- Determine the formula mass of an ionic or molecular compound.
- Determine the percent composition of each element in a compound from the chemical formula.

When objects are very small, it is often inconvenient or inefficient, or even impossible to deal with the objects one at a time. For these reasons, we often deal with very small objects in groups, and have even invented names for various numbers of objects. The most common of these is "dozen" which refers to 12 objects. We frequently buy objects in groups of 12, like doughnuts or pencils. Even smaller objects such as straight pins or staples are usually sold in boxes of 144, or a dozen dozen. A group of 144 is called a "gross".

This problem of dealing with things that are too small to operate with as single items also occurs in chemistry. Atoms and molecules are too small to see, let alone to count or measure. Chemists needed to select a group of atoms or molecules that would be convenient to operate with.

Avogadro's Number:Counting Atoms

Owing to their tiny size, atoms and molecules cannot be counted by direct observation. But much as we do when "counting" beans in a jar, we can estimate the number of particles in a sample of an element or compound if we have some idea of the volume occupied by each particle and the volume of the container. Once this has been done, we know the number of formula units (to use the most general term for any combination of atoms we wish to define) in any arbitrary weight of the substance. The number will of course depend both on the formula of the substance and on the weight of the sample. However, if we consider a weight of substance that is the same as its formula (molecular) weight expressed in grams, we have only one number to know: **Avogadro's number**.

Avogadro's number

Avogadro's number is known to ten significant digits:

$$N_A = 6.022141527 \times 10^{23}.$$

However, you only need to know it to three significant figures:

$$N_A \approx 6.02 \times 10^{23}. \tag{5.7.1}$$

So 6.02×10^{23} of *what*? Well, of anything you like: apples, stars in the sky, burritos. However, the only *practical* use for N_A is to have a more convenient way of expressing the huge numbers of the tiny particles such as atoms or molecules that we deal with in chemistry. Avogadro's number is a *collective number*, just like a dozen. Students can think of 6.02×10^{23} as the "chemist's dozen".

Before getting into the use of Avogadro's number in problems, take a moment to convince yourself of the reasoning embodied in the following examples.

Things to understand about Avogadro's number

- It is a *number*, just as is "dozen", and thus is *dimensionless*.
- It is a *huge* number, far greater in magnitude than we can visualize
- Its practical use is limited to counting tiny things like atoms, molecules, "formula units", electrons, or photons.

- The value of N_A can be known only to the precision that the number of atoms in a measurable weight of a substance can be estimated. Because large numbers of atoms cannot be counted directly, a variety of ingenious indirect measurements have been made involving such things as Brownian motion and X-ray scattering.

The Mole: "A Dozen Eggs and a Mole of Sugar, Please"

The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly $6.02214076 \times 10^{23}$ particles, e.g., atoms, molecules, ions or electrons. The current definition was adopted in November 2018, revising its old definition based on the number of atoms in 12 grams of carbon-12 (^{12}C) (the isotope of carbon with relative atomic mass 12 Daltons by definition).

It is not obvious why eggs come in dozens rather than 10s or 14s, or why a ream of paper contains 500 sheets rather than 400 or 600. The definition of a mole—that is, the decision to base it on 12 g of carbon-12—is also arbitrary. The important point is that 1 mole of carbon—or of anything else, whether atoms, compact discs, or houses—always has the same number of objects: 6.02×10^{23}.

The Mole

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video $5.7.1$ *How big is a mole?*

Converting Between Number of Atoms to Moles and Vice Versa

We can use Avogadro's number as a conversion factor, or ratio, in dimensional analysis problems. If we are given the number of atoms of an element X, we can convert it into moles of by using the relationship

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

An example on the use of Avogadro's number as a conversion factor is given below for carbon.

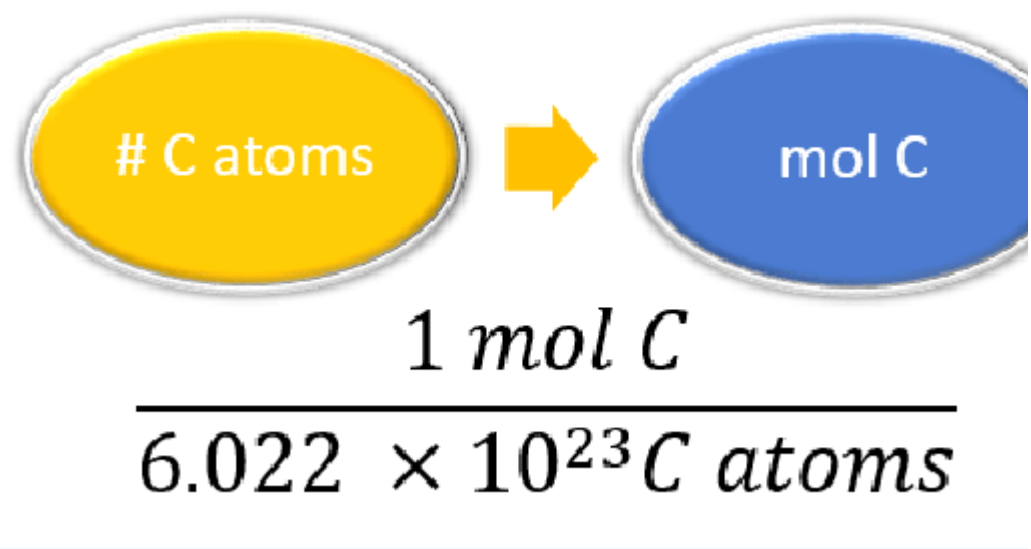
Example $5.7.1$: Moles of Carbon

The element carbon exists in two primary forms: graphite and diamond. How many moles of carbon atoms is 4.72×10^{24} atoms of carbon?

Solution

Example $5.7.1$*:Steps for Problem Solving how many moles of carbon atoms is* 4.72×10^{24} *atoms of carbon?*

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: 4.72×10^{24} C atoms Find: mol C

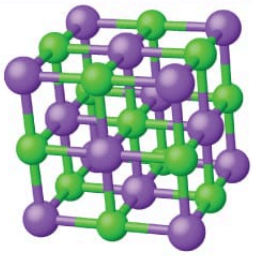
Steps for Problem Solving	Example
List other known quantities	$1\ mol = 6.022 \times 10^{23}$ C atoms
Prepare a concept map and use the proper conversion factor.	 $\frac{1\ mol\ C}{6.022 \times 10^{23} C\ atoms}$
Cancel units and calculate.	$4.72 \times 10^{24}\ \cancel{\text{C atoms}} \times \frac{1\ \text{mol C}}{6.02 \times 10^{23}\ \cancel{\text{C atoms}}} = 7.84\ \text{mol C}$
Think about your result.	The given number of carbon atoms was greater than Avogadro's number,so the number of moles of C atoms mole. Since Avogadro's number is a measured quantity with three significant figures, the result of the calcul to three significant figures

Formula Mass

One skill needed in future chapters is the ability to determine the mass of the formula of various chemical substances. This quantity is called the formula mass. The **formula mass** is obtained by adding the masses of each individual atom in the formula of the substance. Because a proper formula is electrically neutral (with no net electrons gained or lost), the ions can be considered atoms for the purpose of calculating the formula mass.

Let us start by calculating the formula mass of sodium chloride (NaCl). This formula mass is the sum of the atomic masses of one sodium atom and one chlorine atom, which we find from the periodic table; here, we use the masses to two decimal places:

Element	Quantity		Average atomic mass (amu)		Subtotal
Na	1	×	22.99	=	22.99
Cl	1	×	35.45	=	35.45
				Formula mass	58.44

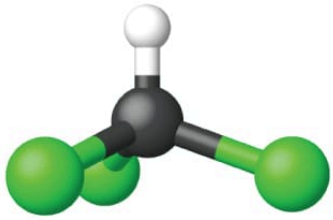


A table and diagram are shown. The table is made up of six columns and four rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space and “Subtotal (a m u).” The first column contains the symbols “N a”, “C l,” and a merged cell. The merged cell runs the length of the first five columns. The second column contains the numbers “1” and “1” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “22.99” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “22.99,” “35.45,” and “58.44.” There is a thick black line below the number “35.45.” The merged cell under the first five columns reads “Formula mass.” To the left of the table is a diagram of a chemical structure. The diagram shows green and purple spheres placed in an alternating pattern, making up the corners of eight stacked cubes to form one larger cube. The green spheres are slightly smaller than the purple spheres.

To two decimal places, the formula mass of NaCl is 58.44 amu.

For covalent substances, the formula represents the numbers and types of atoms composing a single molecule of the substance; therefore, the formula mass may be correctly referred to as a **molecular mass**. Consider chloroform ($CHCl_3$), a covalent compound once used as a surgical anesthetic and now primarily used in the production of tetrafluoroethylene, the building block for the “anti-stick” polymer, Teflon. The molecular formula of chloroform indicates that a single molecule contains one carbon atom, one hydrogen atom, and three chlorine atoms. The average molecular mass of a chloroform molecule is therefore equal to the sum of the average atomic masses of these atoms.

Element	Quantity		Average atomic mass (amu)		Subtotal (amu)
C	1	×	12.01	=	12.01
H	1	×	1.008	=	1.008
Cl	3	×	35.45	=	106.35
				Molecular mass	119.37



A table and diagram are shown. The table is made up of six columns and five rows. The header row reads: “Element,” “Quantity,” a blank space, “Average atomic mass (a m u),” a blank space, and “Subtotal (a m u).” The first column contains the symbols “C,” “H,” “C l” and a blank, merged cell that runs the width of the first five columns. The second column contains the numbers “1,” “1,” and “3” as well as the merged cell. The third column contains the multiplication symbol in each cell except for the last, merged cell. The fourth column contains the numbers “12.01,” “1.008,” and “35.45” as well as the merged cell. The fifth column contains the symbol “=” in each cell except for the last, merged cell. The sixth column contains the values “12.01,” “1.008,” “106.35,” and “119.37.” There is a thick black line below the number 106.35. The merged cell under the first five columns reads “Molecular mass.” To the left of the table is a diagram of a molecule. Three green spheres are attached to a slightly smaller black sphere, which is also attached to a smaller white sphere. The green spheres lie beneath and to the sides of the black sphere while the white sphere is located straight up from the black sphere.

To two decimal places, the formula mass of $CHCl_3$ is 119.37 amu.

For ionic compounds with polyatomic ions, the sum must include the number and mass of each atom in the formula for the polyatomic ion. as shown in the example below for aluminum sulfate, $Al_2(SO_4)_3$.

Example 5.7.2 Formula Mass for an Ionic Compound

Aluminum sulfate, $Al_2(SO_4)_3$, is an ionic compound that is used in the manufacture of paper and in various water purification processes. What is the formula mass (amu) of this compound?

Solution

The formula for this compound indicates it contains Al^{3+} and SO_4^{2-} ions combined in a 2:3 ratio. For purposes of computing a formula mass, it is helpful to rewrite the formula in the simpler format, $Al_2S_3O_{12}$. Following the approach outlined above, the formula mass for this compound is calculated as follows:

<div data-mt-source="1""A

" style="width: 791px; height: 210px;" width="791px" height="210px" src="/@api/deki/files/56171/CNX_Chem_03_01_alsulfatemass_img.jpg">

The formula mass for $Al_2(SO_4)_3$, is 342.14 amu.

Exercise 5.7.1

Use the atomic masses (rounded to two decimal places) to determine the formula mass for each ionic compound.

a. TiO_2
b. AgBr
c. $Au(NO_3)_3$
d. $Fe_3(PO_4)_2$

Answer

a. 79.87 amu

b. 187.77 amu

c. 383.0 amu

Percent Composition of a Compound from a Chemical Formula

The percent composition of a compound can also be determined from the formula of the compound. The subscripts in the formula are first used to calculate the mass of each element in one mole of the compound. That is divided by the molar mass of the compound and multiplied by 100%.

$$\%\text{ by mass} = \frac{\text{mass of element in 1 mol}}{\text{molar mass of compound}} \times 100\%$$

The percent composition of a given compound is always the same as long as the compound is pure.

Example 5.7.3

Dichlorine heptoxide (Cl_2O_7) is a highly reactive compound used in some organic synthesis reactions. Calculate the percent composition of dichlorine heptoxide.

Solution

Example $5.7.3$:*Steps for Problem Solving calculate the percent composition of dichlorine heptoxide* (Cl_2O_7).

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given : Cl_2O_7 Find: % Composition (% Cl and %O)
List other known quantities	Mass of Cl in 1 mol Cl_2O_7 , 2 Cl : 2 x 35.45 g = 70.90 g Mass of O in 1 mol Cl_2O_7 , 7 O: 7 x 16.00 g = 112.00 g Molar mass of Cl_2O_7 = 182.90 g/mol
Cancel units and calculate.	$\%\mathrm{Cl} = \frac{70.90\ \mathrm{g\ Cl}}{182.90\ \mathrm{g}} \times 100\% = 38.76\%\ \mathrm{Cl}$ $\%\ \mathrm{O} = \frac{112.00\ \mathrm{g\ O}}{182.90\ \mathrm{g}} \times 100\% = 61.24\%\ \mathrm{O}$ Calculate the percent by mass of each element by dividing the mass of that element in 1 mole of the co molar mass of the compound and multiplying by 100%.
Think about your result.	The percentages add up to 100%.

Percent composition can also be used to determine the mass of a certain element that is contained in any mass of a compound. In the previous sample problem, it was found that the percent composition of dichlorine heptoxide is $38.76\%\ \mathrm{Cl}$ and $61.24\%\ \mathrm{O}$. Suppose that you needed to know the masses of chlorine and oxygen present in a $12.50\ \mathrm{g}$ sample of dichlorine heptoxide. You can set up a conversion factor based on the percent by mass of each element.

$$12.50\ \mathrm{g\ Cl_2O_7} \times \frac{38.76\ \mathrm{g\ Cl}}{100\ \mathrm{g\ Cl_2O_7}} = 4.845\ \mathrm{g\ Cl}$$

$$12.50\ \mathrm{g\ Cl_2O_7} \times \frac{61.24\ \mathrm{g\ O}}{100\ \mathrm{g\ Cl_2O_7}} = 7.655\ \mathrm{g\ O}$$

The sum of the two masses is $12.50\ \mathrm{g}$, the mass of the sample size.

Exercise $5.7.2$

Barium fluoride is a transparent crystal that can be found in nature as the mineral frankdicksonite. Determine the percent composition of barium fluoride.

Answer a: available at the end of this section.

Summary

- The **mole** (symbol: **mol**) is the base unit of amount of substance ("number of substance") in the International System of Units or System International (SI), defined as exactly 6.02214076×10^{23} particles, e.g., atoms, molecules, ions or electrons.
- Avogadro's number is related to moles of any substance X as follows:

$$1 \text{ mol X} = 6.022 \times 10^{23} \text{ X atoms.}$$

- Formula masses of ionic and molecular compounds can be determined from the masses of the atoms in their formulas.
- Processes are described for calculating the percent composition of a compound based on the chemical formula.

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- Stephen Lower, Professor Emeritus (Simon Fraser U.) Chem1 Virtual Textbook
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- Wikipedia

Answers

Exercise 5.7.2

78.32% Ba and 21.67% F

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5.8: Concentrations

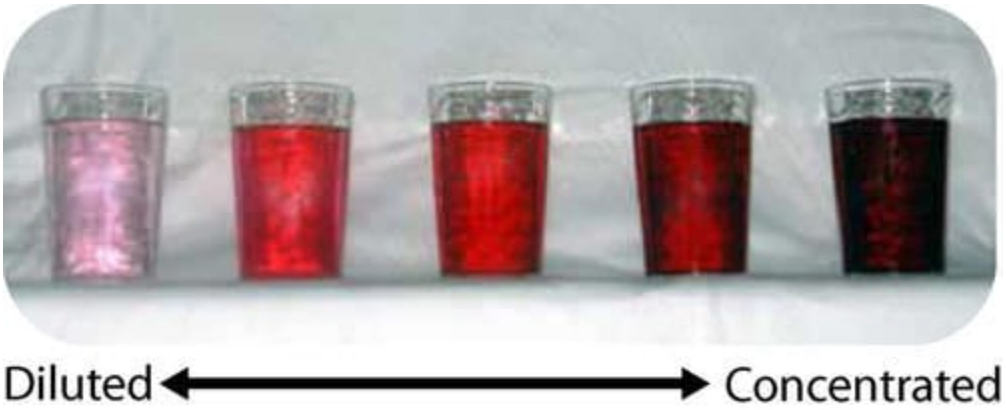
Learning Objectives

- Learn some terminology involving solutions.
- Express the amount of solute in a solution in various concentration units.

When one substance dissolves into another, a solution is formed. A **solution** *is a homogenous mixture consisting of a solute dissolved into a solvent*. The **solute** *is the substance that is being dissolved,* while the **solvent** *is the dissolving medium*. Solutions can be formed with many different types and forms of solutes and solvents. In this chapter, we will focus on solution where the solvent is water. An **aqueous solution** *is water that contains one or more dissolved substance*. The dissolved substances in an aqueous solution may be solids, gases, or other liquids.

In order to be a true solution, a mixture must be stable. When sugar is fully dissolved into water, it can stand for an indefinite amount of time, and the sugar will not settle out of the solution. Further, if the sugar-water solution is passed through a filter, it will remain with the water. This is because the dissolved particles in a solution are very small, usually less than $1\ \text{nm}$ in diameter. Solute particles can be atoms, ions, or molecules, depending on the type of substance that has been dissolved. Water typically dissolves most ionic compounds and polar molecules. Nonpolar molecules, such as those found in grease or oil, do not dissolve in water.

Solution Concentrations



Solutions of a red dye in water from the most dilute (on the left)
to the most concentrated (on the right).

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. What you have to do is take the frozen juice from inside these containers and usually empty it into 3 or 4 times the container size full of water to mix with the juice concentrate and make your container of juice. Therefore, you are diluting the concentrated juice. When we talk about solute and solvent, the concentrated solution

has a lot of solute versus the dilute solution that would have a smaller amount of solute.

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

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but

these are relative terms whose meanings depend on various factors. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution and we know exactly by how much.

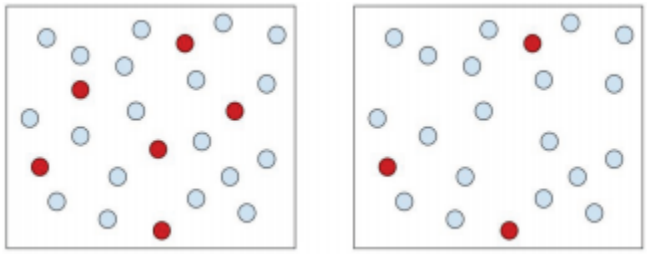


Figure 5.8.2 *The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).*

Molarity

One way of expressing concentration is to give the number of moles of solute per unit volume of solution. Of all the quantitative measures of concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

$$\text{molarity} = \frac{\text{number of moles of solute}}{\text{number of liters of solution}} \tag{5.8.1}$$

The symbol for molarity is M or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[\mathrm{Ag^+}]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to calculate with but the most difficult to make in the lab. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

It is important to remember that “mol” in this expression refers to moles of solute and that “L” refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is therefore

$$\frac{1.5\ \text{mol NaCl}}{0.500\ \text{L solution}} = 3.0\ \text{M NaCl}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” You would read as "a 3.00 ***molar*** sodium chloride solution," meaning that there are 3.00 moles of NaOH dissolved per one liter of solution.

> Be sure to note that molarity is calculated as the total volume of the **entire** solution, not just volume of solvent! The solute contributes to total volume.

If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L?

First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ \cancel{g HCl} \times \frac{1\ mol\ HCl}{36.5\ \cancel{g HCl}} = 0.614\ mol\ HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M\ =\ \frac{0.614\ mol\ HCl}{1.56 L\ solution} = 0.394\ M HCl$$

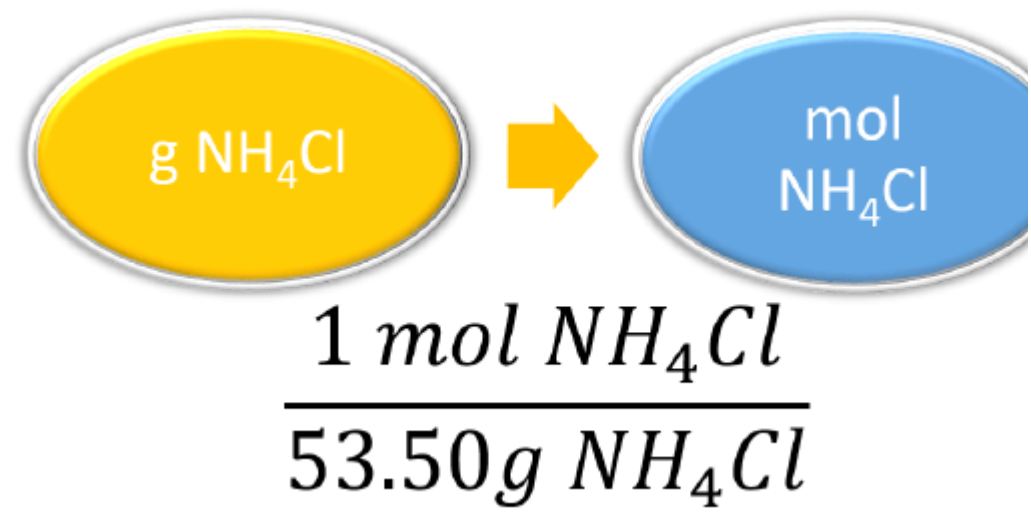
Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

Example 5.8.1

A solution is prepared by dissolving $42.23\ \text{g}$ of NH_4Cl into enough water to make $500.0\ \text{mL}$ of solution. Calculate its molarity.

Solution

Example 5.8.1: Steps for Problem Solving Molarity

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: Mass $= 42.23\ \text{g}\ \text{NH}_4\text{Cl}$ Volume solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$ Find: Molarity = ? M
List other known quantities.	Molar mass $\text{NH}_4\text{Cl} = 53.50\ \text{g/mol}$
Plan the problem.	1. The mass of the ammonium chloride is first converted to moles.  2. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters $\text{M} = \frac{\text{mol NH}_4\text{Cl}}{\text{L solution}}$

Steps for Problem Solving	Example
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23\ \cancel{\text{g NH}_4\text{Cl}} \times \frac{1\ \text{mol}\ \cancel{\text{NH}_4\text{Cl}}}{53.50\ \text{g NH}_4\text{Cl}} = 0.7893\ \text{mol NH}_4\text{Cl}$ $\frac{0.7893\ \text{mol NH}_4\text{Cl}}{0.5000\ \text{L solution}} = 1.579\ \text{M}$
Think about your result.	The molarity is $1.579\ \text{M}$, meaning that a liter of the solution would contain $1.579\ \text{mol}\ \text{NH}_4\text{Cl}$. Four significant appropriate.

Exercise 5.8.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

Exercise 5.8.2

What is the concentration, in mol/L, where $137\ \text{g}$ of NaCl has been dissolved in enough water to make $500\ \text{mL}$ of solution?

Answer available at the end of this section.

Figure 5.8.1 illustrates how a solution is prepared with the measured amount of a solid solute and the desired final volume of the solution.

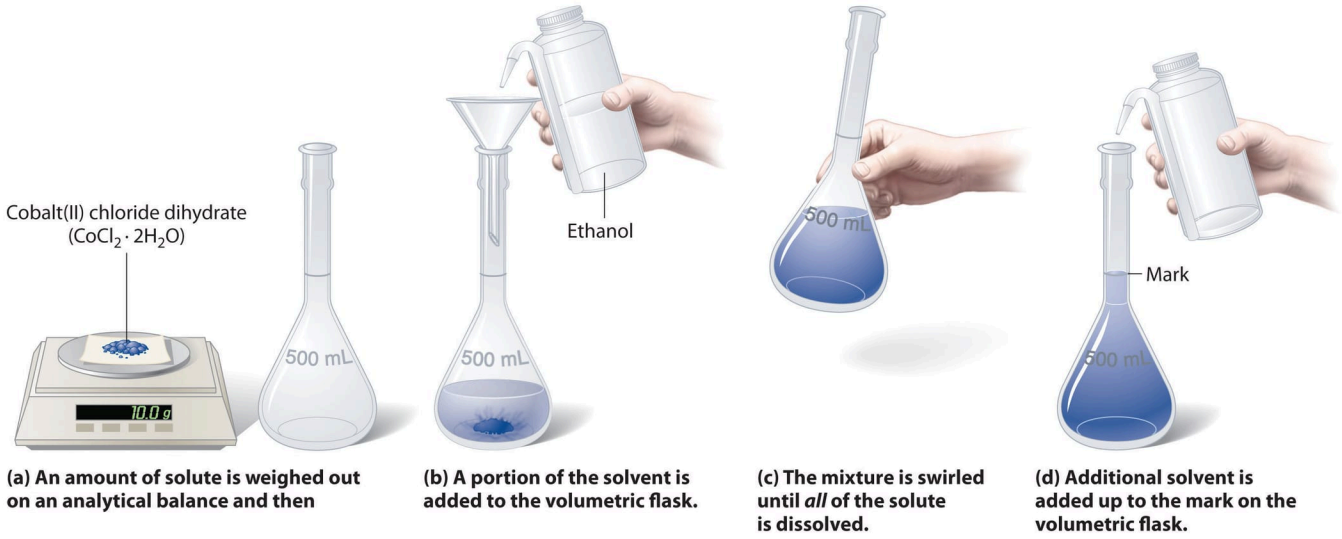


Figure 5.8.3: Preparation of a solution of known concentration using a solid solute.

(a) An amount of solute is weighted out on an analytical balance, (b) a portion of the solvent is added to the volumetric flask. (c) the mixture is swirled until all of the solute is dissolved. (d) additional solvent is added up to the mark on the volumetric flask.

Percent Concentrations

There are several ways of expressing the concentration of a solution by using a percentage. The percent can further be determined in one of two ways: (1) the ratio of the mass of the solute divided by the mass of the solution or (2) the ratio of the volume of the solute divided by the volume of the solution.

Percent by Volume

The percentage of solute in a solution can more easily be determined by volume when the solute and solvent are both liquids. The volume of the solute divided by the volume of the solution expressed as a percent yields the percent by volume $\left(\frac{\text{volume}}{\text{volume}}\right)$ of the solution. If a solution is made by adding $40\ \text{mL}$ of ethanol to $20\ \text{mL}$ of water, the percent by volume is:

$$\text{Percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{5.8.4}$$

$$= \frac{40\ \text{mL ethanol}}{240\ \text{mL solution}} \times 100\% \tag{5.8.5}$$

$$= 16.7\%\ \text{ethanol} \tag{5.8.6}$$

Frequently, ingredient labels on food products and medicines have amounts listed as percentages (see figure below).

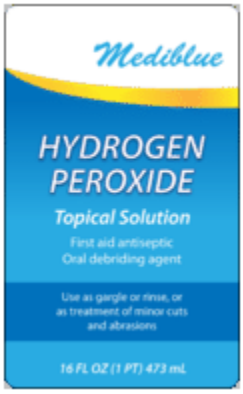


Figure $5.8.4$ Hydrogen peroxide is commonly sold as *a* 3% *by volume solution for use as a disinfectant.*

Percent by Mass

The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\ \text{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving $25.0\ \text{g}$ of sugar into $100.0\ \text{g}$ of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0\ \text{g sugar}}{125.0\ \text{g solution}} \times 100\% = 20.0\%\ \text{sugar}$$

Example 5.8.2

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$\%\ \mathrm{m/m} = \frac{36.5\ \mathrm{g}}{355\ \mathrm{g}} \times 100\% = 10.3\%$$

Exercise 5.8.3

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer available at the end of this section.

Using Mass Percent in Calculations

Sometimes you may want to make up a particular mass of solution of a given percent by mass and need to calculate what mass of the solute to use. Using mass percent as a conversion can be useful in this type of problem. The mass percent can be expressed as a conversion factor in the form $\frac{g\ \text{solute}}{100\ \text{gsolution}}$ or $\frac{100\ \text{gsolution}}{g\ \text{solute}}$

For example, if you need to make $3000.0\ \mathrm{g}$ of a 5.00% solution of sodium chloride, the mass of solute needs to be determined.

Solution

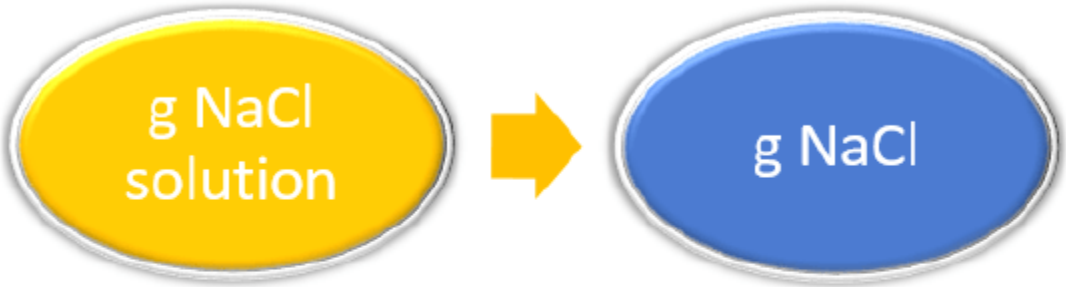
Given: 3000.0 g NaCl solution

5.00% NaCl solution

Find: mass of solute = ? g NaCl

Other known quantities: 5.00 g NaCl is to 100 g solution

The appropriate conversion factor (based on the given mass percent) can be used follows:



$$\frac{5.00\ g\ NaCl}{100\ g\ NaCl\ solution}$$

To solve for the mass of NaCl, the given mass of solution is multiplied by the conversion factor.

$$gNaCl = 3,000.0\ \cancel{g\ NaCl\ solution} \times \frac{5.00\ g\ NaCl}{100\ \cancel{g\ NaCl\ solution}} = 150.0g\ NaCl$$

You would need to weigh out $150\ \text{g}$ of NaCl and add it to $2850\ \text{g}$ of water. Notice that it was necessary to subtract the mass of the $\text{NaCl}\ (150\ \text{g})$ from the mass of solution $(3000\ \text{g})$ to calculate the mass of the water that would need to be added.

Exercise 5.8.4

What is the amount (in g) of hydrogen peroxide (H_2O_2) needed to make a 6.00 kg , 3.00 % (by mass) H_2O_2 solution?

Answer

180. g H_2O_2

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Concentration of solutions can be more precisely expressed in terms of molarity, percent by volume, and percent by mass.

Contributors and Attributions

-
- Allison Soult, Ph.D. (Department of Chemistry, University of Kentucky)
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- Henry Agnew (UC Davis)

Learning Objective

- Learn to determine specific concentrations with several common units.

Rather than qualitative terms (Section 11.2 - Definitions), we need quantitative ways to express the amount of solute in a solution; that is, we need specific units of concentration. In this section, we will introduce several common and useful units of concentration.

Molarity (M) is defined as the number of moles of solute divided by the number of liters of solution:

$$molarity = \frac{moles\ of\ solute}{liters\ of\ solution}$$

which can be simplified as

$$M = \frac{mol}{L},\ or\ mol/L$$

As with any mathematical equation, if you know any two quantities, you can calculate the third, unknown, quantity.

For example, suppose you have 0.500 L of solution that has 0.24 mol of NaOH dissolved in it. The concentration of the solution can be calculated as follows:

$$molarity = \frac{0.24\ mol\ NaOH}{0.500 L} = 0.48\ M\ NaOH$$

The concentration of the solution is 0.48 M, which is spoken as "zero point forty-eight molarity" or "zero point forty-eight molar." If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L? First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ \cancel{g HCl} \times \frac{1\ mol\ Hcl}{36.5\ \cancel{g HCl}} = 0.614\ M\ HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M = \frac{0.614\ mol\ HCl}{1.56 L} = 0.394\ M$$

Example 5.8.1

What is the molarity of a solution made when 32.7 g of NaOH are dissolved to make 445 mL of solution?

Solution

To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$445\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}} = 0.445\ L$$

Now we convert the amount of solute to moles, using the molar mass of NaOH, which is 40.0 g/mol:

$$32.7\ \cancel{g NaOH} \times \frac{1\ mol\ NaOH}{40.0\ \cancel{g NaOH}} = 0.818\ mol\ NaOH$$

Now we can use the definition of molarity to determine the molar concentration:

$$M = \frac{0.818\ mol\ NaOH}{0.445L} = 1.84\ M\ NaOH$$

Exercise 5.8.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

The definition of molarity can be used to determine the amount of solute or the volume of solution, if the other information is given. Example 4 illustrates this situation.

Example 5.8.2

How many moles of solute are present in 0.108 L of a 0.887 M NaCl solution?

Solution

We know the volume and the molarity; we can use the definition of molarity to mathematically solve for the amount in moles. Substituting the quantities into the definition of molarity:

$$0.887\ M = \frac{mol\ NaCl}{0.108L}$$

We multiply the 0.108 L over to the other side of the equation and multiply the units together; "molarity × liters" equals moles, according to the definition of molarity. So

mol NaCl = (0.887 M)(0.108 L) = 0.0958 mol

Exercise 5.8.2

How many moles of solute are present in 225 mL of a 1.44 M $CaCl_2$ solution?

Answer available at the end of this section.

If you need to determine volume, remember the rule that the unknown quantity must be by itself and in the numerator to determine the correct answer. Thus, rearrangement of the definition of molarity is required.

Example 5.8.3

What volume of a 2.33 M $NaNO_3$ solution is needed to obtain 0.222 mol of solute?

Solution

Using the definition of molarity, we have

$$2.33\ M = \frac{0.222mol}{L}$$

To solve for the number of liters, we bring the 2.33 M over to the right into the denominator, and the number of liters over to the left in the numerator. We now have

$$L = \frac{0.222mol}{2.33\ M}$$

Dividing, **the volume is 0.0953 L = 95.3 mL.**

Exercise 5.8.3

What volume of a 0.570 M K_2SO_4 solution is needed to obtain 0.872 mol of solute?

Answer available at the end of this section.

A similar unit of concentration is **molality** (*m*), which is defined as the number of moles of solute per kilogram of solvent, not per liter of solution:

$$molality = \frac{moles\ solute}{kilograms\ solvent}$$

Mathematical manipulation of molality is the same as with molarity.

Another way to specify an amount is **percentage composition by mass** (or *mass percentage*, % m/m). It is defined as follows:

$$\%m/m = \frac{mass\ of\ solute}{mass\ of\ entire\ sample} \times 100\%$$

It is not uncommon to see this unit used on commercial products (see Figure 5.8.1 - Concentration in Commercial Applications).



Figure 5.8.1 Concentration in Commercial Applications © Thinkstock. *The percentage of urea in this package is 5% m/m, meaning that there are 5 g of urea per 100 g of product.*

Example 5.8.4

What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

Solution

Using the definition of mass percentage, we have

$$\%m/m = \frac{87.9\ g\ Fe}{113\ g\ sample} \times 100\% = 77.8\%\ Fe$$

Exercise 5.8.4

What is the mass percentage of H_2O_2 in a solution with 1.67 g of H_2O_2 in a 55.5 g sample?

Answer available at the end of this section.

Related concentration units are **parts per thousand (ppth), parts per million (ppm)** and **parts per billion (ppb).** Parts per thousand is defined as follows:

$$ppth = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1000$$

There are similar definitions for parts per million and parts per billion:

$$ppm = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000$$

and

$$ppb = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000,000$$

Each unit is used for progressively lower and lower concentrations. The two masses must be expressed in the same unit of mass, so conversions may be necessary.

Example 5.8.5

If there are 0.6 g of Pb present in 277 g of solution, what is the Pb concentration in parts per thousand?

Solution

Use the definition of parts per thousand to determine the concentration. Substituting

$$\frac{0.6g\ Pb}{277g\ solution} \times 1000 = 2.17\ ppth$$

Exercise 5.8.5

If there are 0.551 mg of As in 348 g of solution, what is the As concentration in ppm?

Answer available at the end of this section.

As with molarity and molality, algebraic rearrangements may be necessary to answer certain questions.

Example 5.8.6

The concentration of Cl^- ion in a sample of H_2O is 15.0 ppm. What mass of Cl^- ion is present in 240.0 mL of H_2O, which has a density of 1.00 g/mL?

Solution

First, use the density of H_2O to determine the mass of the sample:

$$240.0 \cancel{mL} \times \frac{1.00\ g}{\cancel{mL}} = 240.0\ g$$

Now we can use the definition of ppm:

$$15.0\,ppm\ = \frac{mass\ of\ solute}{240.0\ g\ solution} \times 1,000,000$$

Rearranging to solve for the mass of solute,

$$mass\ solute = \frac{(15.0\,ppm)(240.0\ g\ solution)}{1,000,000} = 0.0036g = 3.6\,mg$$

Exercise 5.8.6

The concentration of Fe^{3+} ion in a sample of H_2O is 335.0 ppm. What mass of Fe^{3+} ion is present in 3,450 mL of H_2O, which has a density of 1.00 g/mL?

Answer available at the end of this section.

For ionic solutions, we need to differentiate between the concentration of the salt versus the concentration of each individual ion. Because the ions in ionic compounds go their own way when a compound is dissolved in a solution, the resulting concentration of the ion may be different from the concentration of the complete salt. For example, if 1 M NaCl were prepared, the solution could also be described as a solution of 1 M $Na^+(aq)$ and 1 M $Cl^-(aq)$ because there is one Na^+ ion and one Cl^- ion per formula unit of the salt. However, if the solution were 1 M $CaCl_2$, there are two $Cl^-(aq)$ ions for every formula unit dissolved, so the concentration of $Cl^-(aq)$ would be 2 M, not 1 M.

In addition, the total ion concentration is the sum of the individual ion concentrations. Thus for the 1 M NaCl, the total ion concentration is 2 M; for the 1 M $CaCl_2$, the total ion concentration is 3 M.

Key Takeaway

- Quantitative units of concentration include molarity, molality, mass percentage, parts per thousand (ppth), parts per million (ppm), and parts per billion (ppb).

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Mathematical manipulation of molality is the same as with molarity.

Another way to specify an amount is **percentage composition by mass** (or *mass percentage*, % m/m). It is defined as follows:

$$\%m/m\;=\;\frac{mass\;of\;solute}{mass\;of\;entire\;sample}\times 100\%$$

It is not uncommon to see this unit used on commercial products (see Figure 5.8.1 - Concentration in Commercial Applications).



Figure 5.8.1 Concentration in Commercial Applications © Thinkstock. *The percentage of urea in this package is 5% m/m, meaning that there are 5 g of urea per 100 g of product.*

Example 5.8.4

What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

Solution

Using the definition of mass percentage, we have

$$\%m/m = \frac{87.9\,g\,Fe}{113\,g\,sample} \times 100\% = 77.8\%\,Fe$$

Exercise 5.8.4

What is the mass percentage of H_2O_2 in a solution with 1.67 g of H_2O_2 in a 55.5 g sample?

Answer available at the end of this section.

Related concentration units are **parts per thousand (ppth), parts per million (ppm)** and **parts per billion (ppb).** Parts per thousand is defined as follows:

$$ppth = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1000$$

There are similar definitions for parts per million and parts per billion:

$$ppm = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000$$

and

$$ppb = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000,000$$

Each unit is used for progressively lower and lower concentrations. The two masses must be expressed in the same unit of mass,

so conversions may be necessary.

Example 5.8.5

If there are 0.6 g of Pb present in 277 g of solution, what is the Pb concentration in parts per thousand?

Solution

Use the definition of parts per thousand to determine the concentration. Substituting

$$\frac{0.6g\,Pb}{277g\,solution} \times 1000 = 2.17\,ppth$$

Exercise 5.8.5

If there are 0.551 mg of As in 348 g of solution, what is the As concentration in ppm?

Answer available at the end of this section.

As with molarity and molality, algebraic rearrangements may be necessary to answer certain questions.

Example 5.8.6

The concentration of Cl^- ion in a sample of H_2O is 15.0 ppm. What mass of Cl^- ion is present in 240.0 mL of H_2O, which has a density of 1.00 g/mL?

Solution

First, use the density of H_2O to determine the mass of the sample:

$$240.0\,\cancel{mL} \times \frac{1.00\,g}{\cancel{mL}} = 240.0\,g$$

Now we can use the definition of ppm:

$$15.0\,ppm\ = \frac{mass\ of\ solute}{240.0\,g\ solution} \times 1,000,000$$

Rearranging to solve for the mass of solute,

$$mass\ solute = \frac{(15.0\,ppm)(240.0\,g\ solution)}{1,000,000} = 0.0036g = 3.6\,mg$$

Exercise 5.8.6

The concentration of Fe^{3+} ion in a sample of H_2O is 335.0 ppm. What mass of Fe^{3+} ion is present in 3,450 mL of H_2O, which has a density of 1.00 g/mL?

Answer available at the end of this section.

For ionic solutions, we need to differentiate between the concentration of the salt versus the concentration of each individual ion. Because the ions in ionic compounds go their own way when a compound is dissolved in a solution, the resulting concentration of the ion may be different from the concentration of the complete salt. For example, if 1 M NaCl were prepared, the solution could

also be described as a solution of 1 M $Na^+(aq)$ and 1 M $Cl^-(aq)$ because there is one Na^+ ion and one Cl^- ion per formula unit of the salt. However, if the solution were 1 M $CaCl_2$, there are two $Cl^-(aq)$ ions for every formula unit dissolved, so the concentration of $Cl^-(aq)$ would be 2 M, not 1 M.

In addition, the total ion concentration is the sum of the individual ion concentrations. Thus for the 1 M NaCl, the total ion concentration is 2 M; for the 1 M $CaCl_2$, the total ion concentration is 3 M.

Key Takeaway

- Quantitative units of concentration include molarity, molality, mass percentage, parts per thousand (ppth), parts per million (ppm), and parts per billion (ppb).

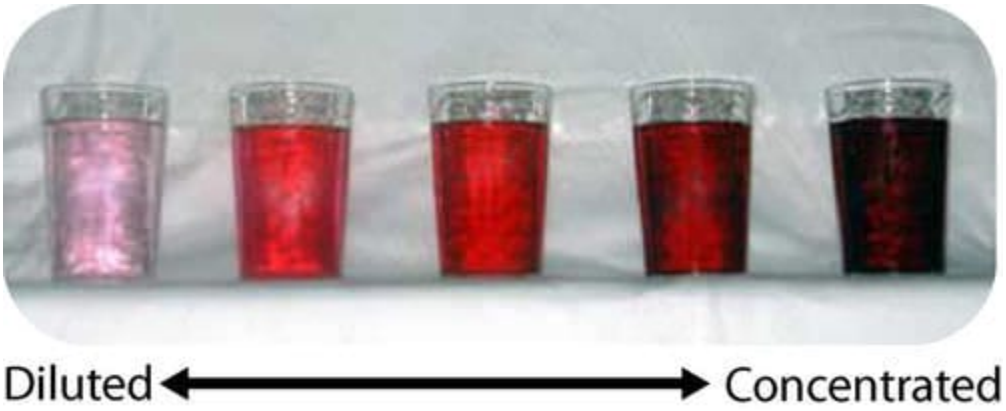
Learning Objectives

- Learn some terminology involving solutions.
- Express the amount of solute in a solution in various concentration units.

When one substance dissolves into another, a solution is formed. A **solution** *is a homogenous mixture consisting of a solute dissolved into a solvent*. The **solute** *is the substance that is being dissolved,* while the **solvent** *is the dissolving medium*. Solutions can be formed with many different types and forms of solutes and solvents. In this chapter, we will focus on solution where the solvent is water. An **aqueous solution** *is water that contains one or more dissolved substance*. The dissolved substances in an aqueous solution may be solids, gases, or other liquids.

In order to be a true solution, a mixture must be stable. When sugar is fully dissolved into water, it can stand for an indefinite amount of time, and the sugar will not settle out of the solution. Further, if the sugar-water solution is passed through a filter, it will remain with the water. This is because the dissolved particles in a solution are very small, usually less than $1\ \text{nm}$ in diameter. Solute particles can be atoms, ions, or molecules, depending on the type of substance that has been dissolved. Water typically dissolves most ionic compounds and polar molecules. Nonpolar molecules, such as those found in grease or oil, do not dissolve in water.

Solution Concentrations



Solutions of a red dye in water from the most dilute (on the left) to the most concentrated (on the right).

A **concentrated** solution is one in which there is a large amount of solute in a given amount of solvent. A **dilute** solution is one in which there is a small amount of solute in a given amount of solvent. A dilute solution is a concentrated solution that has been, in essence, watered down. Think of the frozen juice containers you buy in the grocery store. What you have to do is take the frozen juice from inside these containers and usually empty it into 3 or 4 times the container size full of water to mix with the juice concentrate and make your container of juice. Therefore, you are diluting the concentrated juice. When we talk about solute and

solvent, the concentrated solution

has a lot of solute versus the dilute solution that would have a smaller amount of solute.

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

To define a solution precisely, we need to state its concentration: how much solute is dissolved in a certain amount of solvent. Words such as *dilute* or *concentrated* are used to describe solutions that have a little or a lot of dissolved solute, respectively, but these are relative terms whose meanings depend on various factors. When we say that vinegar is 5% acetic acid in water, we are giving the concentration. If we said the mixture was 10% acetic acid, this would be more concentrated than the vinegar solution and we know exactly by how much.

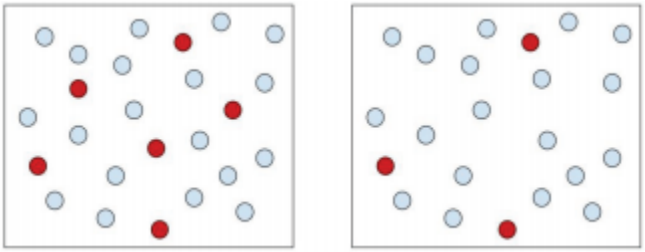


Figure $5.8.2$ *The solution on the left is more concentrated than the solution on the right because there is a greater ratio of solute (red balls) to solvent (blue balls) particles. The solution particles are closer together. The solution on the right is more dilute (less concentrated). (CC-SA-BY-3.0 Tracy Poulsen).*

Molarity

One way of expressing concentration is to give the number of moles of solute per unit volume of solution. Of all the quantitative measures of concentration, molarity is the one used most frequently by chemists. **Molarity** is defined as the number of moles of solute per liter of solution.

Label 'defMolarity' multiply defined

The symbol for molarity is M or moles/liter. Chemists also use square brackets to indicate a reference to the molarity of a substance. For example, the expression $[\mathrm{Ag^+}]$ refers to the molarity of the silver ion in solution. Solution concentrations expressed in molarity are the easiest to calculate with but the most difficult to make in the lab. Such concentration units are useful for discussing chemical reactions in which a solute is a product or a reactant. Molar mass can then be used as a conversion factor to convert amounts in moles to amounts in grams.

It is important to remember that “mol” in this expression refers to moles of solute and that “L” refers to liters of solution. For example, if you have 1.5 mol of NaCl dissolved in 0.500 L of solution, its molarity is therefore

$$\frac{1.5\ \mathrm{mol\ NaCl}}{0.500\ \mathrm{L\ solution}} = 3.0\ \mathrm{M\ NaCl}$$

Sometimes (aq) is added when the solvent is water, as in “3.0 M NaCl (aq).” You would read as "a 3.00 ***molar*** sodium chloride solution," meaning that there are 3.00 moles of NaOH dissolved per one liter of solution.

> Be sure to note that molarity is calculated as the total volume of the **entire** solution, not just volume of solvent! The solute contributes to total volume.

If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L?

First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\,\cancel{g HCl} \times \frac{1\,mol\,HCl}{36.5\,\cancel{g HCl}} = 0.614\,mol\,HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M\ =\ \frac{0.614\,mol\,HCl}{1.56L\,solution} = 0.394\,MHCl$$

Before a molarity concentration can be calculated, the amount of the solute must be expressed in moles, and the volume of the solution must be expressed in liters, as demonstrated in the following example.

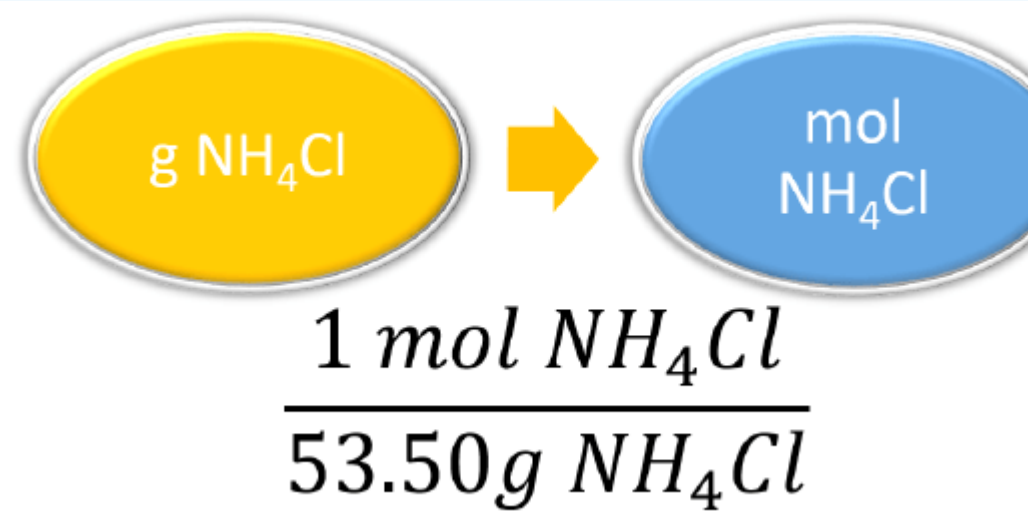
Example 5.8.1

A solution is prepared by dissolving $42.23\ \text{g}$ of NH_4Cl into enough water to make $500.0\ \text{mL}$ of solution. Calculate its molarity.

Solution

Example 5.8.1: Steps for Problem Solving Molarity

Steps for Problem Solving	Example
Identify the "given"information and what the problem is asking you to "find."	Given: Mass $= 42.23\ \text{g}\ \text{NH}_4\text{Cl}$ Volume solution $= 500.0\ \text{mL} = 0.5000\ \text{L}$ Find: Molarity = ? M
List other known quantities.	Molar mass $\text{NH}_4\text{Cl} = 53.50\ \text{g/mol}$

Steps for Problem Solving	Example
Plan the problem.	1. The mass of the ammonium chloride is first converted to moles.  2. Then the molarity is calculated by dividing by liters. Note the given volume has been converted to liters $\mathrm{M} = \frac{\text{mol } \mathrm{NH_4Cl}}{\text{L solution}}$
Cancel units and calculate.	Now substitute the known quantities into the equation and solve. $42.23\ \cancel{\text{g } \mathrm{NH_4Cl}} \times \frac{1\ \text{mol } \cancel{\mathrm{NH_4Cl}}}{53.50\ \text{g } \mathrm{NH_4Cl}} = 0.7893\ \text{mol } \mathrm{NH_4Cl}$ $\frac{0.7893\ \text{mol } \mathrm{NH_4Cl}}{0.5000\ \text{L solution}} = 1.579\ \mathrm{M}$
Think about your result.	The molarity is $1.579\ \mathrm{M}$, meaning that a liter of the solution would contain $1.579\ \text{mol } \mathrm{NH_4Cl}$. Four significant appropriate.

Exercise 5.8.1

What is the molarity of a solution made when 66.2 g of C6H12O6 are dissolved to make 235 mL of solution?

Answer available at the end of this section.

Exercise 5.8.2

What is the concentration, in $\mathrm{mol/L}$, where $137\ \mathrm{g}$ of NaCl has been dissolved in enough water to make $500\ \mathrm{mL}$ of solution?

Answer available at the end of this section.

Figure 5.8.1 illustrates how a solution is prepared with the measured amount of a solid solute and the desired final volume of the solution.

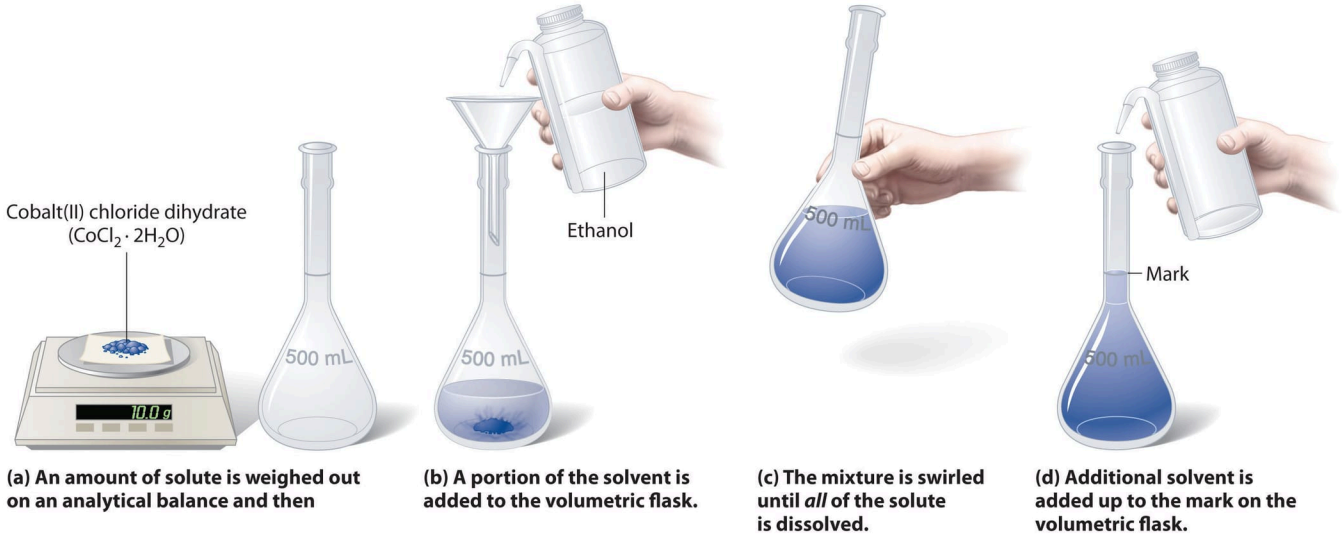


Figure 5.8.3: Preparation of a solution of known concentration using a solid solute.

(a) An amount of solute is weighted out on an analytical balance, (b) a portion of the solvent is added to the volumetric flask. (c) the mixture is swirled until all of the solute is dissolved. (d) additional solvent is added up to the mark on the volumetric flask.

Percent Concentrations

There are several ways of expressing the concentration of a solution by using a percentage. The percent can further be determined in one of two ways: (1) the ratio of the mass of the solute divided by the mass of the solution or (2) the ratio of the volume of the solute divided by the volume of the solution.

Percent by Volume

The percentage of solute in a solution can more easily be determined by volume when the solute and solvent are both liquids. The volume of the solute divided by the volume of the solution expressed as a percent yields the percent by volume $\left(\frac{\text{volume}}{\text{volume}}\right)$ of the solution. If a solution is made by adding $40\ \text{mL}$ of ethanol to $20\ \text{mL}$ of water, the percent by volume is:

$$\text{Percent by volume} = \frac{\text{volume of solute}}{\text{volume of solution}} \times 100\% \tag{5.8.9}$$

$$= \frac{40\ \text{mL ethanol}}{240\ \text{mL solution}} \times 100\% \tag{5.8.10}$$

$$= 16.7\%\ \text{ethanol} \tag{5.8.11}$$

Frequently, ingredient labels on food products and medicines have amounts listed as percentages (see figure below).

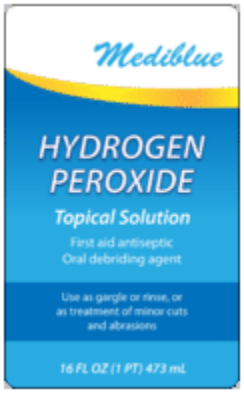


Figure 5.8.4 Hydrogen peroxide is commonly sold as *a 3% by volume solution for use as a disinfectant.*

Percent by Mass

The mass/mass percent (% m/m) is defined as the mass of a solute divided by the mass of a solution times 100:

$$\%\ \text{m/m} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

mass of solution = mass of solute + mass solvent

If you can measure the masses of the solute and the solution, determining the mass/mass percent is easy. Each mass must be expressed in the same units to determine the proper concentration.

Suppose that a solution was prepared by dissolving $25.0\ \text{g}$ of sugar into $100.0\ \text{g}$ of water.

The mass of the solution is

mass of solution = 25.0g sugar + 100.0g water = 125.0 g

The percent by mass would be calculated by:

$$\text{Percent by mass} = \frac{25.0\ \text{g sugar}}{125.0\ \text{g solution}} \times 100\% = 20.0\%\ \text{sugar}$$

Example 5.8.2

A saline solution with a mass of 355 g has 36.5 g of NaCl dissolved in it. What is the mass/mass percent concentration of the solution?

Solution

We can substitute the quantities given in the equation for mass/mass percent:

$$\%\ \text{m/m} = \frac{36.5\ \text{g}}{355\ \text{g}} \times 100\% = 10.3\%$$

Exercise 5.8.3

A dextrose (also called D-glucose, $C_6H_{12}O_6$) solution with a mass of 2.00×10^2 g has 15.8 g of dextrose dissolved in it. What is the mass/mass percent concentration of the solution?

Answer available at the end of this section.

Using Mass Percent in Calculations

Sometimes you may want to make up a particular mass of solution of a given percent by mass and need to calculate what mass of the solute to use. Using mass percent as a conversion can be useful in this type of problem. The mass percent can be expressed as a conversion factor in the form $\frac{g\ \text{solute}}{100\ \text{gsolution}}$ or $\frac{100\ \text{gsolution}}{g\ \text{solute}}$

For example, if you need to make $3000.0\ \text{g}$ of a 5.00% solution of sodium chloride, the mass of solute needs to be determined.

Solution

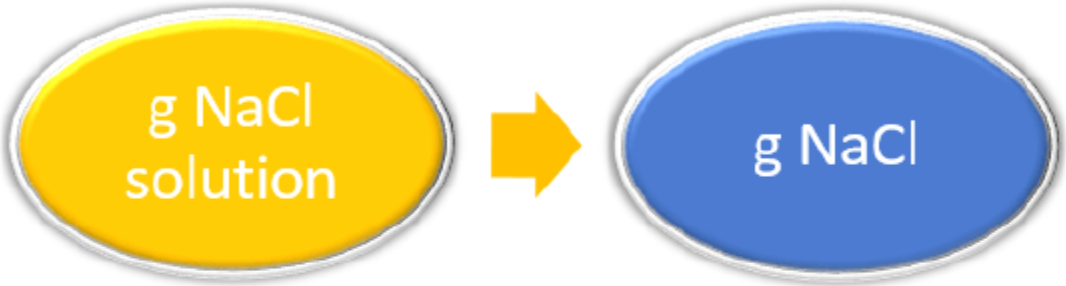
Given: 3000.0 g NaCl solution

5.00% NaCl solution

Find: mass of solute = ? g NaCl

Other known quantities: 5.00 g NaCl is to 100 g solution

The appropriate conversion factor (based on the given mass percent) can be used follows:



$$\frac{5.00\ g\ NaCl}{100\ g\ NaCl\ solution}$$

To solve for the mass of NaCl, the given mass of solution is multiplied by the conversion factor.

$$gNaCl = 3,000.0\ \cancel{g\ NaCl\ solution} \times \frac{5.00\ g\ NaCl}{100\ \cancel{g\ NaCl\ solution}} = 150.0g\ NaCl$$

You would need to weigh out $150\ \text{g}$ of NaCl and add it to $2850\ \text{g}$ of water. Notice that it was necessary to subtract the mass of the $\text{NaCl}\ (150\ \text{g})$ from the mass of solution $(3000\ \text{g})$ to calculate the mass of the water that would need to be added.

Exercise 5.8.4

What is the amount (in g) of hydrogen peroxide (H_2O_2) needed to make a 6.00 kg , 3.00 % (by mass) H_2O_2 solution?

Answer

180. g H_2O_2

Summary

- Solutions are composed of a solvent (major component) and a solute (minor component).
- Concentration of solutions can be more precisely expressed in terms of molarity, percent by volume, and percent by mass.

Contributors and Attributions

-
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Learning Objective

- Learn to determine specific concentrations with several common units.

Rather than qualitative terms (Section 11.2 - Definitions), we need quantitative ways to express the amount of solute in a solution; that is, we need specific units of concentration. In this section, we will introduce several common and useful units of concentration.

Molarity (M) is defined as the number of moles of solute divided by the number of liters of solution:

$$molarity = \frac{moles\ of\ solute}{liters\ of\ solution}$$

which can be simplified as

$$M = \frac{mol}{L},\ or\ mol/L$$

As with any mathematical equation, if you know any two quantities, you can calculate the third, unknown, quantity.

For example, suppose you have 0.500 L of solution that has 0.24 mol of NaOH dissolved in it. The concentration of the solution can be calculated as follows:

$$molarity = \frac{0.24\ mol\ NaOH}{0.500 L} = 0.48\ M\ NaOH$$

The concentration of the solution is 0.48 M, which is spoken as "zero point forty-eight molarity" or "zero point forty-eight molar." If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L? First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\ \cancel{g HCl} \times \frac{1\ mol\ Hcl}{36.5\ \cancel{g HCl}} = 0.614\ M\ HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M = \frac{0.614\ mol\ HCl}{1.56 L} = 0.394\ M$$

Example 5.8.1

What is the molarity of a solution made when 32.7 g of NaOH are dissolved to make 445 mL of solution?

Solution

To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$445\ \cancel{mL} \times \frac{1\ L}{1000\ \cancel{mL}} = 0.445\ L$$

Now we convert the amount of solute to moles, using the molar mass of NaOH, which is 40.0 g/mol:

$$32.7\ \cancel{g NaOH} \times \frac{1\ mol\ NaOH}{40.0\ \cancel{g NaOH}} = 0.818\ mol\ NaOH$$

Now we can use the definition of molarity to determine the molar concentration:

$$M = \frac{0.818\ mol\ NaOH}{0.445 L} = 1.84\ M\ NaOH$$

Exercise 5.8.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

The definition of molarity can be used to determine the amount of solute or the volume of solution, if the other information is given. Example 4 illustrates this situation.

Example 5.8.2

How many moles of solute are present in 0.108 L of a 0.887 M NaCl solution?

Solution

We know the volume and the molarity; we can use the definition of molarity to mathematically solve for the amount in moles. Substituting the quantities into the definition of molarity:

$$0.887\ M = \frac{mol\ NaCl}{0.108 L}$$

We multiply the 0.108 L over to the other side of the equation and multiply the units together; "molarity × liters" equals moles, according to the definition of molarity. So

mol NaCl = (0.887 M)(0.108 L) = 0.0958 mol

Exercise 5.8.2

How many moles of solute are present in 225 mL of a 1.44 M $CaCl_2$ solution?

Answer available at the end of this section.

If you need to determine volume, remember the rule that the unknown quantity must be by itself and in the numerator to determine the correct answer. Thus, rearrangement of the definition of molarity is required.

Example 5.8.3

What volume of a 2.33 M $NaNO_3$ solution is needed to obtain 0.222 mol of solute?

Solution

Using the definition of molarity, we have

$$2.33\ M = \frac{0.222 mol}{L}$$

To solve for the number of liters, we bring the 2.33 M over to the right into the denominator, and the number of liters over to the left in the numerator. We now have

$$L = \frac{0.222 mol}{2.33\ M}$$

Dividing, **the volume is 0.0953 L = 95.3 mL.**

Exercise 5.8.3

What volume of a 0.570 M K_2SO_4 solution is needed to obtain 0.872 mol of solute?

Answer available at the end of this section.

A similar unit of concentration is **molality** (*m*), which is defined as the number of moles of solute per kilogram of solvent, not per liter of solution:

$$molality = \frac{moles\ solute}{kilograms\ solvent}$$

Mathematical manipulation of molality is the same as with molarity.

Another way to specify an amount is **percentage composition by mass** (or *mass percentage*, % m/m). It is defined as follows:

$$\%m/m = \frac{mass\ of\ solute}{mass\ of\ entire\ sample} \times 100\%$$

It is not uncommon to see this unit used on commercial products (see Figure 5.8.1 - Concentration in Commercial Applications).



Figure 5.8.1 Concentration in Commercial Applications © Thinkstock. *The percentage of urea in this package is 5% m/m, meaning that there are 5 g of urea per 100 g of product.*

Example 5.8.4

What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

Solution

Using the definition of mass percentage, we have

$$\%m/m = \frac{87.9\,g\,Fe}{113\,g\,sample} \times 100\% = 77.8\%\,Fe$$

Exercise 5.8.4

What is the mass percentage of H_2O_2 in a solution with 1.67 g of H_2O_2 in a 55.5 g sample?

Answer available at the end of this section.

Related concentration units are **parts per thousand (ppth), parts per million (ppm)** and **parts per billion (ppb).** Parts per thousand is defined as follows:

$$ppth = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1000$$

There are similar definitions for parts per million and parts per billion:

$$ppm = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000$$

and

$$ppb = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000,000$$

Each unit is used for progressively lower and lower concentrations. The two masses must be expressed in the same unit of mass, so conversions may be necessary.

Example 5.8.5

If there are 0.6 g of Pb present in 277 g of solution, what is the Pb concentration in parts per thousand?

Solution

Use the definition of parts per thousand to determine the concentration. Substituting

$$\frac{0.6g\,Pb}{277g\,solution} \times 1000 = 2.17\,ppth$$

Exercise 5.8.5

If there are 0.551 mg of As in 348 g of solution, what is the As concentration in ppm?

Answer available at the end of this section.

As with molarity and molality, algebraic rearrangements may be necessary to answer certain questions.

Example 5.8.6

The concentration of Cl^- ion in a sample of H_2O is 15.0 ppm. What mass of Cl^- ion is present in 240.0 mL of H_2O, which has a density of 1.00 g/mL?

Solution

First, use the density of H_2O to determine the mass of the sample:

$$240.0 \cancel{mL} \times \frac{1.00\ g}{\cancel{mL}} = 240.0\ g$$

Now we can use the definition of ppm:

$$15.0\,ppm\ = \frac{mass\ of\ solute}{240.0\ g\ solution} \times 1,000,000$$

Rearranging to solve for the mass of solute,

$$mass\ solute = \frac{(15.0\,ppm)(240.0\ g\ solution)}{1,000,000} = 0.0036g = 3.6\,mg$$

Exercise 5.8.6

The concentration of Fe^{3+} ion in a sample of H_2O is 335.0 ppm. What mass of Fe^{3+} ion is present in 3,450 mL of H_2O, which has a density of 1.00 g/mL?

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In addition, the total ion concentration is the sum of the individual ion concentrations. Thus for the 1 M NaCl, the total ion concentration is 2 M; for the 1 M $CaCl_2$, the total ion concentration is 3 M.

Key Takeaway

- Quantitative units of concentration include molarity, molality, mass percentage, parts per thousand (ppth), parts per million (ppm), and parts per billion (ppb).

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which can be simplified as

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As with any mathematical equation, if you know any two quantities, you can calculate the third, unknown, quantity.

For example, suppose you have 0.500 L of solution that has 0.24 mol of NaOH dissolved in it. The concentration of the solution can be calculated as follows:

$$molarity = \frac{0.24\,mol\,NaOH}{0.500L} = 0.48\,M\,NaOH$$

The concentration of the solution is 0.48 M, which is spoken as "zero point forty-eight molarity" or "zero point forty-eight molar." If the quantity of the solute is given in mass units, you must convert mass units to mole units before using the definition of molarity to calculate concentration. For example, what is the molar concentration of a solution of 22.4 g of HCl dissolved in 1.56 L? First, convert the mass of solute to moles using the molar mass of HCl (36.5 g/mol):

$$22.4\,\cancel{gHCl} \times \frac{1\,mol\,Hcl}{36.5\,\cancel{gHCl}} = 0.614\,M\,HCl$$

Now we can use the definition of molarity to determine a concentration:

$$M = \frac{0.614\,mol\,HCl}{1.56L} = 0.394\,M$$

Example 5.8.1

What is the molarity of a solution made when 32.7 g of NaOH are dissolved to make 445 mL of solution?

Solution

To use the definition of molarity, both quantities must be converted to the proper units. First, convert the volume units from milliliters to liters:

$$445\,\cancel{mL} \times \frac{1\,L}{1000\,\cancel{mL}} = 0.445\,L$$

Now we convert the amount of solute to moles, using the molar mass of NaOH, which is 40.0 g/mol:

$$32.7\,\cancel{gNaOH} \times \frac{1\,mol\,NaOH}{40.0\,\cancel{gNaOH}} = 0.818\,mol\,NaOH$$

Now we can use the definition of molarity to determine the molar concentration:

$$M = \frac{0.818\,mol\,NaOH}{0.445L} = 1.84\,M\,NaOH$$

Exercise 5.8.1

What is the molarity of a solution made when 66.2 g of $C_6H_{12}O_6$ are dissolved to make 235 mL of solution?

Answer available at the end of this section.

The definition of molarity can be used to determine the amount of solute or the volume of solution, if the other information is given. Example 4 illustrates this situation.

Example 5.8.2

How many moles of solute are present in 0.108 L of a 0.887 M NaCl solution?

Solution

We know the volume and the molarity; we can use the definition of molarity to mathematically solve for the amount in moles. Substituting the quantities into the definition of molarity:

$$0.887\,M = \frac{mol\ NaCl}{0.108 L}$$

We multiply the 0.108 L over to the other side of the equation and multiply the units together; "molarity × liters" equals moles, according to the definition of molarity. So

mol NaCl = (0.887 M)(0.108 L) = 0.0958 mol

Exercise 5.8.2

How many moles of solute are present in 225 mL of a 1.44 M $CaCl_2$ solution?

Answer available at the end of this section.

If you need to determine volume, remember the rule that the unknown quantity must be by itself and in the numerator to determine the correct answer. Thus, rearrangement of the definition of molarity is required.

Example 5.8.3

What volume of a 2.33 M $NaNO_3$ solution is needed to obtain 0.222 mol of solute?

Solution

Using the definition of molarity, we have

$$2.33\,M = \frac{0.222 mol}{L}$$

To solve for the number of liters, we bring the 2.33 M over to the right into the denominator, and the number of liters over to the left in the numerator. We now have

$$L = \frac{0.222 mol}{2.33\,M}$$

Dividing, **the volume is 0.0953 L = 95.3 mL.**

Exercise 5.8.3

What volume of a 0.570 M K_2SO_4 solution is needed to obtain 0.872 mol of solute?

Answer available at the end of this section.

A similar unit of concentration is **molality** (*m*), which is defined as the number of moles of solute per kilogram of solvent, not per liter of solution:

$$molality = \frac{moles\ solute}{kilograms\ solvent}$$

Mathematical manipulation of molality is the same as with molarity.

Another way to specify an amount is **percentage composition by mass** (or *mass percentage*, % m/m). It is defined as follows:

$$\%m/m = \frac{mass\ of\ solute}{mass\ of\ entire\ sample} \times 100\%$$

It is not uncommon to see this unit used on commercial products (see Figure 5.8.1 - Concentration in Commercial Applications).



Figure 5.8.1 Concentration in Commercial Applications © Thinkstock. *The percentage of urea in this package is 5% m/m, meaning that there are 5 g of urea per 100 g of product.*

Example 5.8.4

What is the mass percentage of Fe in a piece of metal with 87.9 g of Fe in a 113 g sample?

Solution

Using the definition of mass percentage, we have

$$\%m/m = \frac{87.9\,g\,Fe}{113\,g\,sample} \times 100\% = 77.8\%\,Fe$$

Exercise 5.8.4

What is the mass percentage of H_2O_2 in a solution with 1.67 g of H_2O_2 in a 55.5 g sample?

Answer available at the end of this section.

Related concentration units are **parts per thousand (ppth), parts per million (ppm)** and **parts per billion (ppb).** Parts per thousand is defined as follows:

$$ppth = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1000$$

There are similar definitions for parts per million and parts per billion:

$$ppm = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000$$

and

$$ppb = \frac{mass\ of\ solute}{mass\ of\ sample} \times 1,000,000,000$$

Each unit is used for progressively lower and lower concentrations. The two masses must be expressed in the same unit of mass, so conversions may be necessary.

Example 5.8.5

If there are 0.6 g of Pb present in 277 g of solution, what is the Pb concentration in parts per thousand?

Solution

Use the definition of parts per thousand to determine the concentration. Substituting

$$\frac{0.6g\ Pb}{277g\ solution} \times 1000 = 2.17\ ppth$$

Exercise 5.8.5

If there are 0.551 mg of As in 348 g of solution, what is the As concentration in ppm?

Answer available at the end of this section.

As with molarity and molality, algebraic rearrangements may be necessary to answer certain questions.

Example 5.8.6

The concentration of Cl^- ion in a sample of H_2O is 15.0 ppm. What mass of Cl^- ion is present in 240.0 mL of H_2O, which has a density of 1.00 g/mL?

Solution

First, use the density of H_2O to determine the mass of the sample:

$$240.0\ \cancel{mL} \times \frac{1.00\ g}{\cancel{mL}} = 240.0\ g$$

Now we can use the definition of ppm:

$$15.0\ ppm = \frac{mass\ of\ solute}{240.0\ g\ solution} \times 1,000,000$$

Rearranging to solve for the mass of solute,

$$mass\ solute = \frac{(15.0\ ppm)(240.0\ g\ solution)}{1,000,000} = 0.0036g = 3.6\ mg$$

Exercise 5.8.6

The concentration of Fe^{3+} ion in a sample of H_2O is 335.0 ppm. What mass of Fe^{3+} ion is present in 3,450 mL of H_2O, which has a density of 1.00 g/mL?

Answer available at the end of this section.

For ionic solutions, we need to differentiate between the concentration of the salt versus the concentration of each individual ion. Because the ions in ionic compounds go their own way when a compound is dissolved in a solution, the resulting concentration of the ion may be different from the concentration of the complete salt. For example, if 1 M NaCl were prepared, the solution could also be described as a solution of 1 M $Na^+(aq)$ and 1 M $Cl^-(aq)$ because there is one Na^+ ion and one Cl^- ion per formula unit of the salt. However, if the solution were 1 M $CaCl_2$, there are two $Cl^-(aq)$ ions for every formula unit dissolved, so the concentration of $Cl^-(aq)$ would be 2 M, not 1 M.

In addition, the total ion concentration is the sum of the individual ion concentrations. Thus for the 1 M NaCl, the total ion concentration is 2 M; for the 1 M $CaCl_2$, the total ion concentration is 3 M.

Key Takeaway

- Quantitative units of concentration include molarity, molality, mass percentage, parts per thousand (ppth), parts per million (ppm), and parts per billion (ppb).

Answers

Exercise 5.8.1
1.57 M $C_6H_{12}O_6$

Exercise 5.8.2
4.69 M NaCl

Exercise 5.8.3
7.90 %

Exercise 5.8.1
1.57 M

Exercise 5.8.2
0.324 mol

Exercise 5.8.3
1.53 L

Exercise 5.8.4
3.01%

Exercise 5.8.5
1.58 ppm

Exercise 5.8.6
1.16 g

Exercise 5.8.1
1.57 M

Exercise 5.8.2
0.324 mol

Exercise 5.8.3
1.53 L

Exercise 5.8.4
3.01%

Exercise 5.8.5
1.58 ppm

Exercise 5.8.6
1.16 g

Exercise 5.8.1
1.57 M $C_6H_{12}O_6$

Exercise 5.8.2
4.69 M NaCl

Exercise 5.8.3
7.90 %

Exercise 5.8.1
1.57 M

Exercise 5.8.2
0.324 mol

Exercise 5.8.3
1.53 L

Exercise 5.8.4
3.01%

Exercise 5.8.5
1.58 ppm

Exercise 5.8.6
1.16 g

Exercise 5.8.1
1.57 M

Exercise 5.8.2
0.324 mol

Exercise 5.8.3
1.53 L

Exercise 5.8.4
3.01%

Exercise 5.8.5
1.58 ppm

Exercise 5.8.6
1.16 g

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- 11.3: Quantitative Units of Concentration by Anonymous is licensed CC BY-NC-SA 3.0. Original source: https://2012books.lardbucket.org/books/beginning-chemistry.

5.9: Concentration Conversions

Learning Objective

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108 \cancel{L\,NaCl} \times \frac{0.887\,mol\,NaCl}{\cancel{L\,NaCl}} = 0.0958\,mol\,NaCl$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 5.9.1

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88 \cancel{mol\,CuSO_4} \times \frac{1\,L}{2.35\,\cancel{mol}} = 2.08\,L\,of\,solution$$

Exercise 5.9.1

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer: available at the end of this section.

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

Example 5.9.2

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765 \cancel{L} \times \frac{1.93 \cancel{mol\ NaOH}}{\cancel{L\ solution}} \times \frac{40.0 g\ NaOH}{1 \cancel{mol\ NaOH}} = 59.1\ g\ NaOH$$

Exercise 5.9.2

What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?

Answer available at the end of this section.

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:

$$2\,AgNO_3\,(aq) + CaCl_2\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)$$

If we wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, we first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution:

$$1.25 \cancel{mol\ AgNO_3} \times \frac{1 \cancel{mol\ CaCl_2}}{2 \cancel{mol\ AgNO_3}} \times \frac{1 L\ solution}{0.555 \cancel{mol\ CaCl_2}} = 1.13\ L\ CaCl_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

Example 5.9.3

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\,Ag\,(s) + Al(NO_3)_3\,(aq) \longrightarrow 3\,AgNO_3 + Al\,(s)$$

Solution

Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66 \cancel{g\ Ag} \times \frac{1 \cancel{mol\ Ag}}{107.97 \cancel{g\ Ag}} \times \frac{1 \cancel{mol\ Al(NO_3)_3}}{3 \cancel{mol\ Ag}} \times \frac{1 L\ solution}{0.0995 \cancel{mol\ Al(NO_3)_3}} = 0.114\ L$$

The strikeouts show how the units cancel.

Exercise 5.9.3

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$H_2C_2O_4\,(s) + 2\,NaOH\,(aq) \longrightarrow Na_2C_2O_4\,(aq) + 2\,H_2O(\ell)$$

Answer: available at the end of this section.

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

Example 5.9.4

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of FeCl3. The student carefully adds 0.1074 M Na2C2O4 until all the Fe^{3+}(aq) has precipitated as Fe2(C2O4)3(s). Using a precisely measured tube called a burette, the student finds that 9.04 mL of the Na2C2O4 solution was added to completely precipitate the Fe^{3+}(aq). What was the concentration of the FeCl3 in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:

$$2\,\mathrm{FeCl_3\,(aq)} + 3\,\mathrm{Na_2C_2O_4\,(aq)} \longrightarrow \mathrm{Fe_2(C_2O_4)_3\,(s)} + 6\,\mathrm{NaCl\,(aq)}$$

Solution

First we need to determine the number of moles of Na2C2O4 that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04\,\cancel{mL} \times \frac{1\,\cancel{L}}{1000\,\cancel{mL}} \times \frac{0.1074 mol\,Na_2C_2O_4}{\cancel{L}} = 0.000971\,mol\,Na_2C_2O_4$$

Now we will use the balanced chemical equation to determine the number of moles of Fe^{3+}(aq) that were present in the initial aliquot:

$$0.000971\,\cancel{mol\,Na_2C_2O_4} \times \frac{2 mol\,FeCl_3}{3\,\cancel{mol\,Na_2C_2O_4}} = 0.000647 mol\,FeCl_3$$

Then we determine the concentration of FeCl3 in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{mol}{L} = \frac{0.000647 mol\,FeCl_3}{0.01000 L} = 0.0647 M\,FeCl_3$$

Exercise 5.9.4

A student titrates 25.00 mL of H3PO4 with 0.0987 M KOH. She uses 54.06 mL to complete the chemical reaction. What is the concentration of H3PO4?

$$\mathrm{H_3PO_4\,(aq)} + 3\,\mathrm{KOH\,(aq)} \longrightarrow \mathrm{K_3PO_4\,(aq)} + 3\,\mathrm{H_2O}$$

Answer: available at the end of this section.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

Example 5.9.5

H2O2 is used to determine the amount of Mn according to this balanced chemical equation:

$$2\,\mathrm{MnO_4^-\,(aq)} + 5\,\mathrm{H_2O_2\,(aq)} + 6\,\mathrm{H^+\,(aq)} \longrightarrow 2\,\mathrm{Mn_2^+\,(aq)} + 5\,\mathrm{O_2\,(g)} + 8\,\mathrm{H_2O(\ell)}$$

What mass of 3.00% m/m H2O2 solution is needed to react with 0.355 mol of MnO_4^-(aq)?

Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of H_2O_2 and then convert to find the mass of H_2O_2. Knowing that the H_2O_2 solution is 3.00% by mass, we can determine the mass of solution needed:

$$0.355 \cancel{mol\ MnO_4^-} \times \frac{5\ \cancel{mol\ H_2O_2}}{2\ \cancel{mol MnO_4^-}} \times \frac{34.02\ \cancel{g\ H_2O_2}}{\cancel{mol\ H_2O_2}} \times \frac{100g\ solution}{3\ \cancel{g\ H_2O_2}} = 1006g\ sol$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of H_2O_2, and the third conversion factor comes from the definition of percentage concentration by mass.

Exercise 5.9.5

Use the balanced chemical reaction for MnO_4^- and H_2O_2 to determine what mass of O_2 is produced if 258 g of 3.00% m/m H_2O_2 is reacted with MnO_4^-.

Answer available at the end of this section.

Summary

Know how to apply concentration units as conversion factors.

Answers

Exercise 5.9.1
1.74 L

Exercise 5.9.2
6.12 g

Exercise 5.9.3
0.777 L

Exercise 5.9.4
0.0711 M



Figure 5.9.1 Titration © Thinkstock. When a student performs a titration, a measured amount of one solution is added to another reactant.

Exercise 5.9.5
7.28 g

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5.10: Density Conversions

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6: CHANGES IN MATTER (UNDER CONSTRUCTION)

6.1: Physical Changes

6.2: Chemical Changes

6.3: Energy Changes

6.4: Chemical Equations

6.5: Types of Chemical Reactions

6.6: Predicting Products of Double Replacement Reactions

6.7: Ionic Equations

6.8: Introduction to Stoichiometry

6.9: Solution Stoichiometry

6.10: Titrations

6.11: Stoichiometry with Reaction Enthalpies

6.12: Theoretical and Percent Yields

6.13: Limiting and Excess Reactants

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6.1: Physical 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. Each water molecule still contains two hydrogen atoms covalently bonded to one oxygen atom. When you have a jar containing a mixture of pennies and nickels, and you sort the mixture so that you have one pile of pennies and another pile of nickels, you have not altered the identity of either the pennies or the nickels—you've merely separated them into two groups. This would be an example of a physical change. Similarly, if you have a piece of paper, you don't change it into something other than a piece of paper by ripping it up. What was paper before you started tearing is still paper when you're done. Again, this is an example of a physical change.



Figure 6.1.3: 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). Dissolution 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.

Summary

- Physical changes are changes that do not alter the identity of a substance.

Contributors and Attributions

-
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Learning Objectives

- Describe the different states of matter.
- Describe the different changes that occur between solid, liquid, and gas.

Matter can exist in one of several different states, including a gas, liquid, or solid state. The amount of energy in molecules of matter determines the **state of matter**.

- A **gas** is a state of matter in which atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide.
- A **liquid** is a state of matter in which atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another.
- A **solid** is a state of matter in which atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another.

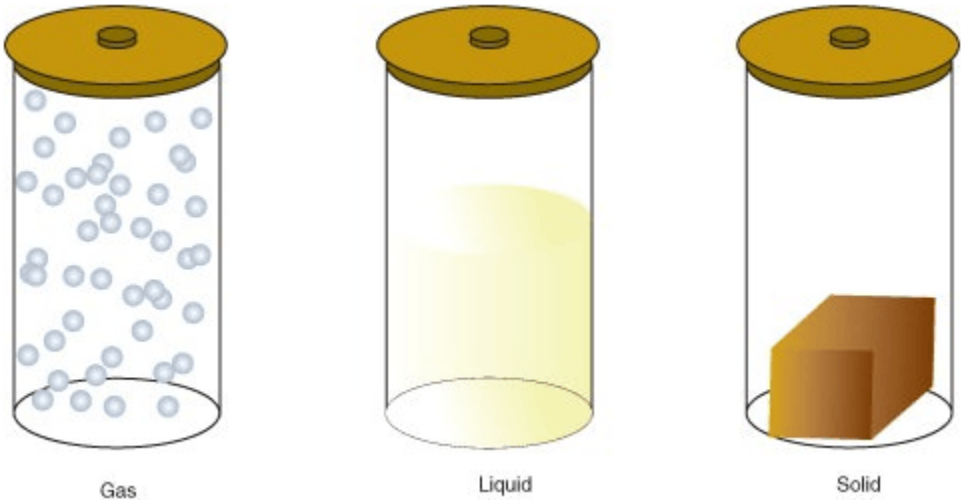


Figure 6.1.1: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

In Figure 6.1.2, all three phases are shown at the submicroscopic level in animations. Notice how the movement and freedom of molecules steadily increases as attractive forces decrease from solid to liquid to gas phase.

Solid matter Liquid matter Gas

SOLID LIQUID GAS

Figure 6.1.2 *Animation of all three phases at the submicroscopic level.*

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. Figures 6.1.3 and 6.1.4 show the differences among solids, liquids, and gases at the molecular level, while Table 6.1.1 lists the different characteristics of these states.

Table 6.1.1 Characteristics of the Three States of Matter

Characteristic	Solid	Liquid	Gas
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 6.1.1

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

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

Solution

a. This statement describes either the liquid state or the solid state.
b. This statement describes either the liquid state or the gas state.
c. This statement describes the liquid state.

Exercise 6.1.1

What state or states of matter does each statement describe?

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

Answer a available at the end of this section.

Phase Changes

A phase change is a physical process in which a substance goes from one phase to another. Usually the change occurs when adding or removing heat at a particular temperature, known as the melting point or the boiling point of the substance. The **melting point** is the temperature at which the substance goes from a solid to a liquid (or from a liquid to a solid). The **boiling point** is the temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid). The nature of the phase change depends on the direction of the heat transfer. Heat going *into* a substance changes it from a solid to a liquid or a liquid to a gas. Removing heat *from* a substance changes a gas to a liquid or a liquid to a solid.

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C. At 0°C, both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O. If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O. A similar process can occur at 100°C: adding heat increases the amount of gaseous H_2O, while removing heat increases the amount of liquid H_2O (Figure 6.1.3).

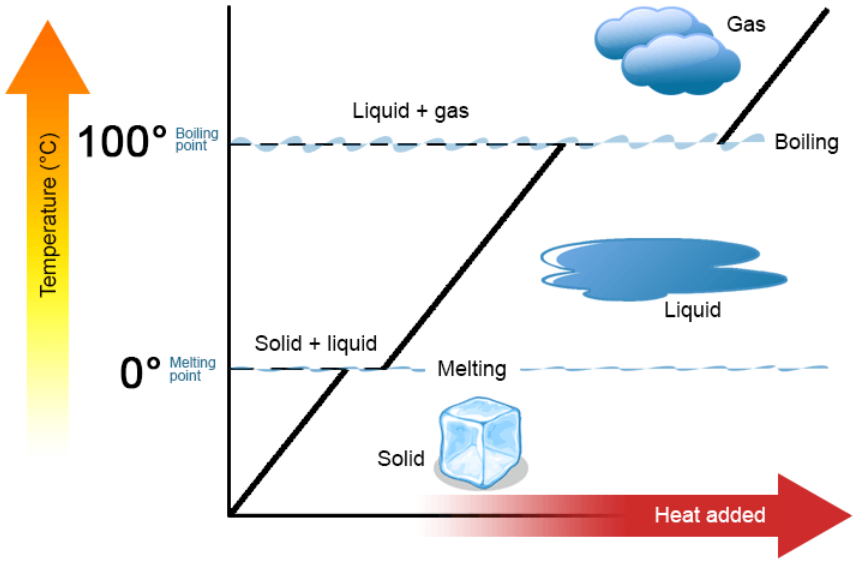


Figure 6.1.3: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0 °C, *the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).*

> Water is a good substance to use as an example because many people are already familiar with it. Other substances have melting points and boiling points as well.

Second, as shown in Figure 6.1.3, **the temperature of a substance does not change** as the substance goes **from one phase to another**. In other words, phase changes are isothermal (isothermal means "constant temperature"). Again, consider H_2O as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of H_2O can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

Summary of Phase Changes

All of the changes of state that occur between solid, liquid, and gas are summarized in the diagram in the figure below. **Freezing** is the opposite of melting, and both represent the equilibrium between the solid and liquid states. **Evaporation** occurs when a liquid turns to a gas. **Condensation** is the opposite of vaporization, and both represent the equilibrium between the liquid and gas states. **Deposition** is the opposite of sublimation, and both represent the equilibrium between the solid and gas states.

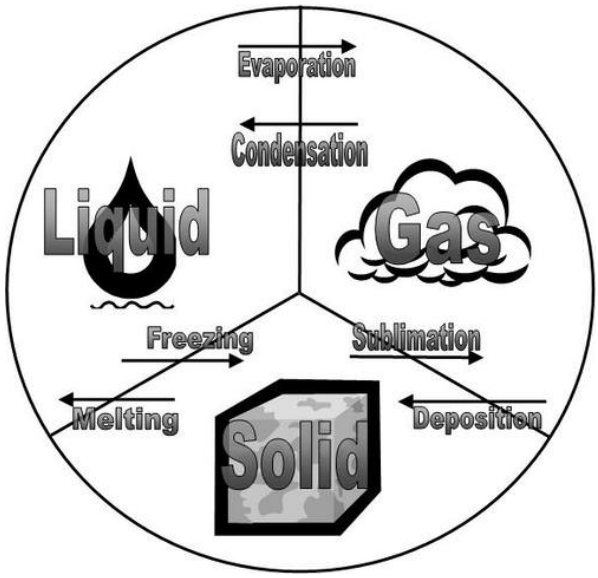


Figure 6.1.4: Solid, liquid, and gas states with the terms for each change of state that occurs between them.
Liquid evaporates into gas in freezes into a solid; gas condenses into a liquid depositions into a solid; solids melt into a liquid and sublimation into a gas

Summary

- Solids, liquids, and gases are different phases that have their own unique properties.
- A change of state (or phase chnage) can be brought about by putting heat into a system or removing it from the system.

Contributors and Attributions

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- Libretexts: The Basics of GOB (Ball et.al.)

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In many areas, winter ice on the streets and sidewalks represents a serious walking and driving hazard. One common approach to melting the ice is to put some form of deicing salt on the surface. Materials such as sodium chloride or calcium chloride are frequently employed for this purpose. In order to be effective, the solid material must first dissolve and break up into the ions that make up the compound.

Dissociation

An ionic crystal lattice breaks apart when it is dissolved in water. **Dissociation** is the separation of ions that occurs when a solid ionic compound dissolves. It is important to be able to write dissociation equations. Simply undo the crisscross method that you learned when writing chemical formulas of ionic compounds. The subscripts for the ions in the chemical formulas become the coefficients of the respective ions on the product side of the equation. Shown below are dissociation equations for NaCl, $\mathrm{Ca(NO_3)_2}$, and $\mathrm{(NH_4)_3PO_4}$.

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

$$\mathrm{Ca(NO_3)_2}\,(s) \rightarrow \mathrm{Ca^{2+}}\,(aq) + \mathrm{2NO_3{}^-}\,(aq)$$

$$\mathrm{(NH_4)_3PO_4}\,(s) \rightarrow \mathrm{3NH_4}\,(aq) + \mathrm{PO_4}\,(aq)$$

The formula unit of sodium chloride dissociates into one sodium ion and one chloride ion. The calcium nitrate formula unit

dissociates into one calcium ion and two nitrate ions. This is because of the $2+$ charge of the calcium ion. Two nitrate ions, each with a $1-$ charge are required to make the equation balance electrically. The ammonium phosphate formula unit dissociates into three ammonium ions and one phosphate ion. Note that the polyatomic ions themselves do not dissociate further, but remain intact.

Do not confuse the subscripts of the atoms within the polyatomic ion for the subscripts that result from the crisscrossing of the charges that make up the original compound neutral. The 3 subscript of the nitrate ion and the 4 subscript of the ammonium ion are part of the polyatomic ion and simply remain as part of its formula after the compound dissociates. Notice that the compounds are solids (s) which then become ions in aqueous solution (aq).



Figure $6.1.1$: Calcium nitrate is a typical ionic compound. In an aqueous solution, it dissociates into calcium ions and nitrate ions.

Nonelectrolytes do not dissociate when forming an aqueous solution. An equation can still be written that simply shows the solid going into solution. For the dissolution of sucrose:

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

Summary

- Dissociation is the separation of ions that occurs when a solid ionic compound dissolves.
- Nonionic compounds do not dissociate in water.

Answers

Exercise $6.1.1$

solid
gas
liquid or gas

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6.2: Chemical Changes

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, to be exact) from the burning of the wax. Another example of a chemical change is what occurs when natural gas is burned in a furnace. This time, on the left there is a molecule of methane, $\mathrm{CH_4}$, and two molecules of oxygen, $\mathrm{O_2}$; on the right are two molecules of water, $\mathrm{H_2O}$, and one molecule of carbon dioxide, $\mathrm{CO_2}$. 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 6.2.4: The 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 (temperature increase or decrease).
- 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 6.2.2

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

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

Solution

1. Physical: boiling and melting are physical changes. When water boils, no bonds are broken or formed. The change could be written as: $\mathrm{H_2O}\,(l) \rightarrow \mathrm{H_2O}\,(g)$
2. 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: $\mathrm{Fe} + \mathrm{O_2} \rightarrow \mathrm{Fe_2O_3}$
3. Physical: because none of the properties changed, this is a physical change. The green mixture is still green and the colorless solution is still colorless. They have just been combined. No color *change*, or other evidence of chemical change, occurred.
4. Chemical: the formation of a precipitate and the color change from colorless to yellow indicates a chemical change.

Exercise 6.2.2

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

1. A mirror is broken.
2. An iron nail corroded in moist air.
3. Copper metal is melted.
4. A catalytic converter changes nitrogen dioxide to nitrogen gas and oxygen gas.

Answer 1: available at the end of this section.

Answer 2: available at the end of this section.

Answer 3: available at the end of this section.

Answer 4: available at the end of this section.

Summary

- Chemical changes are changes that occur when one substance is turned into another substance.
- Chemical changes are frequently harder to reverse than physical changes. Observations that indicate a chemical change occurred include color change, temperature change, light given off, formation of bubbles, and formation of a precipitate.

Contributors and Attributions

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- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)



Figure 6.2.1 (Public Domain; Jon Sullivan via Wikipedia)

Do you like to cook?

Cooking is a valuable skill that can benefit everyone. Whether it is fixing a simple grilled cheese sandwich or preparing an elaborate meal, cooking demonstrates some basic concepts of chemistry. When you bake bread, you mix flour, sugar, yeast, and water together. After baking, this mixture has been changed to form bread, another substance that has different characteristics and qualities from the original materials. The process of baking has produced chemical changes in the ingredients that result in bread being made.

Chemical Change

Most of the elements we know about do not exist freely in nature. Sodium cannot be found by itself (unless we prepare it in the laboratory) because it interacts too easily with other materials. On the other hand, the element helium does not interact with other elements to any extent. We can isolate helium from natural gas during the process of drilling for oil.

A chemical change produces different materials than the ones we started with. One aspect of the science of chemistry is the study of the changes that matter undergoes. If chemistry was simply a study of elements that did nothing, life would be very boring (in fact, life would not exist since the elements are what make up our bodies and sustain us). But the processes of change that take place when different chemicals are combined produce all the materials that we use daily.

One type of chemical change (already mentioned) is when two elements combine to form a compound. Another type involves the breakdown of a compound to produce the elements that make it up. If we pass an electric current through bauxite (aluminum oxide, the raw material for aluminum metal), we get metallic aluminum as a product.



Figure 6.2.1: Electrolytic production of aluminum. (Public Domain; Ronny Jacques via Wikipedia)

However, the vast majority of chemical changes involve one compound being transformed into another compound. There are literally millions of possibilities when we take this approach to chemical change. New compounds can be made to produce better fabrics that are easier to clean and maintain; they can help preserve food so it doesn't spoil as quickly; we can make new medicines to treat diseases; even products, such as airbags (as seen in the simulation below), use a chemical change to improve the world around us.

Simulation

How does a chemical reaction inflate an airbag?

Summary

A chemical change produces different materials than the ones we started with.

Review

1. What is a chemical change?
2. List three types of chemical changes.

Answers

Exercise 6.2.2
physical change

Exercise 6.2.2
chemical change

Exercise 6.2.2
physical change

Exercise 6.2.2
chemical change

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6.3: Energy Changes

Learning Objectives

- Define a type of work in terms of pressure and volume.
- Define *heat*.
- Relate the amount of heat to a temperature change.

We have already defined work as a force acting through a distance. There are other equivalent definitions of work that are also important in chemistry.

When a certain volume of a gas expands, it works against an external pressure to expand (Figure 6.3.1 - Volume versus Pressure). That is, the gas must perform work. Assuming that the external pressure P_{ext} is constant, the amount of work done by the gas is given by the equation

$$w = -P_{ext} \times \Delta V$$

where ΔV is the change in volume of the gas. This term is always the final volume minus the initial volume,

$$\Delta V = V_{final} - V_{initial}$$

and can be positive or negative, depending on whether V_{final} is larger (is expanding) or smaller (is contracting) than $V_{initial}$. The negative sign in the equation for work is important and implies that as volume expands (ΔV is positive), the gas in the system is *losing* energy as work. On the other hand, if the gas is contracting, ΔV is negative, and the two negative signs make the work positive, so energy is being added to the system.

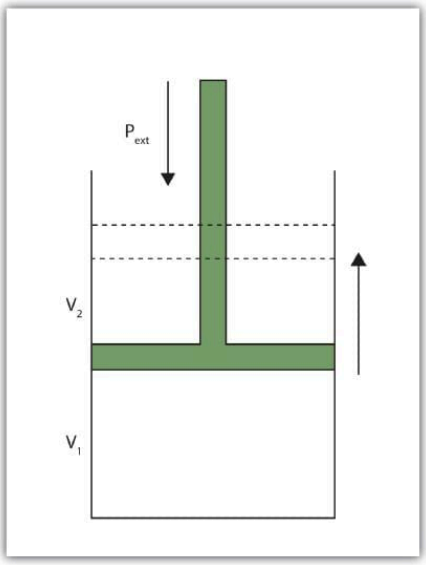


Figure 6.3.1 Volume versus Pressure. *When a gas expands against an external pressure, the gas does work.*

Finally, let us consider units. Volume changes are usually expressed in units like liters, while pressures are usually expressed in atmospheres. When we use the equation to determine work, the unit for work comes out as liter·atmospheres, or L·atm. This is not a very common unit for work. However, there is a conversion factor between L·atm and the common unit of work, joules:

$$1 \text{ L·atm} = 101.32 \text{ J}$$

Using this conversion factor and the previous equation for work, we can calculate the work performed when a gas expands or contracts.

Example 6.3.1

What is the work performed by a gas if it expands from 3.44 L to 6.19 L against a constant external pressure of 1.26 atm? Express the final answer in joules.

Solution

First we need to determine the change in volume, ΔV. A change is always the final value minus the initial value:

$$\Delta V = V_{\text{final}} - V_{\text{initial}} = 6.19\ \text{L} - 3.44\ \text{L} = 2.75\ \text{L}$$

Now we can use the definition of work to determine the work done:

$$w = -P_{\text{ext}} \cdot \Delta V = -(1.26\ \text{atm})(2.75\ \text{L}) = -3.47\ \text{L·atm}$$

Now we construct a conversion factor from the relationship between liter·atmospheres and joules:

$$-3.47\ \cancel{L.atm} \times \times \frac{101.32\ J}{1\ \cancel{L.atm}} = -351\ J$$

We limit the final answer to three significant figures, as appropriate.

Exercise 6.3.1

What is the work performed when a gas expands from 0.66 L to 1.33 L against an external pressure of 0.775 atm?

Answer available at the end of this section.

Heat is another aspect of energy. **Heat** is the transfer of energy from one body to another due to a difference in temperature. For example, when we touch something with our hands, we interpret that object as either hot or cold depending on how energy is transferred: If energy is transferred into your hands, the object feels hot. If energy is transferred from your hands to the object, your hands feel cold. Because heat is a measure of energy transfer, heat is also measured in joules.

For a given object, the amount of heat (q) involved is proportional to two things: the mass of the object (m) and the temperature change (ΔT) evoked by the energy transfer. We can write this mathematically as

$$q \alpha\, m \times \Delta T$$

where $\propto$ means "is proportional to." To make a proportionality an equality, we include a proportionality constant. In this case, the proportionality constant is labeled c and is called the **specific heat capacity**, or, more succinctly, **specific heat**:

\[q = mc\Delta T \nonumber \]

where the mass, specific heat, and change in temperature are multiplied together. Specific heat is a measure of how much energy is needed to change the temperature of a substance; the larger the specific heat, the more energy is needed to change the temperature. The units for specific heat are

$$\frac{J}{g.C} \text{ or } \frac{J}{g.K}$$

depending on what the unit of ΔT is. You may note a departure from the insistence that temperature be expressed in Kelvin. That is because a *change* in temperature has the same value whether the temperatures are expressed in degrees Celsius or kelvins.

Example 6.3.2

Calculate the heat involved when 25.0 g of Fe increase temperature from 22°C to 76°C. The specific heat of Fe is 0.449 J/g·°C.

Solution

First we need to determine ΔT. A change is always the final value minus the initial value:

$$\Delta T = 76°C - 22°C = 54°C$$

Now we can use the expression for q, substitute for all variables, and solve for heat:

$$q = (25.0\,g)(0.449\frac{J}{g.°C})(54°C) = 610\,J$$

Note how the g and °C units cancel, leaving J, a unit of heat. Also note that this value of q is inherently positive, meaning that energy is going into the system.

Exercise 6.3.2

Calculate the heat involved when 76.5 g of Ag increase temperature from 17.8°C to 144.5°C. The specific heat of Ag is 0.233 J/g·°C.

Answer available at the end of this section.

As with any equation, when you know all but one variable in the expression for q, you can determine the remaining variable by using algebra.

Example 6.3.3

It takes 5,408 J of heat to raise the temperature of 373 g of Hg by 104°C. What is the specific heat of Hg?

Solution

We can start with the equation for q, but now different values are given, and we need to solve for specific heat. Note that ΔT is given directly as 104°C. Substituting,

$$5{,}408\text{ J} = (373\text{ g})c(104°C)$$

We divide both sides of the equation by 373 g and 104°C:

$$c = \frac{5408\,J}{(373g)(104°C)}$$

Combining the numbers and bringing together all the units, we get

$$c = 0.139\frac{J}{g.°C}$$

Exercise 6.3.3

Gold has a specific heat of 0.129 J/g·°C. If 1,377 J are needed to increase the temperature of a sample of gold by 99.9°C, what is the mass of the gold?

Answer available at the end of this section.

Table 6.3.1: Specific Heats of Various Substances, lists the specific heats of some substances. Specific heat is a physical property of substances, so it is a characteristic of the substance. The general idea is that the lower the specific heat, the less energy is required to change the temperature of the substance by a certain amount.

Table 6.3.1: Specific Heats of Various Substances.

Substance	Specific Heat (J/g·°C)
water	4.184
iron	0.449
gold	0.129
mercury	0.139
aluminum	0.900
ethyl alcohol	2.419
magnesium	1.03
helium	5.171
oxygen	0.918

Key Takeaways

- Work can be defined as a gas changing volume against a constant external pressure.
- Heat is the transfer of energy due to temperature differences.
- Heat can be calculated in terms of mass, temperature change, and specific heat.

Learning Objectives

- Describe the different states of matter.
- Describe the different changes that occur between solid, liquid, and gas.

Matter can exist in one of several different states, including a gas, liquid, or solid state. The amount of energy in molecules of matter determines the **state of matter**.

- A **gas** is a state of matter in which atoms or molecules have enough energy to move freely. The molecules come into contact with one another only when they randomly collide.
- A **liquid** is a state of matter in which atoms or molecules are constantly in contact but have enough energy to keep changing positions relative to one another.
- A **solid** is a state of matter in which atoms or molecules do not have enough energy to move. They are constantly in contact and in fixed positions relative to one another.

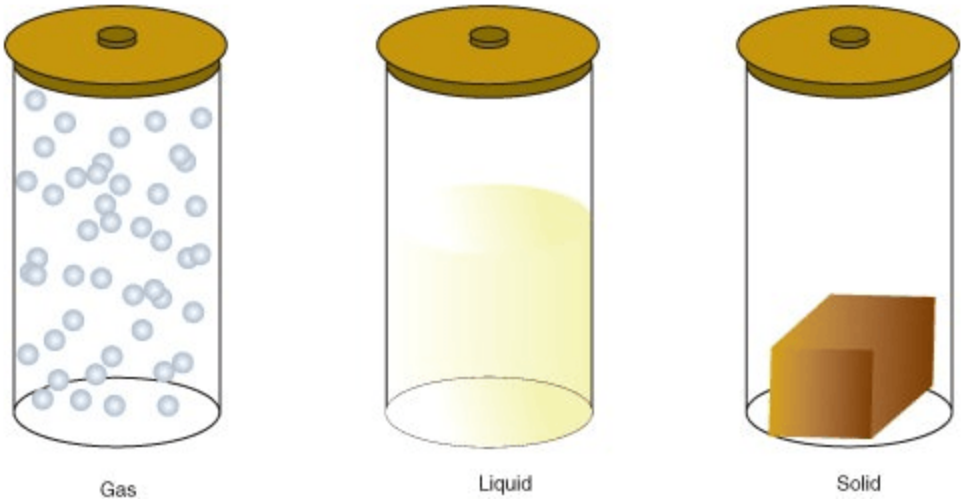


Figure 6.3.1: States of Matter. All three containers contain a substance with the same mass, but the substances are in different states. In the left-hand container, the substance is a gas, which has spread to fill its container. It takes both the shape and volume of the container. In the middle container, the substance is a liquid, which has spread to take the shape of its container but not the volume. In the right-hand container, the substance is a solid, which takes neither the shape nor the volume of its container.

In Figure 6.3.2, all three phases are shown at the submicroscopic level in animations. Notice how the movement and freedom of molecules steadily increases as attractive forces decrease from solid to liquid to gas phase.

Solid matter Liquid matter Gas

SOLID LIQUID GAS

Figure 6.3.2 *Animation of all three phases at the submicroscopic level.*

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. Figures 6.3.3 and 6.3.4 show the differences among solids, liquids, and gases at the molecular level, while Table 6.3.1 lists the different characteristics of these states.

Table 6.3.1 *Characteristics of the Three States of Matter*

Characteristic	Solid	Liquid	Gas
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 6.3.1

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

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

Solution

a. This statement describes either the liquid state or the solid state.
b. This statement describes either the liquid state or the gas state.
c. This statement describes the liquid state.

Exercise 6.3.1

What state or states of matter does each statement describe?

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

Answer a available at the end of this section.

Phase Changes

A phase change is a physical process in which a substance goes from one phase to another. Usually the change occurs when adding or removing heat at a particular temperature, known as the melting point or the boiling point of the substance. The **melting point** is the temperature at which the substance goes from a solid to a liquid (or from a liquid to a solid). The **boiling point** is the temperature at which a substance goes from a liquid to a gas (or from a gas to a liquid). The nature of the phase change depends on the direction of the heat transfer. Heat going *into* a substance changes it from a solid to a liquid or a liquid to a gas. Removing heat *from* a substance changes a gas to a liquid or a liquid to a solid.

Two key points are worth emphasizing. First, at a substance's melting point or boiling point, two phases can exist simultaneously. Take water (H_2O) as an example. On the Celsius scale, H_2O has a melting point of 0°C and a boiling point of 100°C. At 0°C, both the solid and liquid phases of H_2O can coexist. However, if heat is added, some of the solid H_2O will melt and turn into liquid H_2O. If heat is removed, the opposite happens: some of the liquid H_2O turns into solid H_2O. A similar process can occur at 100°C: adding heat increases the amount of gaseous H_2O, while removing heat increases the amount of liquid H_2O (Figure 6.3.3).

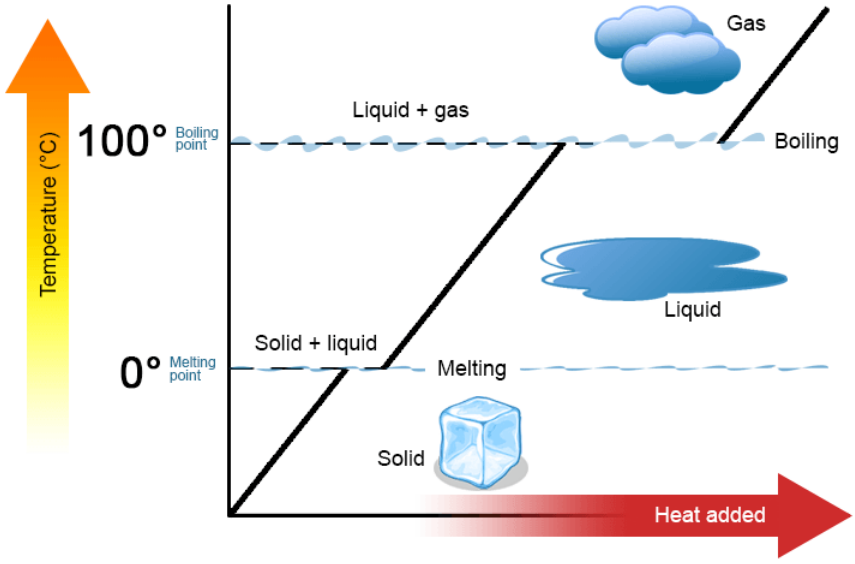


Figure 6.3.3: Heating curve for water. As heat is added to solid water, the temperature increases until it reaches 0 °C, *the melting point. At this point, the phase change, added heat goes into changing the state from a solid to liquid. Only when this phase change is complete, the temperature can increase. (CC BY 3.0 Unported; Community College Consortium for Bioscience Credentials).*

Water is a good substance to use as an example because many people are already familiar with it. Other substances have melting points and boiling points as well.

Second, as shown in Figure 6.3.3, **the temperature of a substance does not change** as the substance goes **from one phase to another**. In other words, phase changes are isothermal (isothermal means “constant temperature”). Again, consider H_2O as an example. Solid water (ice) can exist at 0°C. If heat is added to ice at 0°C, some of the solid changes phase to make liquid, which is also at 0°C. Remember, the solid and liquid phases of H_2O can coexist at 0°C. Only after all of the solid has melted into liquid does the addition of heat change the temperature of the substance.

Summary of Phase Changes

All of the changes of state that occur between solid, liquid, and gas are summarized in the diagram in the figure below. **Freezing** is the opposite of melting, and both represent the equilibrium between the solid and liquid states. **Evaporation** occurs when a liquid turns to a gas. **Condensation** is the opposite of vaporization, and both represent the equilibrium between the liquid and gas states. **Deposition** is the opposite of sublimation, and both represent the equilibrium between the solid and gas states.

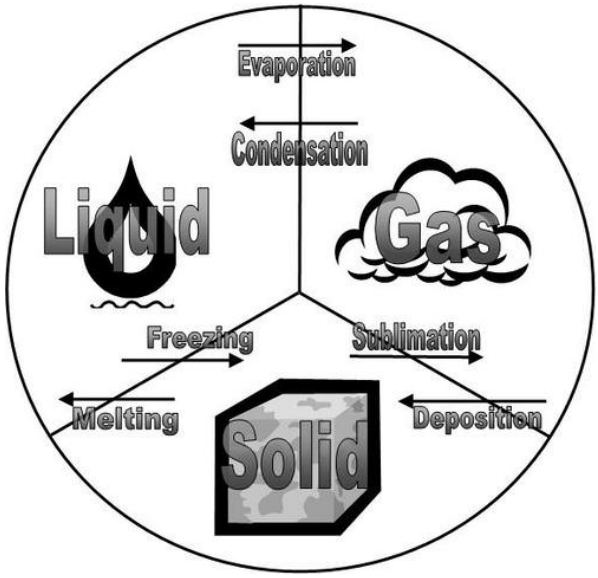


Figure 6.3.4: Solid, liquid, and gas states with the terms for each change of state that occurs between them.
Liquid evaporates into gas in freezes into a solid; gas condenses into a liquid depositions into a solid; solids melt into a liquid and sublimation into a gas

Summary

- Solids, liquids, and gases are different phases that have their own unique properties.
- A change of state (or phase chnage) can be brought about by putting heat into a system or removing it from the system.

Contributors and Attributions

-
- Libretexts: The Basics of GOB (Ball et.al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Learning Objective

- Describe what happens during a phase change.
- Calculate the energy change needed for a phase change.

Substances can change phase—often because of a temperature change. At low temperatures, most substances are solid; as the temperature increases, they become liquid; at higher temperatures still, they become gaseous.

The process of a solid becoming a liquid is called **melting** (an older term that you may see sometimes is *fusion*). The opposite process, a liquid becoming a solid, is called **solidification**. For any pure substance, the temperature at which melting occurs—known as the **melting point**—is a characteristic of that substance. It requires energy for a solid to melt into a liquid. Every pure substance has a certain amount of energy it needs to change from a solid to a liquid. This amount is called the **enthalpy of fusion (or heat of fusion)** of the substance, represented as ΔH_{fus}. Some ΔH_{fus} values are listed in Table 6.3.1; it is assumed that these values are for the melting point of the substance. Note that the unit of ΔH_{fus} is kilojoules per mole, so we need to know the quantity of material to know how much energy is involved. The ΔH_{fus} is always tabulated as a positive number. However, it can be used for both the melting and the solidification processes, minding that melting is always endothermic (so ΔH will be positive), while solidification is always exothermic (so ΔH will be negative).

Table 6.3.1: Enthalpies of Fusion for Various Substances

Substance (Melting Point)	ΔH_{fus} (kJ/mol)
Water (0°C)	6.01
Aluminum (660°C)	10.7
Benzene (5.5°C)	9.95
Ethanol (–114.3°C)	5.02
Mercury (–38.8°C)	2.29

Example 6.3.1

What is the energy change when 45.7 g of $\mathrm{H_2O}$ melt at 0°C?

Solution

The ΔH_{fus} of $\mathrm{H_2O}$ is 6.01 kJ/mol. However, our quantity is given in units of grams, not moles, so the first step is to convert grams to moles using the molar mass of $\mathrm{H_2O}$, which is 18.0 g/mol. Then we can use ΔH_{fus} as a conversion factor. Because the substance is melting, the process is endothermic, so the energy change will have a positive sign.

$$45.7\ \cancel{g\ H_2O} \times \frac{1\ \cancel{mol\ H_2O}}{18.0\ \cancel{g}} \times \frac{6.01 kJ}{\cancel{mol}} = 15.3\ kJ$$

Without a sign, the number is assumed to be positive.

Exercise 6.3.1

What is the energy change when 108 g of $\mathrm{C_6H_6}$ freeze at 5.5°C?

Answer available at the end of this section.

During melting, energy goes exclusively to changing the phase of a substance; it does not go into changing the temperature of a substance. Hence melting is an **isothermal** process because a substance stays at the same temperature. Only when all of a substance is melted does additional energy go to changing its temperature.

What happens when a solid becomes a liquid? In a solid, individual particles are stuck in place because the intermolecular forces cannot be overcome by the energy of the particles. When more energy is supplied (e.g., by raising the temperature), there comes a point at which the particles have enough energy to move around, but not enough energy to separate. This is the liquid phase: particles are still in contact, but are able to move around each other. This explains why liquids can assume the shape of their containers: the particles move around and, under the influence of gravity, fill the lowest volume possible (unless the liquid is in a zero-gravity environment—see Figure 6.3.1.



Figure 6.3.1: Liquids and Gravity. (a) A liquid fills the bottom of its container as it is drawn downward by gravity and the particles slide over each other. (b) A liquid floats in a zero-gravity environment. The particles still slide over each other because they are in the liquid phase, but now there is no gravity to pull them down. Source: Photo on the left © Thinkstock. Photo on the right courtesy of NASA, http://www.nasa.gov/mission_pages/st...image_009.html.

The phase change between a liquid and a gas has some similarities to the phase change between a solid and a liquid. At a certain temperature, the particles in a liquid have enough energy to become a gas. The process of a liquid becoming a gas is called **boiling (or vaporization)**, while the process of a gas becoming a liquid is called **condensation**. However, unlike the solid/liquid conversion process, the liquid/gas conversion process is noticeably affected by the surrounding pressure on the liquid because gases are strongly affected by pressure. This means that the temperature at which a liquid becomes a gas, the **boiling point**, can change with surrounding pressure. Therefore, we define the **normal boiling point** as the temperature at which a liquid changes to a gas when the surrounding pressure is exactly 1 atm, or 760 torr. Unless otherwise specified, it is assumed that a boiling point is for 1 atm of pressure.

Like the solid/liquid phase change, the liquid/gas phase change involves energy. The amount of energy required to convert a liquid to a gas is called the **enthalpy of vaporization** (or heat of vaporization), represented as ΔH_{vap}. Some ΔH_{vap} values are listed in Table 6.3.2; it is assumed that these values are for the normal boiling point temperature of the substance, which is also given in the table. The unit for ΔH_{vap} is also kilojoules per mole, so we need to know the quantity of material to know how much energy is involved. The ΔH_{vap} is also always tabulated as a positive number. It can be used for both the boiling and the condensation processes as long as you keep in mind that boiling is always endothermic (so ΔH will be positive), while condensation is always exothermic (so ΔH will be negative).

Table 6.3.1: Enthalpies of Vaporization for Various Substances

Substance (Normal Boiling Point)	ΔH_{vap} (kJ/mol)
Water (100°C)	40.68
Bromine (59.5°C)	15.4
Benzene (80.1°C)	30.8
Ethanol (78.3°C)	38.6
Mercury (357°C)	59.23

Example 6.3.2

What is the energy change when 66.7 g of $Br_2(g)$ condense to a liquid at 59.5°C?

Solution

The ΔH_{vap} of Br_2 is 15.4 kJ/mol. Even though this is a condensation process, we can still use the numerical value of ΔH_{vap} as long as we realize that we must take energy out, so the ΔH value will be negative. To determine the magnitude of the energy change, we must first convert the amount of Br_2 to moles. Then we can use ΔH_{vap} as a conversion factor.

$$66.7\ \cancel{g\ Br_2} \times \frac{1\ \cancel{mol\ Br_2}}{159.8\ \cancel{g}} \times \frac{15.4 kJ}{\cancel{mol}} = 6.43\ kJ$$

Because the process is exothermic, the actual value will be negative: ΔH = −6.43 kJ.

Exercise 6.3.2

What is the energy change when 822 g of $C_2H_5OH(\ell)$ boil at its normal boiling point of 78.3°C?

Answer available at the end of this section.

As with melting, the energy in boiling goes exclusively to changing the phase of a substance; it does not go into changing the temperature of a substance. So boiling is also an isothermal process. Only when all of a substance has boiled does any additional energy go to changing its temperature.

What happens when a liquid becomes a gas? We have already established that a liquid is composed of particles in contact with each other. When a liquid becomes a gas, the particles separate from each other, with each particle going its own way in space. This is how gases tend to fill their containers. Indeed, in the gas phase most of the volume is empty space; only about 1/1,000th of the volume is actually taken up by matter (Figure 6.3.1). It is this property of gases that explains why they can be compressed, a fact that is considered in Chapter 6.

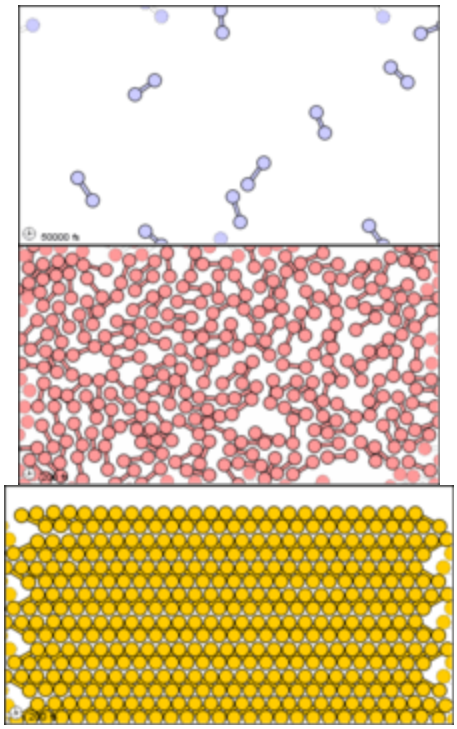


Figure 6.3.2: Sub-microscopic view of the diatomic molecules of the element bromine (a) in the gaseous state (above 58°C); (b) in liquid form (between -7.2 and 58.8°C); and (c) in solid form (below -7.2°C). As a solid, the molecules are fixed, but fluctuate. As a liquid, the molecules are in contact but are also able to move around each other. As a gas, most of the volume is actually empty space. The particles are not to scale; in reality, the dots representing the particles would be about 1/100th of the size depicted.

Under some circumstances, the solid phase can transition directly to the gas phase without going through a liquid phase, and a gas can directly become a solid. The solid-to-gas change is called **sublimation**, while the reverse process is called **deposition.** Sublimation is isothermal, like the other phase changes. There is a measurable energy change during sublimation—this energy change is called the **enthalpy of sublimation**, represented as ΔH_{sub}. The relationship between the ΔH_{sub} and the other enthalpy changes is as follows:

$$\Delta H_{sub} = \Delta H_{fus} + \Delta H_{vap}$$

As such, ΔH_{sub} is not always tabulated because it can be simply calculated from ΔH_{fus} and ΔH_{vap}.

There are several common examples of sublimation. A well-known product, dry ice, is actually solid CO_2. Dry ice is dry because it sublimes, with the solid bypassing the liquid phase and going straight to the gas phase. The sublimation occurs at temperature of –77°C, so it must be handled with caution. If you have ever noticed that ice cubes in a freezer tend to get smaller over time, it is because the solid water is very slowly subliming. "Freezer burn" isn't actually a burn; it occurs when certain foods, such as meats, slowly lose solid water content because of sublimation. The food is still good, but looks unappetizing. Reducing the temperature of a freezer will slow the sublimation of solid water.



Figure 6.3.3: Freezer burn on a piece of beef. (Public Domain; RolloM.)

Chemical equations can be used to represent a phase change. In such cases, it is crucial to use phase labels on the substances. For example, the chemical equation for the melting of ice to make liquid water is as follows:

$$H_2O(s) \rightarrow H_2O(\ell)$$

No chemical change is taking place; however, a physical change is taking place.

Summary

- Phase changes can occur between any two phases of matter.
- All phase changes occur with a simultaneous change in energy.
- All phase changes are isothermal.

Answers

Exercise 6.3.1
–53 J

Exercise 6.3.2
2,260 J

Exercise 6.3.3
107 g

Exercise 6.3.1
solid
gas
liquid or gas

Exercise 6.3.1
–13.8 kJ

Exercise 6.3.2
689 kJ

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6.4: Chemical Equations

Learning Objectives

- Define *chemical equation*.
- Identify the parts of a chemical equation.

A chemical reaction expresses a chemical change. For example, one chemical property of hydrogen is that it will react with oxygen to make water. We can write that as follows:

$$\text{hydrogen reacts with oxygen to make water}$$

We can represent this chemical change more succinctly as:

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water}$$

where the + sign means that the two substances interact chemically with each other and the → symbol implies that a chemical reaction takes place. But substances can also be represented by chemical formulas. Remembering that hydrogen and oxygen both exist as diatomic molecules, we can rewrite our chemical change as:

$$\mathrm{H_2} + \mathrm{O_2} \longrightarrow \mathrm{H_2O}$$

This is an example of a chemical equation, which is a concise way of representing a **chemical reaction**. The initial substances are called **reactants** and the final substances are called **products**.

Unfortunately, it is also an *incomplete* chemical equation. The law of conservation of matter says that matter cannot be created or destroyed. In chemical equations, the number of atoms of each element in the reactants must be the same as the number of atoms of each element in the products. If we count the number of hydrogen atoms in the reactants and products, we find two hydrogen atoms. But if we count the number of oxygen atoms in the reactants and products, we find that there are two oxygen atoms in the reactants, but only one oxygen atom in the products.

What can we do? Can we change the subscripts in the formula for water so that it has two oxygen atoms in it? No—you *cannot* change the formulas of individual substances, because the chemical formula for a given substance is characteristic of that substance. What you *can* do, however, is to change the number of molecules that react or are produced. We do this one element at a time, going from one side of the reaction to the other, changing the number of molecules of a substance until all elements have the same number of atoms on each side.

To accommodate the two oxygen atoms as reactants, let us assume that we have two water molecules as products:

$$\mathrm{H_2} + \mathrm{O_2} \longrightarrow 2\,\mathrm{H_2O}$$

The 2 in front of the formula for water is called a **coefficient**. Now there are the same number of oxygen atoms in the reactants as there are in the product. But in satisfying the need for the same number of oxygen atoms on both sides of the reaction, we have also changed the number of hydrogen atoms on the product side, so the number of hydrogen atoms is no longer equal. No problem, simply go back to the reactant side of the equation, and add a coefficient in front of the H2. The coefficient that works is 2:

$$2\,\mathrm{H_2} + \mathrm{O_2} \longrightarrow 2\,\mathrm{H_2O}$$

There are now four hydrogen atoms in the reactants and also four atoms of hydrogen in the product. There are two oxygen atoms in the reactants and two atoms of oxygen in the product. The law of conservation of matter has been satisfied. When the reactants and products of a chemical equation have the same number of atoms of all elements present, we say that an equation is **balanced**. All proper chemical equations are balanced. If a substance does not have a coefficient written in front of it, it is assumed to be 1. Also, the convention is to use all whole numbers when balancing chemical equations. This sometimes makes us do a bit more "back and forth" work when balancing a chemical equation.

Example 6.4.1

Write and balance the chemical equation for each given chemical reaction.

1. Hydrogen and chlorine react to make HCl
2. Ethane, $\mathrm{C_2H_6}$, reacts with oxygen to make carbon dioxide and water.

Solution

1. Let us start by simply writing a chemical equation in terms of the formulas of the substances, remembering that both elemental hydrogen and chlorine are diatomic:

$$\mathrm{H_2 + Cl_2 \longrightarrow HCl}$$

There are two hydrogen atoms and two chlorine atoms in the reactants and one of each atom in the product. We can fix this by including the coefficient 2 on the product side:

$$\mathrm{H_2 + Cl_2 \longrightarrow 2\,HCl}$$

Now there are two hydrogen atoms and two chlorine atoms on both sides of the chemical equation, so it is balanced.

2. Start by writing the chemical equation in terms of the substances involved:

$$\mathrm{C_2H_6 + O_2 \longrightarrow CO_2 + H_2O}$$

We have two carbon atoms on the left, so we need two carbon dioxide molecules on the product side, so that each side has two carbon atoms; that element is balanced. We have six hydrogen atoms in the reactants, so we need six hydrogen atoms in the products. We can get this by having three water molecules:

$$\mathrm{C_2H_6 + O_2 \longrightarrow 2\,CO_2 + 3\,H_2O}$$

Now we have seven oxygen atoms in the products (four from the CO2 and three from the H2O). This means we need seven oxygen atoms in the reactants. However, because oxygen is a diatomic molecule, we can only get an even number of oxygen atoms at a time. We can achieve this by multiplying the other coefficients by 2:

$$\mathrm{2\,C_2H_6 + O_2 \longrightarrow 4\,CO_2 + 6\,H_2O}$$

By multiplying everything else by 2, we do not unbalance the other elements, and we now get an even number of oxygen atoms in the product—14. We can get 14 oxygen atoms on the reactant side by having 7 oxygen molecules:

$$\mathrm{2\,C_2H_6 + 7\,O_2 \longrightarrow 4\,CO_2 + 6\,H_2O}$$

As a check, recount everything to determine that each side has the same number of atoms of each element. This chemical equation is now balanced.

Exercise 6.4.1

Write and balance the chemical equation that represents nitrogen and hydrogen reacting to produce ammonia, NH3.

Answer available at the end of this section.

Many chemical equations also include phase labels for the substances: (s) for solid, (ℓ) for liquid, (g) for gas, and (aq) for aqueous (i.e., dissolved in water). Special conditions, such as temperature, may also be listed above the arrow. For example,

$$\mathrm{2\,NaHCO_3\,(s) \overset{200^\circ C}{\rightarrow} Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(l)}$$

Key Takeaways

- A chemical equation is a concise description of a chemical reaction.
- Proper chemical equations are balanced.

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.
- 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.

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, wheat is produced or the result of what happens to the reactants when we put them together in the reaction vessel. If we think about baking chocolate chip cookies, our reactants would be flour, butter, sugar, vanilla, some baking soda, salt, egg, and chocolate chips. What would be the products? Cookies! The reaction vessel would be our mixing bowl.

$$\underbrace{\text{Flour} + \text{Butter} + \text{Sugar} + \text{Vanilla} + \text{Baking Soda} + \text{Eggs} + \text{Chocolate Chips}}_{\text{Ingredients} = \text{Reactants}} \rightarrow \underbrace{\text{Cookies}}_{\text{Product}}$$

Writing Chemical Equations

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

$$\underbrace{2\,\mathrm{SO_2\,(g)} + \mathrm{O_2\,(g)}}_{\text{Reactants}} \rightarrow \underbrace{2\,\mathrm{SO_3\,(g)}}_{\text{Products}}$$

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:

$$\text{Reactants} \rightarrow \text{Products}$$

There are a few special symbols that we need to know in order to "talk" in chemical shorthand. In the table below is the summary of the major symbols used in chemical equations. You will find there are others but these are the main ones that we need to know. Table $6.4.1$ shows a listing of symbols used in chemical equations.

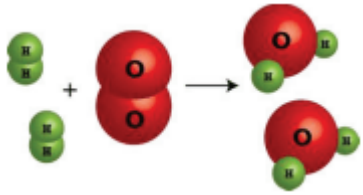
Table $6.4.1$ Symbols Used in Chemical Equations

Symbol	Description
$+$	used to separate multiple reactants or products

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

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

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



2. They could write a word equation for the chemical reaction:
"*Two molecules of hydrogen gas react with one molecule of oxygen gas to produce two molecules of water vapor.*"

3. They could write the equation in chemical shorthand.

$$2\text{H}_2\,(g) + \text{O}_2\,(g) \rightarrow 2\text{H}_2\text{O}\,(g)$$

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

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

$$\text{Ca}(\text{NO}_3)_2\,(aq) + 2\text{NaOH}\,(aq) \rightarrow \text{Ca}(\text{OH})_2\,(s) + 2\text{NaNO}_3\,(aq)$$

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:

$$\text{Cu}\,(s) + \text{AgNO}_3\,(aq) \rightarrow \text{Cu}(\text{NO}_3)_2\,(aq) + \text{Ag}\,(s)$$

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

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

1. Identify the reactants and products. This will help you know what symbols go on each side of the arrow and where the $+$ signs go.
2. Write the correct formulas for all compounds. You will need to use the rules you learned in Chapter 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 they are always found in pairs in nature. They include those elements listed in the table.

Table 6.4.2 *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 6.4.1

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

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

Solution

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

b. Reactants: propane ($\mathrm{C_3H_8}$) and oxygen ($\mathrm{O_2}$)

Product: carbon dioxide ($\mathrm{CO_2}$) and water ($\mathrm{H_2O}$)

$$\mathrm{C_3H_8}\,(g) + \mathrm{O_2}\,(g) \rightarrow \mathrm{CO_2}\,(g) + \mathrm{H_2O}\,(l)$$

c. Reactants: hydrogen fluoride and potassium carbonate

Products: potassium fluoride, water, and carbon dioxide

$$\mathrm{HF}\,(g) + \mathrm{K_2CO_3}\,(aq) \rightarrow \mathrm{KF}\,(aq) + \mathrm{H_2O}\,(l) + \mathrm{CO_2}\,(g)$$

Exercise 6.4.1

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

a. hydrogen gas reacts with nitrogen gas to produce gaseous ammonia
b. $\mathrm{HCl}\,(aq) + \mathrm{LiOH}\,(aq) \rightarrow \mathrm{NaCl}\,(aq) + \mathrm{H_2O}\,(l)$
c. copper metal is heated with oxygen gas to produce solid copper(II) oxide.

Answer a available at the end of this section.

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.

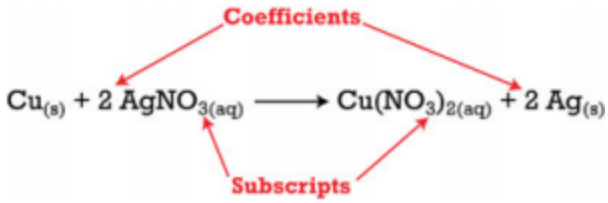


Figure 6.4.1Balancing Equations. You cannot change subscripts in a chemical formula to balance a chemical equation; you can change only the coefficients. Changing subscripts changes the ratios of atoms in the molecule and the resulting chemical properties. For example, water (H_2O) and hydrogen peroxide (H_2O_2) are chemically distinct substances.

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 6.4.2.

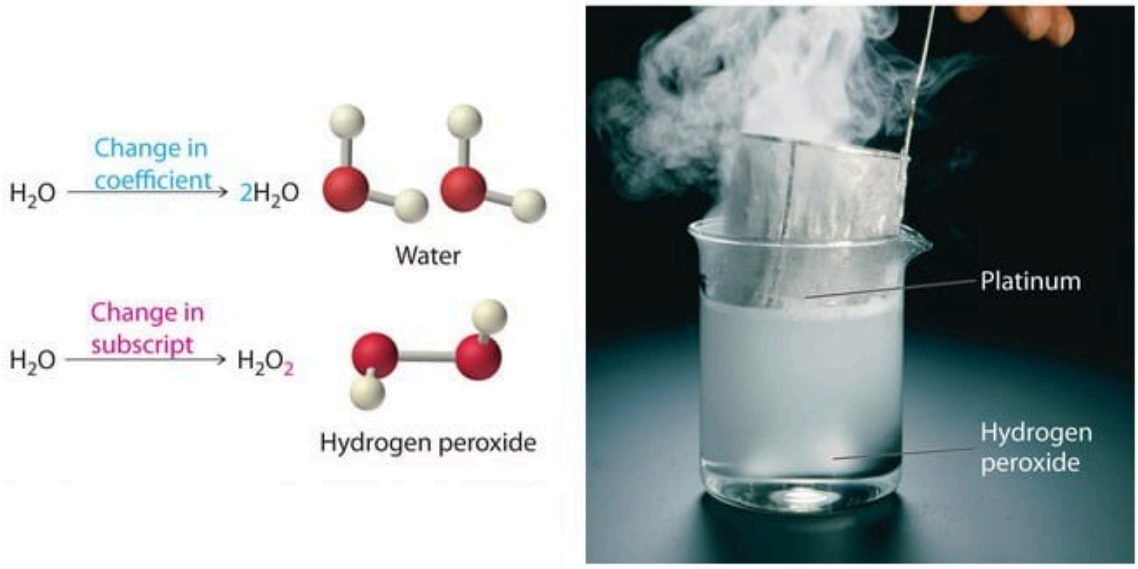


Figure 6.4.2 H_2O_2 decomposes to H_2O and O_2 gas when it comes in contact with the metal platinum (left), whereas no such reaction occurs between water and platinum (right).

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 6.4.2: Combustion of Heptane

Balancing the chemical Equation for the combustion of Heptane (C_7H_{16})

$$C_7H_{16}\,(l) + O_2\,(g) \longrightarrow CO_2\,(g) + H_2O\,(g)$$

Solution

Example 6.4.2: Steps for Problem Solving the Combustion Of Heptane

Steps	Explanation
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 assum 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 cont carbon atom, on the right side: $$C_7H_{16}\,(l) + O_2\,(g) \longrightarrow \underline{7}CO_2\,(g) + H_2O\,(g)$$ *Reactants and Products of Elements*

Reactants	Element/Polyatomic Ion	Produ
7	C	7

b. Because one molecule of n-heptane contains 16 hydrogen atoms, we need 8 H_2O molecules, each of which hydrogen atoms, on the right side:

$$C_7H_{16}\,(l) + O_2\,(g) \longrightarrow 7\,CO_2\,(g) + \underline{8}H_2O\,(g)$$

Reactants and Products of Elements

Reactants	Element/Polyatomic Ion	Produ
7	C	7
16	H	16

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 atoms on the left. We can balance the oxygen atoms by adjusting the coefficient in front of the least complex s O_2, on the reactant side: $$C_7H_{16}\,(l) + \underline{11}O_2\,(g) \longrightarrow 7\,CO_2\,(g) + 8\,H_2O\,(g)$$ *Reactants and Products of Elements*

Reactants	Element/Polyatomic Ion	Produ
7	C	7
16	H	16
22	O	22

5. **Check your work.**	The equation is now balanced, and there are no fractional coefficients: there are 7 carbon atoms, 16 hydrogen ato oxygen atoms on each side. Always check to be sure that a chemical equation is balanced.

Example 6.4.3: Combustion of Isooctane

Combustion of Isooctane (C_8H_{18})

$$C_8H_{18}\,(l) + O_2\,(g) \longrightarrow CO_2\,(g) + H_2O\,(g)$$

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.

Example 6.4.3: Solution for the combustion of isooctane

steps	examples
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 assum 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 must be 8 CO_2 molecules in the products: $$C_8H_{18}(l) + O_2(g) \longrightarrow \underline{8}CO_2(g) + H_2O(g)$$ *Reactants and Products of Elements* **Reactants** | **Element/Polyatomic Ion** | **Produ** 8 | C | 8 b. Eighteen hydrogen atoms in isooctane means that there must be 9 H_2O molecules in the products: $$C_8H_{18}(l) + O_2(g) \longrightarrow 8\,CO_2(g) + \underline{9}H_2O(g)$$ *Reactants and Products of Elements* **Reactants** | **Element/Polyatomic Ion** | **Produ** 8 | C | 8 18 | H | 18
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 atoms on the left. We can balance the least complex substance, O_2, but because there are 2 oxygen atoms per molecule, we must use a fractional coefficient ($\frac{25}{2}$) to balance the oxygen atoms: $$C_8H_{18}(l) + \underline{\frac{25}{2}}O_2\ (g) \longrightarrow 8\,CO_2(g) + 9\,H_2O(g)$$ *Reactants and Products of Elements* **Reactants** | **Element/Polyatomic Ion** | **Produ** 8 | C | 8 18 | H | 18 25 | O | 25 The equation is now balanced, but we usually write equations with whole-number coefficients. We can eliminat fractional coefficient by multiplying all coefficients on both sides of the chemical equation by **2**: $$\underline{2}C_8H_{18}(l) + \underline{25}O_2(g) \longrightarrow \underline{16}CO_2(g) + \underline{18}H_2O(g)$$

5. **Check your work.**	The balanced chemical equation has 16 carbon atoms, 36 hydrogen atoms, and 50 oxygen atoms on each side *Reactants and Products of Elements* **Reactants** | **Element/Polyatomic Ion** | **Produ** 16 | C | 16 36 | H | 36 50 | O | 50 Balancing equations requires some practice on your part as well as some common sense. If you find yourself u large coefficients or if you have spent several minutes without success, go back and make sure that you have w formulas of the reactants and products correctly.

Example 6.4.4: 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

Example 6.4.4: Solution for the combustion of isooctane

Steps for Problem Solving	**Example 6.4.4**
1. **Identify the most complex substance.**	The most complex substance is lead (II) chloride. $$\mathrm{Pb(NO_3)_2\,(aq) + NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$$
2. **Adjust the coefficients.**	There are twice as many chloride ions in the reactants than in the products. Place a 2 in front of the NaCl in or balance the chloride ions. $$\mathrm{Pb(NO_3)_2\,(aq) + \underline{2}NaCl\,(aq) \longrightarrow NaNO_3\,(aq) + PbCl_2\,(s)}$$ *Reactants and Products of Elements* **Reactants** | **Element/Polyatomic Ion** | **Produ** 1 | Pb | 1 2 | Na | 1 2 | Cl | 2

3. **Balance polyatomic ions as a unit.**	The nitrate ions are still unbalanced. Place a 2 in front of the $NaNO_3$. The result is: $$Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow \underline{2}NaNO_3\,(aq) + PbCl_2\,(s)$$ *Reactants and Products of Elements*

Reactants	Element/Polyatomic Ion	Produ
1	Pb	1
2	Na	2
2	Cl	2
2	NO_3^-	2

4. **Balance the remaining atoms.**	There is no need to balance the remaining atoms because they are already balanced.
5. **Check your work.**	$$Pb(NO_3)_2\,(aq) + 2\,NaCl\,(aq) \longrightarrow 2\,NaNO_3\,(aq) + PbCl_2\,(s)$$ *Reactants and Products of Elements*

Reactants	Element/Polyatomic Ion	Produ
1	Pb	1
2	Na	2
2	Cl	2
2	NO_3^-	2

Exercise 6.4.2

Is each chemical equation balanced?

a. $2\,Hg(\ell)^+\ O_2\,(g) \ \rightarrow Hg_2O_2\,(s)$
b. $C_2H_4\,(g) + 2\,O_2(g) \longrightarrow 2\,CO_2\,(g) + 2\,H_2O\,(g)$
c. $Mg(NO_3)_2\,(s) + 2\,Li\,(s) \ \rightarrow Mg(s)^+\ 2\,LiNO_3\,(s)$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Exercise 6.4.3

Balance the following chemical equations.

a. $N_2\,(g) + O_2\,(g) \longrightarrow NO_2\,(g)$
b. $Pb(NO_3)_2\,(aq) + FeCl_3\,(aq) \longrightarrow Fe(NO_3)_3\,(aq) + PbCl_2\,(s)$
c. $C_6H_{14}\,(l) + O_2(g) \longrightarrow CO_2\,(g) + H_2O\,(g)$

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Summary

- A chemical reaction is the process in which one or more substances are changed into one or more new substances.
- Chemical reactions are represented by chemical equations.
- Chemical equations have reactants on the left, an arrow that is read as "yields", and the products on the right.
- To be useful, chemical equations must always be balanced. Balanced chemical equations have the same number and type of each atom on both sides of the equation.
- The coefficients in a balanced equation must be the simplest whole number ratio. Mass is always conserved in chemical reactions.

Contributors and Attributions

-

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Learning Objectives

- Explain the law of conservation of mass.
- Explain the law of definite proportions.

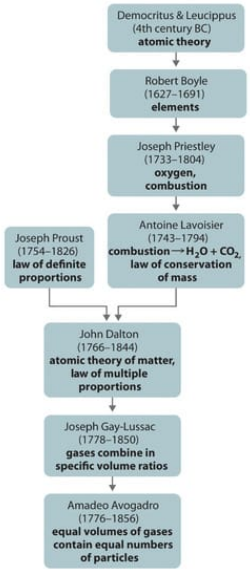


Figure 6.4.1: A summary of the historical development of the concept of the atom.

The Law of Conservation of Mass

It may seem as though burning destroys matter, but the same amount, or mass, of matter still exists after a campfire as before. Figure 6.4.2 (below) 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 assume that 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.

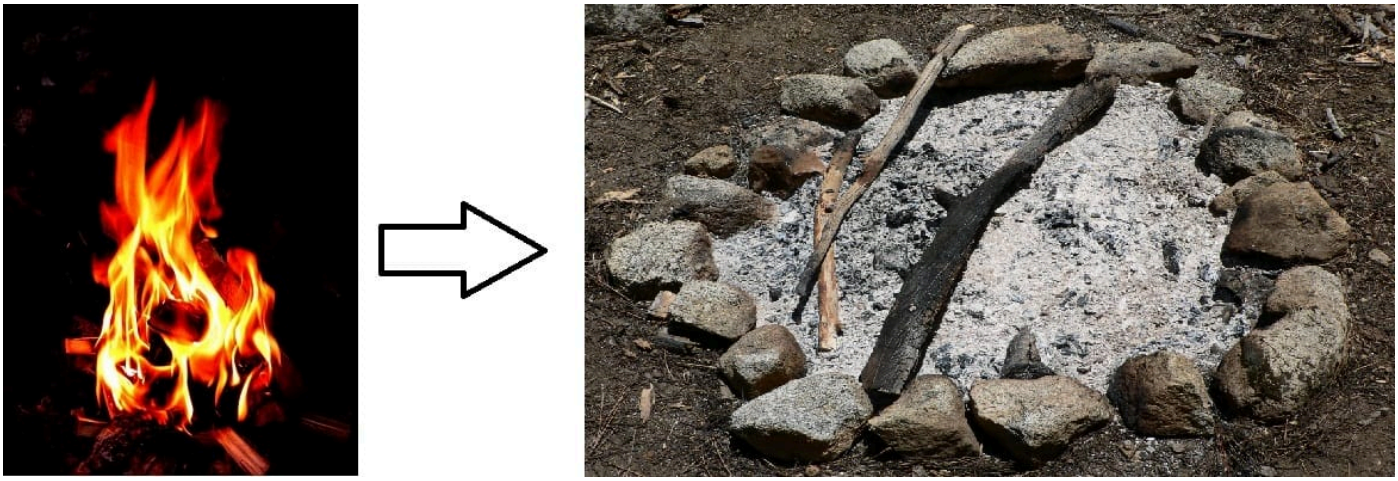


Figure 6.4.2: Burning is a chemical process. The flames are caused as a result of a fuel undergoing combustion (burning). Images used with permission (CC BY-SA 2.5; Einar Helland Berger for fire and for ash).

The law of conservation of mass was created in 1789 by French chemist Antoine Lavoisier. The **law of conservation of mass** states that matter cannot be created or destroyed in a chemical reaction. For example—in Figure 6.4.3—when wood burns, the mass of the soot, ashes, and gases equals the original mass of the wood and the oxygen when it first reacted.

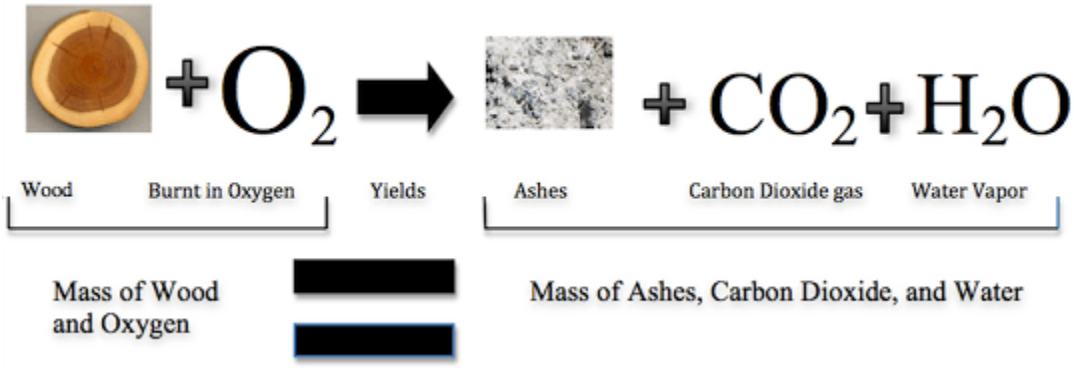


Figure 6.4.3: Image of the wood courtesy of Ehamberg and Stannered on Wikimedia Commons, available under Creative Commons Attribution 2.5 Generic license. Image of ashes courtesy of Walter Siegmund. Image as a whole constructed by Jessica Thornton (UCD).

So, the mass of the product(s) equals the mass of the reactants. Reactants are two or more elements chemically interacting to make a new substance; product is the substance(s) formed as the result of a chemical reaction (Video 6.4.1). Matter (and mass) may not be able to be created or destroyed, but can change forms to other substances like liquids, gases, solids, etc.

YouTube Video

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Video 6.4.1 This is a nice little demonstration showing the conservation of mass in action.

It's important to know the law well. If a 300 kg tree is burning down—when the process is complete—there are only ashes left, and all of them together weigh 10 kg. You may wonder where the other 290 kg went. The missing 290 kg was released into the atmosphere as smoke, so the only visible matter left is the 10 kg of ash. If you understand the law of conservation of mass, then you know that the other 290 kg had to go somewhere, because it had to equal the mass of the tree before it burnt down.

Example 6.4.1

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

$$\text{Mass of the reactants} = \text{Mass of the products}$$
$$10.0\,\text{g of } CaCO_3 = 4.4\,\text{g of } CO_2 + 5.6\,\text{g of CaO}$$
$$10.0\,\text{g of reactant} = 10.0\,\text{g of products}$$

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 6.4.1

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 react with 88.0 g of CO_2? The reaction also produces 36.0 g of water.

Answer available at the end of this section.

The Law of Definite Proportions

Joseph Proust (1754-1826) formulated the ***law of definite proportions*** (also called the **Law of Constant *Composition*** or **Proust's Law**). This law states that if a compound is broken down into its constituent elements, the masses of the constituents will always have the same proportions, regardless of the quantity or source of the original substance. Joseph Proust based this law primarily on his experiments with basic copper carbonate. The illustration below depicts this law in action.

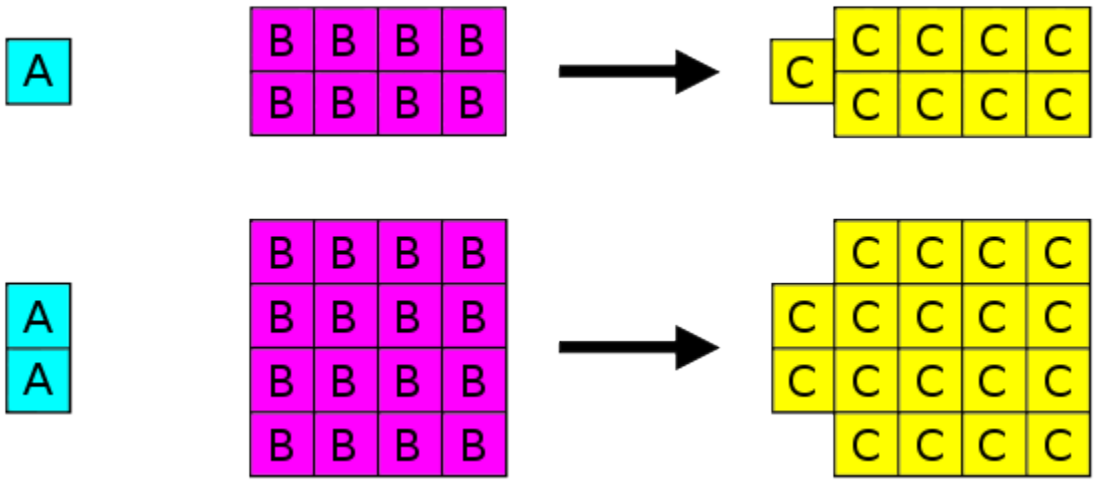


Figure 6.4.4: If 1 gram of A reacts with 8 grams of B, then by the Law of Definite Proportions, 2 grams of A must react with 16 grams of B. If 1 gram of A reacts with 8 grams of B, then by the Law of Conservation of Mass, they must produce 9 grams of C. Similarly, when 2 grams of A react with 16 grams of B, they must produce 18 grams of C.

Law of Definite Proportions states that in a given type of chemical substance, the elements are always combined in

the same proportions by mass.

The Law of Definite Proportions applies when elements are reacted together to form *the same* product. Therefore, while the Law of Definite Proportions can be used to compare two experiments in which hydrogen and oxygen react to form water, the Law of Definite Proportions can *not* be used to compare one experiment in which hydrogen and oxygen react to form water, and another experiment in which hydrogen and oxygen react to form hydrogen peroxide (peroxide is another material that can be made from hydrogen and oxygen).

Example 6.4.2: water

Oxygen makes up 88.8% of the mass of any sample of pure water, while hydrogen makes up the remaining 11.2% of the mass. You can get water by melting ice or snow; or by condensing steam from a river, sea, pond, etc. It can be from different places: USA, UK, Australia, or anywhere. Water can be made by chemical reactions, such as burning hydrogen in oxygen.

However, if the water is **pure**, it will **always** consist of 88.8 % oxygen by mass and 11.2 % hydrogen by mass, irrespective of its source or method of preparation.

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Video 6.4.2 Law of definite proportions.

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.
- The law of definite proportions states that a given chemical compound always contains the same elements in the exact same proportions by mass.

References

1. Petrucci, Ralph, William Harwood, Geoffrey Herring, and Jeffry Madura. General Chemistry. 9th ed. Upper Saddle River, New Jersey: Pearson Prentince Hall, 2007.
2. Moore, John. Chemistry for Dummies. John Wiley & Sons Inc, 2002.
3. Asimov, Isaac. A Short History of Chemistry. , CT.: Greenwood Press, 1965.
4. Patterson, Elizabeth C. John Dalton and the Atomic Theory. Garden City, NY: Doubleday, 1970.
5. Myers, Richard. The Basics of Chemistry. Greenwood, 2003.
6. Demtröder, Wolfgang. Atoms, Molecules and Photons: An Introduction to Atomic- Molecular- and Quantum Physics. 1st ed. Springer. 2002.

Contributors and Attributions

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Answers

Exercise 6.4.1

\[\ce{N2 + 3H2 → 2NH3 \nonumber \nonumber \]

Exercise 6.4.1

$H_2(g) + N_2(g) \rightarrow NH_3(g)$

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.

$Cu(s) + O_2(g) \rightarrow CuO(s)$

Exercise 6.4.2

yes

Exercise 6.4.2

no

Exercise 6.4.2

yes

Exercise 6.4.3

N_2 (g) + $2O_2$ (g) → $2NO_2$ (g)

Exercise 6.4.3

$3Pb(NO_3)_2$(aq) + $2FeCl_3$(aq) → $2Fe(NO_3)_3$(aq) + $3PbCl_2$(s)

Exercise 6.4.3

$2C_6H_{14}$(l) + $19O_2$(g)→ $12CO_2$(g) + $14H_2O$(g)

Exercise 6.4.1

276.4 g of potassium carbonate

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6.5: Types of Chemical Reactions



Figure 6.5.1 (CC-BY 2.0; Wisewire - "reaction of calcium carbonate with HCl" via Flickr)

Does the term *chemical reaction* bring to mind an image like this one? In the picture, a chemist is mixing chemicals in a lab. Many chemical reactions take place in labs. However, most chemical reactions do not. Where do they occur? They happen in the world all around you. They even happen inside your own body. In fact, you are alive only because of the many chemical reactions that constantly take place inside your cells.

What Is a Chemical Reaction?

A **chemical reaction** is a process in which some substances change into different substances. Substances that start a chemical reaction are called reactants. Substances that are produced in the reaction are called products. Reactants and products can be elements or compounds. Chemical reactions are represented by chemical equations, like the one below, in which reactants (on the left) are connected by an arrow to products (on the right).

$$\text{Reactants} \rightarrow \text{Products}$$

Chemical reactions may occur quickly or slowly. Look at the two pictures in the figure below. Both represent chemical reactions. In the picture on the left, a reaction inside a fire extinguisher causes foam to shoot out of the extinguisher. This reaction occurs almost instantly. In the picture on the right, a reaction causes the iron tool to turn to rust. This reaction occurs very slowly. In fact, it might take many years for all of the iron in the tool to turn to rust.



Figure 6.5.2 (CC BY 2.0; Fire extinguisher: Official U.S. Navy Page; Screw: Paulnasca via Fire extinguisher: http://www.flickr.com/photos/usnavy/7949633444; Screw: http://commons.wikimedia.org/wiki/File:Rust_screw.jpg)

Q: What happens during a chemical reaction? Where do the reactants go, and where do the products come from?

A: During a chemical reaction, chemical changes take place. Some chemical bonds break and new chemical bonds form.

Same Atoms, New Bonds

The reactants and products in a chemical reaction contain the same atoms, but they are rearranged during the reaction. As a result, the atoms are in different combinations in the products than they were in the reactants. This happens because chemical bonds break in the reactants and new chemical bonds form in the products.

Consider the chemical reaction in which water forms from oxygen and hydrogen gases. The figure below represents this reaction. Bonds break in molecules of hydrogen and oxygen, and then new bonds form in molecules of water. In both reactants and products there are four hydrogen atoms and two oxygen atoms, but the atoms are combined differently in water.

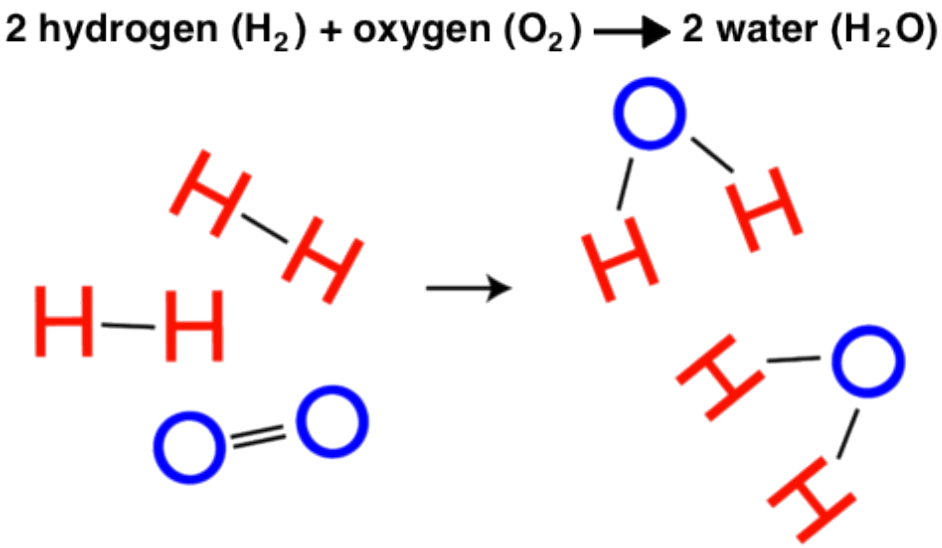


Figure 6.5.3: (CC BY-NC 3.0; CK-12 Foundation via CK-12 Foundation)

Types of Chemical Reactions

The chemical reaction in the figure above, in which water forms from hydrogen and oxygen, is an example of a synthesis reaction. In this type of reaction, two or more reactants combine to synthesize a single product. There are several other types of chemical reactions, including decomposition, replacement, and combustion reactions. The table below compares these four types of chemical reactions.

Four Types of Chemical Reactions

Type of Reaction	General Equation	Example
Synthesis	A+B → C	$2Na + Cl_2 \rightarrow 2NaCl$
Decomposition	AB → A + B	$2H_2O \rightarrow 2H_2 + O_2$
Single Replacement	A+BC → B+ AC	$2K + 2H_2O \rightarrow 2KOH + H_2$
Double Replacement	AB+ CD → AD + CB	NaCl+ AgF → NaF + AgCl
Combustion	fuel + oxygen → carbon dioxide + water	$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Q: The burning of wood is a chemical reaction. Which type of reaction is it?

A: The burning of wood—or of anything else—is a combustion reaction. In the combustion example in the table, the fuel is methane gas (CH_4).

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Energy In and Energy Out

All chemical reactions involve energy. Energy is used to break bonds in reactants, and energy is released when new bonds form in products. In terms of energy, there are two types of chemical reactions: endothermic reactions and exothermic reactions.

- In exothermic reactions, more energy is released when bonds form in products than is used to break bonds in reactants. These reactions release energy to the environment, often in the form of heat or light.
- In endothermic reactions, more energy is used to break bonds in reactants than is released when bonds form in products. These reactions absorb energy from the environment.

Q: When it comes to energy, which type of reaction is the burning of wood? Is it an endothermic reaction or an exothermic reaction? How can you tell?

A: The burning of wood is an exothermic reaction. You can tell by the heat and light energy given off by a wood fire.

Summary

- A chemical reaction is a process in which some substances, called reactants, change into different substances, called products. During the reaction, chemical bonds break in the reactants and new chemical bonds form in the products.
- Types of chemical reactions include synthesis, decomposition, replacement, and combustion reactions.
- All chemical reactions involve energy. Exothermic reactions release more energy than they use. Endothermic reactions use more energy than they release.

Review

1. What is a chemical reaction?
2. Write a general chemical equation that shows the relationship of products to reactants in a chemical reaction.
3. Contrast exothermic and endothermic chemical reactions.

Learning Objectives

- Recognize composition, decomposition, and combustion reactions.
- Predict the products of a combustion reaction.

Three classifications of chemical reactions will be reviewed in this section. Predicting the products in some of them may be difficult, but the reactions are still easy to recognize.

A **composition reaction** (sometimes also called a *combination reaction* or a *synthesis reaction*) produces a single substance from multiple reactants. A single substance as a product is the key characteristic of the composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction. In the reaction

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

water is produced from hydrogen and oxygen. Although there are two molecules of water being produced, there is only one substance—water—as a product. So this is a composition reaction.

A **decomposition reaction** starts from a single substance and produces more than one substance; that is, it decomposes. The key characteristics of a decomposition reaction are: one substance as a reactant and more than one substance as the products. For example, in the decomposition of sodium hydrogen carbonate (also known as sodium bicarbonate):

$$2\,NaHCO_3\,(s) \longrightarrow Na_2CO_3\,(s) + CO_2\,(g) + H_2O(\ell)$$

sodium carbonate, carbon dioxide, and water are produced from the single substance sodium hydrogen carbonate.

Composition and decomposition reactions are difficult to predict; however, they should be easy to recognize.

Example 6.5.1: Identifying Reactions

Identify each equation as a composition reaction, a decomposition reaction, or neither.

a. $Fe_2O_3 + 3\,SO_3 \longrightarrow Fe_2(SO_4)_3$
b. $NaCl + AgNO_3 \longrightarrow AgCl + NaNO_3$
c. $(NH_4)_2Cr_2O_7 \longrightarrow Cr_2O_3 + 4\,H_2O + N_2$

Solution

a. In this equation, two substances combine to make a single substance. This is a composition reaction.
b. Two different substances react to make two new substances. This does not fit the definition of either a composition reaction or a decomposition reaction, so it is neither. In fact, you may recognize this as a double-replacement reaction.
c. A single substance reacts to make multiple substances. This is a decomposition reaction.

Exercise 6.5.1

Identify the equation as a composition reaction, a decomposition reaction, or neither.

$$C_3H_8 \longrightarrow C_3H_4 + 2\,H_2$$

Answer available at the end of this section.

A **combustion reaction** occurs when a reactant combines with oxygen, many times from the atmosphere, to produce oxides of all other elements as products; any nitrogen in the reactant is converted to elemental nitrogen, N_2. Many reactants, called *fuels*, contain mostly carbon and hydrogen atoms, reacting with oxygen to produce CO_2 and H_2O. For example, the balanced chemical equation for the combustion of methane, CH_4, is as follows:

$$\mathrm{CH_4 + 2\,O_2 \longrightarrow CO_2 + 2\,H_2O}$$

Kerosene can be approximated with the formula $\mathrm{C_{12}H_{26}}$, and its combustion equation is:

$$\mathrm{2\,C_{12}H_{26} + 37\,O_2 \longrightarrow 24\,CO\text{-}2 + 26\,H_2O}$$

Sometimes fuels contain oxygen atoms, which must be counted when balancing the chemical equation. One common fuel is ethanol, $\mathrm{C_2H_5OH}$, whose combustion equation is:

$$\mathrm{C_2H_5OH + 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O}$$

If nitrogen is present in the original fuel, it is converted to $\mathrm{N_2}$, not to a nitrogen-oxygen compound. Thus, for the combustion of the fuel dinitroethylene, whose formula is $\mathrm{C_2H_2N_2O_4}$, we have:

$$\mathrm{2\,C_2H_2N_2O_4 + O_2 \longrightarrow 4\,CO_2 + 2\,H_2O + 2\,N_2}$$

Example 6.5.2: Combustion Reactions

Complete and balance each combustion equation.

1. the combustion of propane ($\mathrm{C_3H_8}$)
2. the combustion of ammonia ($\mathrm{NH_3}$)

Solution

1. The products of the reaction are CO2 and H2O, so our unbalanced equation is

   $$\mathrm{C_3H_8 + O_2 \longrightarrow CO_2 + H_2O}$$

   Balancing (and you may have to go back and forth a few times to balance this), we get

   $$\mathrm{C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O}$$

2. The nitrogen atoms in ammonia will react to make N2, while the hydrogen atoms will react with O2 to make H2O:

   $$\mathrm{NH_3 + O_2 \longrightarrow N_2 + H_2O}$$

   To balance this equation without fractions (which is the convention), we get

   $$\mathrm{4\,NH_3 + 3\,O_2 \longrightarrow 2\,N_2 + 6\,H_2O}$$

Exercise 6.5.2

Complete and balance the combustion equation for cyclopropanol (\ce{C3H6O}\)).

Answer available at the end of this section.

Key Takeaways

- A composition reaction produces a single substance from multiple reactants.
- A decomposition reaction produces multiple products from a single reactant.
- Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products (although nitrogen atoms react to make N2).

Learning Objectives

- Recognize composition, decomposition, and combustion reactions.

- Predict the products of a combustion reaction.

Three classifications of chemical reactions will be reviewed in this section. Predicting the products in some of them may be difficult, but the reactions are still easy to recognize.

A **composition reaction** (sometimes also called a *combination reaction* or a *synthesis reaction*) produces a single substance from multiple reactants. A single substance as a product is the key characteristic of the composition reaction. There may be a coefficient other than one for the substance, but if the reaction has only a single substance as a product, it can be called a composition reaction. In the reaction

$$2\,\mathrm{H_2\,(g)} + \mathrm{O_2\,(g)} \longrightarrow 2\,\mathrm{H_2O}(\ell)$$

water is produced from hydrogen and oxygen. Although there are two molecules of water being produced, there is only one substance—water—as a product. So this is a composition reaction.

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$$2\,\mathrm{NaHCO_3\,(s)} \longrightarrow \mathrm{Na_2CO_3\,(s)} + \mathrm{CO_2\,(g)} + \mathrm{H_2O}(\ell)$$

sodium carbonate, carbon dioxide, and water are produced from the single substance sodium hydrogen carbonate.

Composition and decomposition reactions are difficult to predict; however, they should be easy to recognize.

Example 6.5.1: Identifying Reactions

Identify each equation as a composition reaction, a decomposition reaction, or neither.

a. $\mathrm{Fe_2O_3} + 3\,\mathrm{SO_3} \longrightarrow \mathrm{Fe_2(SO_4)_3}$
b. $\mathrm{NaCl} + \mathrm{AgNO_3} \longrightarrow \mathrm{AgCl} + \mathrm{NaNO_3}$
c. $\mathrm{(NH_4)_2Cr_2O_7} \longrightarrow \mathrm{Cr_2O_3} + 4\,\mathrm{H_2O} + \mathrm{N_2}$

Solution

a. In this equation, two substances combine to make a single substance. This is a composition reaction.
b. Two different substances react to make two new substances. This does not fit the definition of either a composition reaction or a decomposition reaction, so it is neither. In fact, you may recognize this as a double-replacement reaction.
c. A single substance reacts to make multiple substances. This is a decomposition reaction.

Exercise 6.5.1

Identify the equation as a composition reaction, a decomposition reaction, or neither.

$$\mathrm{C_3H_8} \longrightarrow \mathrm{C_3H_4} + 2\,\mathrm{H_2}$$

Answer available at the end of this section.

A **combustion reaction** occurs when a reactant combines with oxygen, many times from the atmosphere, to produce oxides of all other elements as products; any nitrogen in the reactant is converted to elemental nitrogen, N_2. Many reactants, called *fuels*, contain mostly carbon and hydrogen atoms, reacting with oxygen to produce CO_2 and H_2O. For example, the balanced chemical equation for the combustion of methane, CH_4, is as follows:

$$\mathrm{CH_4} + 2\,\mathrm{O_2} \longrightarrow \mathrm{CO_2} + 2\,\mathrm{H_2O}$$

Kerosene can be approximated with the formula $\mathrm{C_{12}H_{26}}$, and its combustion equation is:

$$2\,C_{12}H_{26} + 37\,O_2 \longrightarrow 24\,CO\text{-}2 + 26\,H_2O$$

Sometimes fuels contain oxygen atoms, which must be counted when balancing the chemical equation. One common fuel is ethanol, C_2H_5OH, whose combustion equation is:

$$C_2H_5OH + 3\,O_2 \longrightarrow 2\,CO_2 + 3\,H_2O$$

If nitrogen is present in the original fuel, it is converted to N_2, not to a nitrogen-oxygen compound. Thus, for the combustion of the fuel dinitroethylene, whose formula is $C_2H_2N_2O_4$, we have:

$$2\,C_2H_2N_2O_4 + O_2 \longrightarrow 4\,CO_2 + 2\,H_2O + 2\,N_2$$

Example 6.5.2: Combustion Reactions

Complete and balance each combustion equation.

1. the combustion of propane (C_3H_8)
2. the combustion of ammonia (NH_3)

Solution

1. The products of the reaction are CO2 and H2O, so our unbalanced equation is

   $$C_3H_8 + O_2 \longrightarrow CO_2 + H_2O$$

   Balancing (and you may have to go back and forth a few times to balance this), we get

   $$C_3H_8 + 5\,O_2 \longrightarrow 3\,CO_2 + 4\,H_2O$$

2. The nitrogen atoms in ammonia will react to make N2, while the hydrogen atoms will react with O2 to make H2O:

   $$NH_3 + O_2 \longrightarrow N_2 + H_2O$$

   To balance this equation without fractions (which is the convention), we get

   $$4\,NH_3 + 3\,O_2 \longrightarrow 2\,N_2 + 6\,H_2O$$

Exercise 6.5.2

Complete and balance the combustion equation for cyclopropanol (\ce{C3H6O}\)).

Answer available at the end of this section.

Key Takeaways

- A composition reaction produces a single substance from multiple reactants.
- A decomposition reaction produces multiple products from a single reactant.
- Combustion reactions are the combination of some compound with oxygen to make oxides of the other elements as products (although nitrogen atoms react to make N2).

Answers

Exercise 6.5.1

decomposition

Exercise 6.5.2

$$C_3H_6O + 4\,O_2 \longrightarrow 3\,CO_2 + 3\,H_2O$$



Figure 6.5.1 A 20 lb (9.1 kg) steel propane cylinder.

Exercise 6.5.1

decomposition

Exercise 6.5.2

$$C_3H_6O + 4\,O_2 \longrightarrow 3\,CO_2 + 3\,H_2O$$



Figure 6.5.1 A 20 lb (9.1 kg) steel propane cylinder.

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6.6: Predicting Products of Double Replacement Reactions

Learning Objectives

- Recognize chemical reactions as single-replacement reactions and double-replacement reactions.
- Use the periodic table, an activity series, or solubility rules to predict whether single-replacement reactions or double-replacement reactions will occur.

Up until now, we have presented chemical reactions as a topic, but we have not discussed how the products of a chemical reaction can be predicted. Here we will begin our study of certain types of chemical reactions that allow us to predict what the products of the reaction will be.

A **single-replacement reaction** is a chemical reaction in which one element is substituted for another element in a compound, generating a new element and a new compound as products. Presented below:

$$2\,\mathrm{HCl}\,(\mathrm{aq}) + \mathrm{Zn}\,(\mathrm{s}) \longrightarrow \mathrm{ZnCl_2}\,(\mathrm{aq}) + \mathrm{H_2}\,(\mathrm{g})$$

is an example of a single-replacement reaction. The hydrogen atoms in HCl are replaced by Zn atoms, and in the process a new element—hydrogen—is formed. Another example of a single-replacement reaction is

$$2\,\mathrm{NaCl}\,(\mathrm{aq}) + \mathrm{F_2}\,(\mathrm{g}) \longrightarrow 2\,\mathrm{NaF}\,(\mathrm{s}) + \mathrm{Cl_2}\,(\mathrm{g})$$

Here the negatively charged ion changes from chloride to fluoride. A typical characteristic of a single-replacement reaction is that there is one element as a reactant and another element as a product.

Not all proposed single-replacement reactions will occur between two given reactants. This is most easily demonstrated with fluorine, chlorine, bromine, and iodine. Collectively, these elements are called the *halogens* and are in the next-to-last column on the periodic table (Figure $6.6.1$). The elements on top of the column will replace the elements below them on the periodic table, but not the other way around. Thus, the reaction represented by

$$\mathrm{CaI_2}\,(\mathrm{s}) + \mathrm{Cl_2}\,(\mathrm{g}) \longrightarrow \mathrm{CaCl_2}\,(\mathrm{s}) + \mathrm{I_2}\,(\mathrm{s})$$

will occur; but the reaction

$$\mathrm{CaF_2}\,(\mathrm{s}) + \mathrm{Br_2}(\ell) \longrightarrow \mathrm{CaBr_2}\,(\mathrm{s}) + \mathrm{F_2}\,(\mathrm{g})$$

will not, because bromine is below fluorine on the periodic table. This is just one of many ways the periodic table helps us to understand chemistry.

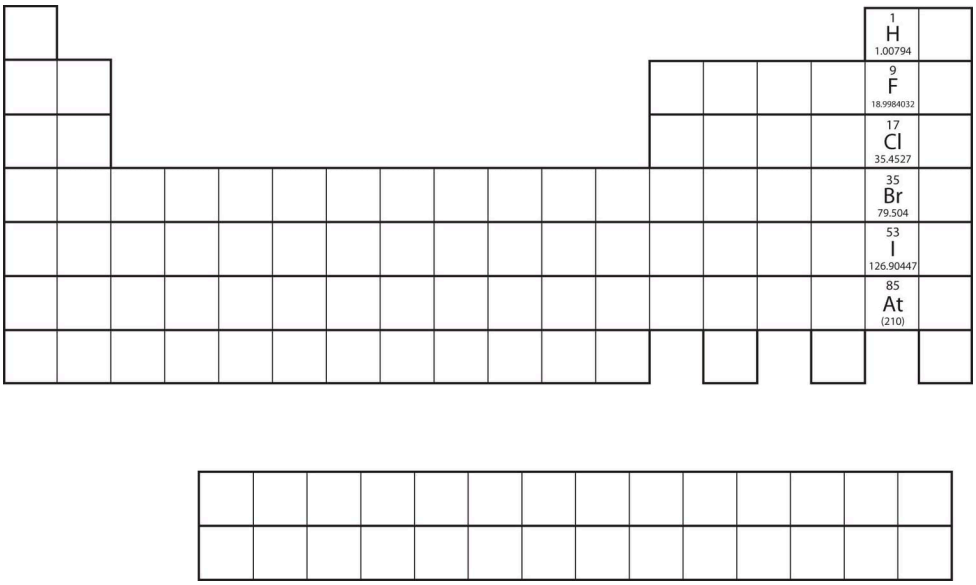


Figure 6.6.1: Halogens on the Periodic Table. The halogens are the elements in the next-to-last column on the periodic table.

Example 6.6.1

Will a single-replacement reaction occur? If so, identify the products.

a. $MgCl_2 + I_2 \rightarrow ?$
b. $CaBr_2 + F_2 \rightarrow ?$

Solution

a. Because iodine is below chlorine on the periodic table, a single-replacement reaction will not occur.
b. Because fluorine is above bromine on the periodic table, a single-replacement reaction will occur, and the products of the reaction will be CaF_2 and Br_2.

Exercise 6.6.1

Will a single-replacement reaction occur? If so, identify the products.

$$FeI_2 + Cl_2 \longrightarrow$$

Answer available at the end of this section.

Chemical reactivity trends are easy to predict when replacing anions in simple ionic compounds—simply use their relative positions on the periodic table. However, when replacing the cations, the trends are not as straightforward. This is partly because there are so many elements that can form cations; an element in one column on the periodic table may replace another element nearby, or it may not. A list called the **activity series** does the same thing the periodic table does for halogens: it lists the elements that will replace elements below them in single-replacement reactions. A simple activity series is shown below.

Activity Series for Cation Replacement in Single-Replacement Reactions

- Li
- K
- Ba
- Sr

- Ca
- Na
- Mg
- Al
- Mn
- Zn
- Cr
- Fe
- Ni
- Sn
- Pb
- H_2
- Cu
- Hg
- Ag
- Pd
- Pt
- Au

Using the activity series is similar to using the positions of the halogens on the periodic table. An element on top will replace an element below it in compounds undergoing a single-replacement reaction. Elements will not replace elements above them in compounds.

Example 6.6.2

Use the activity series to predict the products, if any, of each equation.

1. $FeCl_2 + Zn \rightarrow ?$
2. $HNO_3 + Au \rightarrow ?$

Solution

1. Because zinc is above iron in the activity series, it will replace iron in the compound. The products of this single-replacement reaction are $ZnCl_2$ and Fe.
2. Gold is below hydrogen in the activity series. As such, it will not replace hydrogen in a compound with the nitrate ion. No reaction is predicted.

Exercise 6.6.2

Use the activity series to predict the products, if any, of this equation.

$$\mathrm{AlPO_4 + Mg \longrightarrow}$$

Answer available at the end of this section.

A **double-replacement reaction** occurs when parts of two ionic compounds are exchanged, making two new compounds. A characteristic of a double-replacement equation is that there are two compounds as reactants and two different compounds as products. An example is

$$\mathrm{CuCl_2\,(aq) + 2\,AgNO_3\,(aq) \longrightarrow Cu(NO_3)_2\,(aq) + 2\,AgCl\,(s)}$$

There are two equivalent ways of considering a double-replacement equation: either the cations are swapped, or the anions are swapped. (You cannot swap both; you would end up with the same substances you started with.) Either perspective should allow you to predict the proper products, as long as you pair a cation with an anion, and not a cation with a cation or an anion with an anion.

Example 6.6.3

Predict the products of this double-replacement equation:

$$\mathrm{BaCl_2 + Na_2SO_4 \longrightarrow}$$

Solution

Thinking about the reaction as either switching the cations or switching the anions, we would expect the products to be $BaSO_4$ and NaCl.

Exercise 6.6.3

Predict the products of this double-replacement equation:

$$\mathrm{KBr + AgNO_3 \longrightarrow}$$

Answer available at the end of this section.

Predicting whether a double-replacement reaction occurs is somewhat more difficult than predicting a single-replacement reaction. However, there is one type of double-replacement reaction that we can predict: the precipitation reaction. A **precipitation reaction** occurs when two ionic compounds are dissolved in water and form a new ionic compound that does not dissolve; this new compound falls out of solution as a solid **precipitate.** The formation of a solid precipitate is the driving force that makes the reaction proceed.

To judge whether double-replacement reactions will occur, we need to know what kinds of ionic compounds form precipitates. For this, we use **solubility rules**, which are general statements that predict which ionic compounds dissolve (are soluble) and which do not (are not soluble, or insoluble). Table 6.6.1 lists some general solubility rules. We need to consider each ionic compound (both the reactants and the possible products) in light of the solubility rules. If a compound is soluble, we use the (aq) label with it, indicating that it dissolves. If a compound is not soluble, we use the (s) label with it and assume that it will precipitate out of solution. If everything is soluble, then no reaction will be expected.

Table 6.6.1: Some Useful Solubility Rules (soluble)

These compounds generally dissolve in water (are soluble):	**Exceptions:**
All compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+	None
All compounds of NO_3^- and $C_2H_3O_2^-$	None
Compounds of Cl^-, Br^-, I^-	Ag^+, Hg_2^{2+}, Pb^{2+}
Compounds of SO_4^{2}	Hg_2^{2+}, Pb^{2+}, Sr^{2+}, Ba^{2+}

Table 6.6.2: Some Useful Solubility Rules (insoluble)

These compounds generally do not dissolve in water (are insoluble):	**Exceptions:**
Compounds of CO_3^{2-} and PO_4^{3-}	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, and NH_4^+
Compounds of OH^-	Compounds of Li^+, Na^+, K^+, Rb^+, Cs^+, NH_4^+, Sr^{2+}, and Ba^{2+}

For example, consider the possible double-replacement reaction between Na_2SO_4 and $SrCl_2$. The solubility rules say that all

ionic sodium compounds are soluble and all ionic chloride compounds are soluble, except for Ag^{+}, Hg_2^{2+}, and Pb^{2+}, which are not being considered here. Therefore, Na_2SO_4 and $SrCl_2$ are both soluble. The possible double-replacement reaction products are NaCl and $SrSO_4$. Are these soluble? NaCl is (by the same rule we just quoted), but what about $SrSO_4$? Compounds of the sulfate ion are generally soluble, but Sr^{2+} is an exception: we expect it to be insoluble—a precipitate. Therefore, we expect a reaction to occur, and the balanced chemical equation would be:

$$\mathrm{Na_2SO_4\,(aq) + SrCl_2\,(aq) \longrightarrow 2\,NaCl\,(aq) + SrSO_4\,(s)}$$

You would expect to see a visual change corresponding to $SrSO_4$ precipitating out of solution (Figure 6.6.2).



Figure 6.6.2: Double-Replacement Reactions. Some double-replacement reactions are obvious because you can see a solid precipitate coming out of solution. Source: Photo courtesy of Choij, http://commons.wikimedia.org/wiki/File:Copper_solution.jpg(opens in new window).

Example 6.6.4

Will a double-replacement reaction occur? If so, identify the products.

1. $Ca(NO_3)_2$ + KBr → ?
2. NaOH + $FeCl_2$ → ?

Solution

1. According to the solubility rules, both $Ca(NO_3)_2$ and KBr are soluble. Now we consider what the double-replacement products would be by switching the cations (or the anions)—namely, $CaBr_2$ and KNO_3. However, the solubility rules predict that these two substances would also be soluble, so no precipitate would form. Thus, we predict no reaction in this case.
2. According to the solubility rules, both NaOH and $FeCl_2$ are expected to be soluble. If we assume that a double-replacement reaction may occur, we need to consider the possible products, which would be NaCl and $Fe(OH)_2$. NaCl is soluble, but, according to the solubility rules, $Fe(OH)_2$ is not. Therefore, a reaction would occur, and $Fe(OH)_2$(s) would precipitate out of solution. The balanced chemical equation is

$$\mathrm{2\,NaOH\,(aq) + FeCl_2\,(aq) \longrightarrow 2\,NaCl\,(aq) + Fe(OH)_2\,(s)}$$

Exercise 6.6.4

$$\mathrm{Sr(NO_3)_2 + KCl \longrightarrow}$$

Answer available at the end of this section.

Key Takeaways

- A single-replacement reaction replaces one element for another in a compound.
- The periodic table or an activity series can help predict whether single-replacement reactions occur.
- A double-replacement reaction exchanges the cations (or the anions) of two ionic compounds.
- A precipitation reaction is a double-replacement reaction in which one product is a solid precipitate.
- Solubility rules are used to predict whether some double-replacement reactions will occur.

Learning Objectives

- Identify an acid and a base.
- Identify a neutralization reaction and predict its products.

In Chapter 3, we defined an acid as an ionic compound that contains H^+ as the cation. This is slightly incorrect, but until additional concepts were developed, a better definition needed to wait. Now we can redefine an acid: an **acid** is any compound that increases the amount of hydrogen ion (H^+) in an aqueous solution. The chemical opposite of an acid is a base. The equivalent definition of a base is that a **base** is a compound that increases the amount of hydroxide ion (OH^-) in an aqueous solution. These original definitions were proposed by Arrhenius (the same person who proposed ion dissociation) in 1884, so they are referred to as the **Arrhenius definition** of an acid and a base, respectively.

You may recognize that, based on the description of a hydrogen atom, an H^+ ion is a hydrogen atom that has lost its lone electron; that is, H^+ is simply a proton. Do we really have bare protons moving about in aqueous solution? No. What is more likely is that the H^+ ion has attached itself to one (or more) water molecule(s). To represent this chemically, we define the **hydronium ion** as H_3O^+, which represents an additional proton attached to a water molecule. We use the hydronium ion as the more logical way that a hydrogen ion appears in an aqueous solution, although in many chemical reactions H^+ and H_3O^+ are treated equivalently.

The reaction of an acid and a base is called a **neutralization reaction**. Although acids and bases have their own unique chemistries, the acid and base cancel each other's chemistry to produce a rather innocuous substance—water. In fact, the general reaction between an acid and a base is:

$$\text{acid} + \text{base} \longrightarrow \text{water} + \text{salt}$$

where the term **salt** is generally used to define any ionic compound (soluble or insoluble) that is formed from a reaction between an acid and a base. (In chemistry, the word *salt* refers to more than just table salt.) For example, the balanced chemical equation for the reaction between HCl(aq) and KOH(aq) is

$$\mathrm{HCl\,(aq)} + \mathrm{KOH\,(aq)} \longrightarrow \mathrm{H_2O(\ell)} + \mathrm{KCl\,(aq)}$$

where the salt is KCl. By counting the number of atoms of each element, we find that only one water molecule is formed as a product. However, in the reaction between HCl(aq) and $Mg(OH)_2$(aq), additional molecules of HCl and H_2O are required to balance the chemical equation:

$$2\,\mathrm{HCl\,(aq)} + \mathrm{Mg(OH)_2\,(aq)} \longrightarrow 2\,\mathrm{H_2O(\ell)} + \mathrm{MgCl_2\,(aq)}$$

Here, the salt is $MgCl_2$. (This is one of several reactions that take place when a type of antacid—a base—is used to treat stomach acid.)

Example 6.6.1

Write the neutralization reactions between each acid and base.

a. $HNO_3(aq)$ and $Ba(OH)_2(aq)$
b. $H_3PO_4(aq)$ and $Ca(OH)_2(aq)$

Solution

First, we will write the chemical equation with the formulas of the reactants and the expected products; then we will balance the equation.

a. The expected products are water and barium nitrate, so the initial chemical reaction is

$$\mathrm{HNO_3\,(aq) + Ba(OH)_2\,(aq) \longrightarrow H_2O(\ell) + Ba(NO_3)_2\,(aq)}$$

To balance the equation, we need to realize that there will be two H_2O molecules, so two HNO_3 molecules are required:

$$\mathrm{2\,HNO_3\,(aq) + Ba(OH)_2\,(aq) \longrightarrow 2\,H_2O(\ell) + Ba(NO_3)_2\,(aq)}$$

This chemical equation is now balanced.

b. The expected products are water and calcium phosphate, so the initial chemical equation is

$$\mathrm{H_3PO_4\,(aq) + Ca(OH)_2\,(aq) \longrightarrow H_2O(\ell) + Ca_3(PO_4)_2\,(s)}$$

According to the solubility rules, $Ca_3(PO_4)_2$ is insoluble, so it has an (s) phase label. To balance this equation, we need two phosphate ions and three calcium ions. We end up with six water molecules to balance the equation:

$$\mathrm{2\,H_3PO_4\,(aq) + 3\,Ca(OH)_2\,(aq) \longrightarrow 6\,H_2O(\ell) + Ca_3(PO_4)_2\,(s)}$$

This chemical equation is now balanced.

Exercise 6.6.1

Write the neutralization reaction between $H_2SO_4(aq)$ and $Sr(OH)_2(aq)$.

Answer

$$\mathrm{H_2SO_4\,(aq) + Sr(OH)_2\,(aq) \longrightarrow 2\,H_2O(\ell) + SrSO_4\,(aq)}$$

Neutralization reactions are one type of chemical reaction that proceeds even if one reactant is not in the aqueous phase. For example, the chemical reaction between $HCl(aq)$ and $Fe(OH)_3(s)$ still proceeds according to the equation:

$$\mathrm{3\,HCl\,(aq) + Fe(OH)_3\,(s) \longrightarrow 3\,H_2O(\ell) + FeCl_3\,(aq)}$$

even though $Fe(OH)_3$ is not soluble. When one realizes that $Fe(OH)_3(s)$ is a component of rust, this explains why some cleaning solutions for rust stains contain acids—the neutralization reaction produces products that are soluble and wash away. (Washing with acids like HCl is one way to remove rust and rust stains, but HCl must be used with caution!)

Complete and net ionic reactions for neutralization reactions will depend on whether the reactants and products are soluble, even if the acid and base react. For example, in the reaction of $HCl(aq)$ and $NaOH(aq)$,

$$\mathrm{HCl\,(aq) + NaOH\,(aq) \longrightarrow H_2O(\ell) + NaCl\,(aq)}$$

the complete ionic reaction is

$$\mathrm{H^+\,(aq) + Cl^-\,(aq) + Na^+\,(aq) + OH^-\,(aq) \longrightarrow H_2O(\ell) + Na^+\,(aq) + Cl^-\,(aq)}$$

The $Na^+(aq)$ and $Cl^-(aq)$ ions are spectator ions, so we can remove them to have

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(\ell)$$

as the net ionic equation. If we wanted to write this in terms of the hydronium ion, $H_3O^+(aq)$, we would write it as

$$H_3O^+(aq) + OH^-(aq) \longrightarrow 2\,H_2O(\ell)$$

With the exception of the introduction of an extra water molecule, these two net ionic equations are equivalent.

However, for the reaction between HCl(aq) and Cr(OH)2(s), because chromium(II) hydroxide is insoluble, we cannot separate it into ions for the complete ionic equation:

$$2\,H^+(aq) + 2\,Cl^-(aq) + Cr(OH)_2(s) \longrightarrow 2\,H_2O(\ell) + Cr^{2+}(aq) + 2\,Cl^-(aq)$$

The chloride ions are the only spectator ions here, so the net ionic equation is

$$2\,H^+(aq) + Cr(OH)_2(s) \longrightarrow 2\,H_2O(\ell) + Cr^{2+}(aq)$$

Example 6.6.2

Oxalic acid, $H_2C_2O_4(s)$, and $Ca(OH)_2(s)$ react very slowly. What is the net ionic equation between these two substances if the salt formed is insoluble? (The anion in oxalic acid is the oxalate ion, $C_2O_4^{2-}$.)

Solution

The products of the neutralization reaction will be water and calcium oxalate:

$H_2C_2O_4(s) + Ca(OH)_2(s) \rightarrow 2H_2O(\ell) + CaC_2O_4(s)$

Because nothing is dissolved, there are no substances to separate into ions, so the net ionic equation is the equation of the three solids and one liquid.

Exercise 6.6.2

What is the net ionic equation between $HNO_3(aq)$ and $Ti(OH)_4(s)$?

Answer

$4H^+(aq) + Ti(OH)_4(s) \rightarrow 4H_2O(\ell) + Ti^{4+}(aq)$

Key Takeaways

- The Arrhenius definition of an acid is a substance that increases the amount of H^+ in an aqueous solution.
- The Arrhenius definition of a base is a substance that increases the amount of OH^- in an aqueous solution.
- Neutralization is the reaction of an acid and a base, which forms water and a salt.
- Net ionic equations for neutralization reactions may include solid acids, solid bases, solid salts, and water.

Answers

Exercise 6.6.1

Yes; $FeCl_2$ and I_2

Exercise 6.6.2

$Mg_3(PO_4)_2$ and Al

Exercise 6.6.3

KNO_3 and AgBr

Exercise 6.6.4

No reaction; all possible products are soluble.

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6.7: Ionic Equations

Learning Objectives

- Write ionic equations for chemical reactions between ionic compounds.
- Write net ionic equations for chemical reactions between ionic compounds.

For single-replacement and double-replacement reactions, many of the reactions included ionic compounds—compounds between metals and nonmetals, or compounds that contained recognizable polyatomic ions. Now, we take a closer look at reactions that include ionic compounds.

One important aspect about ionic compounds that differs from molecular compounds has to do with dissolution in a liquid, such as water. When molecular compounds, such as sugar, dissolve in water, the individual molecules drift apart from each other. When ionic compounds dissolve, *the ions physically separate from each other*. We can use a chemical equation to represent this process—for example, with NaCl:

$$\mathrm{NaCl\,(s)} \xrightarrow{\mathrm{H_2O}} \mathrm{Na^+\,(aq) + Cl^-\,(aq)}$$

When NaCl dissolves in water, the ions separate and go their own way in solution; the ions are now written with their respective charges, and the (aq) phase label emphasizes that they are dissolved (Figure 6.7.1). This process is called **dissociation;** we say that the ions *dissociate*.

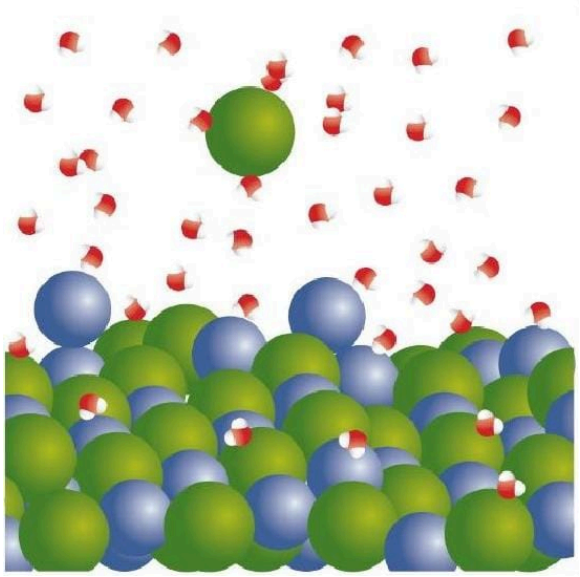


Figure 6.7.1 Ionic Solutions. When an ionic compound dissociates in water, water molecules surround each ion and separate it from the rest of the solid. Each ion goes its own way in solution.

All ionic compounds that dissolve behave this way. This behavior was first suggested by the Swedish chemist Svante August Arrhenius [1859–1927] as part of his PhD dissertation in 1884. Interestingly, his PhD examination team had a hard time believing that ionic compounds would behave like this, so they gave Arrhenius a barely passing grade. Later, this work was cited when Arrhenius was awarded the Nobel Prize in Chemistry. Keep in mind that when the ions separate, *all* the ions separate. Thus, when $CaCl_2$ dissolves, the one Ca^{2+} ion and the two Cl^- ions separate from one another:

$$CaCl_2(s) \overset{H_2O}{\rightarrow} Ca^{2+}(aq) + Cl^-(aq) + Cl^-(aq)$$

$$CaCl_2(s) \overset{H_2O}{\rightarrow} Ca^{2+}(aq) + 2Cl^-(aq)$$

That is, the two chloride ions go off on their own. They do not remain as Cl_2 (that would be elemental chlorine; these are chloride ions), and they do not stick together to make Cl_2^- or Cl_2^{2-}. They become dissociated ions in their own right. Polyatomic ions also retain their overall identity when they are dissolved.

Example 6.7.1

Write the chemical equation that represents the dissociation of each ionic compound.

1. KBr
2. Na_2SO_4

Solution

1. $KBr(s) \rightarrow K^+(aq) + Br^-(aq)$
2. Not only do the two sodium ions go their own way, but the sulfate ion stays together as the sulfate ion. The dissolving equation is $Na_2SO_4(s) \rightarrow 2Na^+(aq) + SO_4^{2-}(aq)$

Exercise 6.7.1

Write the chemical equation that represents the dissociation of $(NH_4)_2S$.

Answer available at the end of this section.

When chemicals in solution react, the proper way of writing the chemical formulas of the dissolved ionic compounds is in terms of the dissociated ions, not the complete ionic formula. A **complete ionic equation** is a chemical equation in which the dissolved ionic compounds are written as separated ions. Solubility rules are very useful in determining which ionic compounds are dissolved and which are not. For example, when NaCl(aq) reacts with $AgNO_3(aq)$ in a double-replacement reaction to precipitate AgCl(s) and form $NaNO_3(aq)$, the complete ionic equation includes NaCl, $AgNO_3$, and $NaNO_3$ written as separate ions:

$$\mathrm{Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO_3^-(aq) \longrightarrow AgCl(s) + Na^+(aq) + NO_3^-(aq)}$$

This is more representative of what is occurring in the solution.

Example 6.7.1

Write the complete ionic equation for each chemical reaction.

1. $KBr(aq) + AgC_2H_3O_2(aq) \rightarrow KC_2H_3O_2(aq) + AgBr(s)$
2. $MgSO_4(aq) + Ba(NO_3)_2(aq) \rightarrow Mg(NO_3)_2(aq) + BaSO_4(s)$

Solution

For any ionic compound that is aqueous, we will write the compound as separated ions.

1. The complete ionic equation is $K^+(aq) + Br^-(aq) + Ag^+(aq) + C_2H_3O_2^-(aq) \rightarrow K^+(aq) + C_2H_3O_2^-(aq) + AgBr(s)$
2. The complete ionic equation is $Mg^{2+}(aq) + SO_4^{2-}(aq) + Ba^{2+}(aq) + 2NO_3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO_3^-(aq) + BaSO_4(s)$

Exercise 6.7.1

Write the complete ionic equation for

$$CaCl_2\,(aq) + Pb(NO_3)_2\,(aq) \longrightarrow Ca(NO_3)_2\,(aq) + PbCl_2\,(s)$$

Answer available at the end of this section.

You may notice that in a complete ionic equation, some ions do not change their chemical form; they stay exactly the same on the reactant and product sides of the equation. For example, in

$Na^+(aq) + Cl^-(aq) + Ag^+(aq) + NO3^-(aq) \rightarrow AgCl(s) + Na^+(aq) + NO3^-(aq)$

the $Ag^+(aq)$ and $Cl^-(aq)$ ions become AgCl(s), but the $Na^+(aq)$ ions and the $NO3^-(aq)$ ions stay as $Na^+(aq)$ ions and $NO3^-(aq)$ ions. These two ions are examples of spectator ions—ions that do nothing in the overall course of a chemical reaction. They are present, but they do not participate in the overall chemistry. It is common to cancel spectator ions (something also done with algebraic quantities) on the opposite sides of a chemical equation:

$$\cancel{Na^+(aq)} + Cl^-(aq) + Ag^+(aq) + \cancel{NO_3^-(aq)} \rightarrow AgCl(s) + \cancel{Na^+(aq)} + \cancel{NO_3^-(aq)}$$

What remains when the spectator ions are removed is called the **net ionic equation**, which represents the actual chemical change occurring between the ionic compounds:

$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

It is important to reiterate that the spectator ions are still present in solution, but they do not experience any net chemical change, so they are not written in a net ionic equation.

Example 6.7.1

Write the net ionic equation for each chemical reaction.

1. $K^+(aq) + Br^-(aq) + Ag^+(aq) + C2H3O2^-(aq) \rightarrow K^+(aq) + C2H3O2^-(aq) + AgBr(s)$
2. $Mg^{2+}(aq) + SO4^{2-}(aq) + Ba^{2+}(aq) + 2NO3^-(aq) \rightarrow Mg^{2+}(aq) + 2NO3^-(aq) + BaSO4(s)$

Solution

1. In the first equation, the $K^+(aq)$ and $C2H3O2^-(aq)$ ions are spectator ions, so they are canceled:

$$\cancel{K^+(aq)} + Br^-(aq) + Ag^+(aq) + \cancel{C_2H_3O_2^-(aq)} \rightarrow K^+(aq) + \cancel{C_2H_3O_2^-(aq)} + AgBr(s)$$

The net ionic equation is

$Br^-(aq) + Ag^+(aq) \rightarrow AgBr(s)$

2. In the second equation, the $Mg^{2+}(aq)$ and $NO3^-(aq)$ ions are spectator ions, so they are canceled:

$$\cancel{Mg^{2+}(aq)} + SO_4^{2-}(aq) + Ba^{2+}(aq) + \cancel{2NO_3^-(aq)} \rightarrow Mg^{2+}(aq) + \cancel{2NO_3^-(aq)} + BaSo_4(s)$$

The net ionic equation is

$SO4^{2-}(aq) + Ba^{2+}(aq) \rightarrow BaSO4(s)$

Exercise 6.7.1

Write the net ionic equation for

$CaCl_2(aq) + Pb(NO_3)_2(aq) \rightarrow Ca(NO_3)_2(aq) + PbCl_2(s)$

Answer available at the end of this section.

Chemistry is Everywhere: Soluble and Insoluble Ionic Compounds

The concept of solubility versus insolubility in ionic compounds is a matter of degree. Some ionic compounds are very soluble, some are only moderately soluble, and some are soluble so little that they are considered insoluble. For most ionic compounds, there is also a limit to the amount of compound that can be dissolved in a sample of water. For example, you can dissolve a maximum of 36.0 g of NaCl in 100 g of water at room temperature, but you can dissolve only 0.00019 g of AgCl in 100 g of water. We consider NaCl soluble but AgCl insoluble.

One place where solubility is important is in the tank-type water heater found in many homes in the United States. Domestic water frequently contains small amounts of dissolved ionic compounds, including calcium carbonate ($CaCO_3$). However, $CaCO_3$ has the relatively unusual property of being less soluble in hot water than in cold water. So as the water heater operates by heating water, $CaCO_3$ can precipitate if there is enough of it in the water. This precipitate, called *limescale*, can also contain magnesium compounds, hydrogen carbonate compounds, and phosphate compounds. The problem is that too much limescale can impede the function of a water heater, requiring more energy to heat water to a specific temperature or even blocking water pipes into or out of the water heater, causing dysfunction.



Figure 6.7.2 Tank water heater © Thinkstock. Most homes in the United States have a tank-type water heater like this one.

Another place where solubility versus insolubility is an issue is the Grand Canyon. We usually think of rock as insoluble. But it is actually ever so slightly soluble. This means that over a period of about two billion years, the Colorado River carved rock from the surface by slowly dissolving it, eventually generating a spectacular series of gorges and canyons. And all because of solubility!



Figure 6.7.3: The Grand Canyon was formed by water running through rock for billions of years, very slowly dissolving it. Note the Colorado River is still present in the lower part of the photo. (Sonaal Bangera via unsplash)

Key Takeaways

- Ionic compounds that dissolve separate into individual ions.
- Complete ionic equations show dissolved ionic solids as separated ions.
- Net ionic equations show only the ions and other substances that change in a chemical reaction.

Answers

Exercise 6.7.1

$(NH_4)_2S(s) \rightarrow 2NH_4^+(aq) + S^{2-}(aq)$

Exercise 6.7.1

$Ca^{2+}(aq) + 2Cl^-(aq) + Pb^{2+}(aq) + 2NO_3^-(aq) \rightarrow Ca^{2+}(aq) + 2NO_3^-(aq) + PbCl_2(s)$

Exercise 6.7.1

$Pb^{2+}(aq) + 2Cl^-(aq) \rightarrow PbCl_2(s)$

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6.8: Introduction to Stoichiometry

Learning Objectives

- Define *stoichiometry*.
- Relate quantities in a balanced chemical reaction on a molecular basis.

Consider a classic recipe for pound cake: 1 pound of eggs, 1 pound of butter, 1 pound of flour, and 1 pound of sugar. (That's why it's called "pound cake.") If you have 4 pounds of butter, how many pounds of sugar, flour, and eggs do you need? You would need 4 pounds each of sugar, flour, and eggs.

Now suppose you have 1.00 g H_2. If the chemical reaction follows the balanced chemical equation

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

then what mass of oxygen do you need to make water?

Curiously, this chemical reaction question is very similar to the pound cake question. Both of them involve relating a quantity of one substance to a quantity of another substance or substances. The relating of one chemical substance to another using a balanced chemical reaction is called **stoichiometry**. Using stoichiometry is a fundamental skill in chemistry; it greatly broadens your ability to predict what will occur and, more importantly, how much is produced.

Let us consider a more complicated example. A recipe for pancakes calls for 2 cups (c) of pancake mix, 1 egg, and 1/2 c of milk. We can write this in the form of a chemical equation:

2 c mix + 1 egg + 1/2 c milk → 1 batch of pancakes

If you have 9 c of pancake mix, how many eggs and how much milk do you need? It might take a little bit of work, but eventually you will find you need 4½ eggs and 2¼ c milk.

How can we formalize this? We can make a conversion factor using our original recipe and use that conversion factor to convert from a quantity of one substance to a quantity of another substance. This is similar to the way we constructed a conversion factor between feet and yards in Chapter 2. Because one recipe's worth of pancakes requires 2 c of pancake mix, 1 egg, and 1/2 c of milk, we actually have the following mathematical relationships that relate these quantities:

2 c pancake mix ⇔ 1 egg ⇔ 1/2 c milk

where ⇔ is the mathematical symbol for "is equivalent to." This does not mean that 2 c of pancake mix equal 1 egg. However, *as far as this recipe is concerned*, these are the equivalent quantities needed for a single recipe of pancakes. So, any possible quantities of two or more ingredients must have the same numerical ratio as the ratios in the equivalence.

We can deal with these equivalences in the same way we deal with equalities in unit conversions: we can make conversion factors that essentially equal 1. For example, to determine how many eggs we need for 9 c of pancake mix, we construct the conversion factor:

$$\frac{1\,egg}{2\,c\,pancake\,mix}$$

This conversion factor is, in a strange way, equivalent to 1 because the recipe relates the two quantities. Starting with our initial quantity and multiplying by our conversion factor,

$$9\,\cancel{c\,pancake\,mix} \times \frac{1\,egg}{\cancel{2\,c\,pancake\,mix}} = 4.5\,eggs$$

Note how the units *cups pancake mix* canceled, leaving us with units of *eggs*. This is the formal, mathematical way of getting our amounts to mix with 9 c of pancake mix. We can use a similar conversion factor for the amount of milk:

$$9\,c\,\cancel{pancake\ mix} \times \frac{\frac{1}{2}c\,milk}{2\,c\,\cancel{pancake\ mix}} = 2.25\,c\,milk$$

Again, units cancel, and new units are introduced.

A balanced chemical equation is nothing more than *a recipe for a chemical reaction*. The difference is that a balanced chemical equation is written in terms of atoms and molecules, not cups, pounds, and eggs.

For example, consider the following chemical equation:

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

We can interpret this as, literally, "two hydrogen molecules react with one oxygen molecule to make two water molecules." That interpretation leads us directly to some equivalencies, just as our pancake recipe did:

2H2 molecules ⇔ 1O2 molecule ⇔ 2H2O molecules

These equivalences allow us to construct conversion factors:

$$\frac{2\,molecules\,H_2}{1\,molecule\,O_2}\ \frac{2\,molecules\,H_2}{2\,molecules\,H_2O}\ \frac{1\,molecule\,H_2}{2\,molecules\,H_2O}$$

and so forth. These conversions can be used to relate quantities of one substance to quantities of another. For example, suppose we need to know how many molecules of oxygen are needed to react with 16 molecules of H2. As we did with converting units, we start with our given quantity and use the appropriate conversion factor:

$$\cancel{16\,molecules\,H_2} \times \frac{1\,molecules\,O_2}{\cancel{2\,molecules\,H_2}} = 8\,molecules\,O_2$$

Note how the unit *molecules H2* cancels algebraically, just as any unit does in a conversion like this. The conversion factor came directly from the coefficients in the balanced chemical equation. This is another reason why a properly balanced chemical equation is important.

Example 6.8.1

How many molecules of SO3 are needed to react with 144 molecules of Fe2O3 given this balanced chemical equation?

$$Fe_2O_3 + 3\,SO_3 \longrightarrow Fe_2(SO_4)_3$$

Solution

We use the balanced chemical equation to construct a conversion factor between Fe2O3 and SO3. The number of molecules of Fe2O3 goes on the bottom of our conversion factor so it cancels with our given amount, and the molecules of SO3 go on the top. Thus, the appropriate conversion factor is

$$\frac{3\,molecules\,SO_3}{1\,molecule\,Fe_2O_3}$$

Starting with our given amount and applying the conversion factor, the result is

$$144\ \cancel{molecules}\ Fe_2O_3 \times \frac{3\,molecules\,SO_3}{1\ \cancel{molecule}\ Fe_2O_3} = 432\,molecules\,SO_3$$

We need 432 molecules of SO3 to react with 144 molecules of Fe2O3.

Exercise 6.8.1

How many molecules of H_2 are needed to react with 29 molecules of N_2 to make ammonia if the balanced chemical equation is:

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3?$$

Answer available at the end of this section.

Chemical equations also allow us to make conversions regarding the number of atoms in a chemical reaction, because a chemical formula lists the number of atoms of each element in a compound. The formula H_2O indicates that there are two hydrogen atoms and one oxygen atom in each molecule, and these relationships can be used to make conversion factors:

$$\frac{2\,atoms\,H}{1\,molecule\,H_2O} \quad \frac{1\,molecule\,H_2O}{1\,atom\,O}$$

Conversion factors like this can also be used in stoichiometry calculations.

Example 6.8.2

How many molecules of NH_3 can you make if you have 228 atoms of H_2?

Solution

From the formula, we know that one molecule of NH_3 has three H atoms. Use that fact as a conversion factor:

$$228\ \cancel{atoms\,H} \times \frac{1\,molecule\,NH_3}{3\ \cancel{atoms\,H}} = 76\,NH_3$$

Exercise 6.8.2

How many molecules of $Fe_2(SO_4)_3$ can you make from 777 atoms of S?

Answer available at the end of this section.

Summary

Quantities of substances can be related to each other using balanced chemical equations.

Learning Objectives

- Balance a chemical equation in terms of moles.
- Use the balanced equation to construct conversion factors in terms of moles.
- Calculate moles of one substance from moles of another substance using a balanced chemical equation.

Consider this balanced chemical equation:

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

We interpret this as "two molecules of hydrogen react with one molecule of oxygen to make two molecules of water." The chemical equation is balanced as long as the coefficients are in the ratio 2:1:2. For instance, this chemical equation is also balanced:

$$100\,H_2 + 50\,O_2 \longrightarrow 100\,H_2O$$

This equation is not conventional—because convention says that we use the lowest ratio of coefficients—but it is balanced. So is this chemical equation:

$$5{,}000\,H_2 + 2{,}500\,O_2 \longrightarrow 5{,}000\,H_2O$$

Again, this is not conventional, but it is still balanced. Suppose we use a much larger number:

$$12.044 \times 10^{23} H_2 + 6.022 \times 10^{23} O_2 \rightarrow 12.044 \times 10^{23} H_2O$$

These coefficients are also in the ratio of 2:1:2. But these numbers are related to the number of things in a mole: the first and last numbers are two times Avogadro's number, while the second number is Avogadro's number. That means that the first and last numbers represent 2 mol, while the middle number is just 1 mol. Well, why not just use the number of moles in balancing the chemical equation?

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

is the same balanced chemical equation we started with! What this means is that chemical equations are not just balanced in terms of molecules; *they are also balanced in terms of moles*. We can just as easily read this chemical equation as "two moles of hydrogen react with one mole of oxygen to make two moles of water." All balanced chemical reactions are balanced in terms of moles.

Example 6.8.1

Interpret this balanced chemical equation in terms of moles.

$$P_4 + 5\,O_2 \longrightarrow P_4O_{10}$$

Solution

The coefficients represent the number of moles that react, not just molecules. We would speak of this equation as "one mole of molecular phosphorus reacts with five moles of elemental oxygen to make one mole of tetraphosphorus decoxide."

Exercise 6.8.1

Interpret this balanced chemical equation in terms of moles.

$$N_2 + 3\,H_2 \longrightarrow 2\,NH_3$$

Answer available at the end of this section.

In Section 4.1, we stated that a chemical equation is simply a recipe for a chemical reaction. As such, chemical equations also give us equivalents—equivalents between the reactants and the products. However, now we understand that *these equivalents are expressed in terms of moles*. Consider the chemical equation

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O$$

This chemical reaction gives us the following equivalents:

$$2\text{ mol }H_2 \Leftrightarrow 1\text{ mol }O_2 \Leftrightarrow 2\text{ mol }H_2O$$

Any two of these quantities can be used to construct a conversion factor that lets us relate the number of moles of one substance to an equivalent number of moles of another substance. If, for example, we want to know how many moles of oxygen will react with 17.6 mol of hydrogen, we construct a conversion factor between 2 mol of H_2 and 1 mol of O_2 and use it to convert from moles of one substance to moles of another:

$$17.6\,\cancel{mol\,H_2} \times \frac{1\,mol\,O_2}{2\,\cancel{mol\,H_2}} = 8.80\,mol\,O_2$$

Note how the mol H_2 unit cancels, and mol O_2 is the new unit introduced. This is an example of a **mole-mole calculation,** when you start with moles of one substance and convert to moles of another substance by using the balanced chemical equation. The example may seem simple because the numbers are small, but numbers won't always be so simple!

Example 6.8.2

For the balanced chemical equation

$$2\,C_4H_{10}\,(g) + 13\,O_2 \longrightarrow 8\,CO_2\,(g) + 10\,H_2O(\ell)$$

if 154 mol of O_2 are reacted, how many moles of CO_2 are produced?

Solution

We are relating an amount of oxygen to an amount of carbon dioxide, so we need the equivalence between these two substances. According to the balanced chemical equation, the equivalence is

$$13 \text{ mol } O_2 \Leftrightarrow 8 \text{ mol } CO_2$$

We can use this equivalence to construct the proper conversion factor. We start with what we are given and apply the conversion factor:

$$154\ \cancel{mol\ O_2} \times \frac{8\ mol\ CO_2}{13\ \cancel{mol\ O_2}} = 94.8\ mol\ CO_2$$

The mol O_2 unit is in the denominator of the conversion factor so it cancels. Both the 8 and the 13 are exact numbers, so they do not contribute to the number of significant figures in the final answer.

Exercise 6.8.2

Using the above equation, how many moles of H_2O are produced when 154 mol of O_2 react?

Answer available at the end of this section.

It is important to reiterate that balanced chemical equations are balanced in terms of *moles*. Not grams, kilograms, or liters—but moles. Any stoichiometry problem will likely need to work through the mole unit at some point, especially if you are working with a balanced chemical reaction.

Summary

Balanced chemical reactions are balanced in terms of moles. A balanced chemical reaction gives equivalents in moles that allow stoichiometry calculations to be performed.

Learning Objectives

- From a given number of moles of a substance, calculate the mass of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the moles of another substance involved using the balanced chemical equation.
- From a given mass of a substance, calculate the mass of another substance involved using the balanced chemical equation.

Mole-mole calculations are not the only type of calculations that can be performed using balanced chemical equations. Recall

that the molar mass can be determined from a chemical formula and used as a conversion factor. We can add that conversion factor as another step in a calculation to make a **mole-mass calculation**, where we start with a given number of moles of a substance and calculate the mass of another substance involved in the chemical equation, or vice versa.

For example, suppose we have the balanced chemical equation

$$2Al + 3Cl_2 \rightarrow 2AlCl_3$$

Suppose we know we have 123.2 g of Cl_2. How can we determine how many moles of $AlCl_3$ we will get when the reaction is complete? First and foremost, *chemical equations are not balanced in terms of grams; they are balanced in terms of moles*. So to use the balanced chemical equation to relate an amount of Cl_2 to an amount of $AlCl_3$, we need to convert the given amount of Cl_2 into moles. We know how to do this by simply using the molar mass of Cl_2 as a conversion factor. The molar mass of Cl_2 (which we get from the atomic mass of Cl from the periodic table) is 70.90 g/mol. We must invert this fraction so that the units cancel properly:

$$123.2 \ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} = 1.738\ mol\,Cl_2$$

Now that we have the quantity in moles, we can use the balanced chemical equation to construct a conversion factor that relates the number of moles of Cl_2 to the number of moles of $AlCl_3$. The numbers in the conversion factor come from the coefficients in the balanced chemical equation:

$$\frac{2\,mol\,AlCl_3}{3\,mol\,Cl_2}$$

Using this conversion factor with the molar quantity we calculated above, we get

$$1.738\ \cancel{mol\,Cl_2} \times \frac{2\,mol\,AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\,AlCl_3$$

So, we will get 1.159 mol of $AlCl_3$ if we react 123.2 g of Cl_2.

In this last example, we did the calculation in two steps. However, it is mathematically equivalent to perform the two calculations sequentially on one line:

$$123.2\ \cancel{g\,Cl_2} \times \frac{1\,mol\,Cl_2}{70.90\ \cancel{g\,Cl_2}} \times \frac{2\,mol\,AlCl_3}{3\ \cancel{mol\,Cl_2}} = 1.159\,mol\,AlCl_3$$

The units still cancel appropriately, and we get the same numerical answer in the end. Sometimes the answer may be slightly different from doing it one step at a time because of rounding of the intermediate answers, but the final answers should be effectively the same.

Example 6.8.1

How many moles of HCl will be produced when 249 g of $AlCl_3$ are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Solution

We will do this in two steps: convert the mass of $AlCl_3$ to moles and then use the balanced chemical equation to find the number of moles of HCl formed. The molar mass of $AlCl_3$ is 133.33 g/mol, which we have to invert to get the appropriate conversion factor:

$$249\ \cancel{g\,AlCl_3} \times \frac{1\,mol\,AlCl_3}{133.33\ \cancel{g\,AlCl_3}} = 1.87\,mol\,AlCl_3$$

Now we can use this quantity to determine the number of moles of HCl that will form. From the balanced chemical equation, we construct a conversion factor between the number of moles of AlCl3 and the number of moles of HCl:

$$frac6\,mol\,HCl2\ \cancel{mol\,AlCl_3}$$

Applying this conversion factor to the quantity of AlCl3, we get

$$1.87\ \cancel{mol\,AlCl_3} \times \frac{6\,mol\,HCl}{2\ \cancel{mol\,AlCl_3}} = 5.61\,mol\,HCl$$

Alternatively, we could have done this in one line:

$$249\ \cancel{g\,AlCl_3} \times \frac{1\ \cancel{mol\,AlCl_3}}{133.33\ \cancel{g\,AlCl_3}} \times \frac{6\,mol\,HCl}{2\ \cancel{mol\,AlCl_3}} = 5.60\,mol\,HCl$$

The last digit in our final answer is slightly different because of rounding differences, but the answer is essentially the same.

Exercise 6.8.1

How many moles of Al2O3 will be produced when 23.9 g of H2O are reacted according to this chemical equation?

$$2AlCl_3 + 3H_2O(\ell) \rightarrow Al_2O_3 + 6HCl(g)$$

Answer available at the end of this section.

A variation of the mole-mass calculation is to start with an amount in moles and then determine an amount of another substance in grams. The steps are the same but are performed in reverse order.

Example 6.8.2

How many grams of NH3 will be produced when 33.9 mol of H2 are reacted according to this chemical equation?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

Solution

The conversions are the same, but they are applied in a different order. Start by using the balanced chemical equation to convert to moles of another substance and then use its molar mass to determine the mass of the final substance. In two steps, we have

$$33.9\ \cancel{mol\,H_2} \times \frac{2\,mol\,NH_3}{3\ \cancel{mol\,H_2}} = 22.6\,mol\,NH_3$$

Now, using the molar mass of NH3, which is 17.03 g/mol, we get

$$22.6\ \cancel{mol\,NH_3} \times \frac{17.03\,g\,NH_3}{1\ \cancel{mol\,NH_3}} = 385\,g\,NH_3$$

Exercise 6.8.2

How many grams of N2 are needed to produce 2.17 mol of NH3 when reacted according to this chemical equation?

$$\mathrm{N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g)}$$

Answer available at the end of this section.

It should be a trivial task now to extend the calculations to **mass-mass calculations**, in which we start with a mass of some substance and end with the mass of another substance in the chemical reaction. For this type of calculation, the molar masses of two different substances must be used—be sure to keep track of which is which. Again, however, it is important to emphasize that before the balanced chemical reaction is used, the mass quantity must first be converted to moles. Then the coefficients of the balanced chemical reaction can be used to convert to moles of another substance, which can then be converted to a mass.

For example, let us determine the number of grams of SO3 that can be produced by the reaction of 45.3 g of SO2 and O2:

$$2SO_2(g) + O_2(g) \rightarrow 2SO_3(g)$$

First, we convert the given amount, 45.3 g of SO2, to moles of SO2 using its molar mass (64.06 g/mol):

$$45.3\ \cancel{g\ SO_2} \times \frac{1\ mol\ SO_2}{64.06\ \cancel{g\ SO_2}} = 0.707\ mol\ SO_2$$

Second, we use the balanced chemical reaction to convert from moles of SO2 to moles of SO3:

$$0.707\ \cancel{mol\ SO_2} \times \frac{2\ mol\ SO_3}{2\ \cancel{mol\ SO_2}} = 0.707\ mol\ SO_3$$

Finally, we use the molar mass of SO3 (80.06 g/mol) to convert to the mass of SO3:

$$0.707\ \cancel{mol\ SO_3} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

We can also perform all three steps sequentially, writing them on one line as

$$45.3\ \cancel{g\ SO_2} \times \frac{1\ mol\ SO_2}{64.06\ \cancel{g\ SO_2}} \times \frac{2\ mol\ SO_3}{2\ \cancel{mol\ SO_2}} \times \frac{80.06\ g\ SO_3}{1\ \cancel{mol\ SO_3}} = 56.6\ g\ SO_3$$

We get the same answer. Note how the initial and all the intermediate units cancel, leaving grams of SO3, which is what we are looking for, as our final answer.

Example 6.8.3

What mass of Mg will be produced when 86.4 g of K are reacted?

$$\mathrm{MgCl_2\,(s) + 2\,K\,(s) \longrightarrow Mg\,(s) + 2\,KCl\,(s)}$$

Solution

We will simply follow the steps

mass K → mol K → mol Mg → mass Mg

In addition to the balanced chemical equation, we need the molar masses of K (39.09 g/mol) and Mg (24.31 g/mol). In one line,

$$86.4\ \cancel{g\ K} \times \frac{1\ mol\ K}{39.09\ \cancel{g\ K}} \times \frac{1\ \cancel{mol\ Mg}}{2\ \cancel{mol\ K}} \times \frac{24.31\ g\ Mg}{1\ \cancel{mol\ Mg}} = 26.87\ g\ Mg$$

Exercise 6.8.3

What mass of H_2 will be produced when 122 g of Zn are reacted?

$$\mathrm{Zn\,(s) + 2\,HCl\,(aq) \longrightarrow ZnCl_2\,(aq) + H_2\,(g)}$$

Answer available at the end of this section.

Summary

- Mole quantities of one substance can be related to mass quantities using a balanced chemical equation.
- Mass quantities of one substance can be related to mass quantities using a balanced chemical equation.
- In all cases, quantities of a substance must be converted to moles before the balanced chemical equation can be used to convert to moles of another substance.

Answers

Exercise 6.8.1
87 molecules

Exercise 6.8.2
259 molecules

Exercise 6.8.1
One mole of elemental nitrogen reacts with three moles of elemental hydrogen to produce two moles of ammonia.

Exercise 6.8.2
118 mol

Exercise 6.8.1
0.442 mol

Exercise 6.8.2
30.4 g (Note: here we go from a product to a reactant, showing that mole-mass problems can begin and end with any substance in the chemical equation.)

Exercise 6.8.3
3.77 g

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6.9: Solution Stoichiometry

Learning Objective

- Apply concentration units as conversion factors.

Concentration can be a conversion factor between the amount of solute and the amount of solution or solvent (depending on the definition of the concentration unit). As such, concentrations can be useful in a variety of stoichiometry problems. In many cases, it is best to use the original definition of the concentration unit; it is that definition that provides the conversion factor.

A simple example of using a concentration unit as a conversion factor is one in which we use the definition of the concentration unit and rearrange; we can do the calculation again as a unit conversion, rather than as a definition. For example, suppose we ask how many moles of solute are present in 0.108 L of a 0.887 M NaCl solution. Because 0.887 M means 0.887 mol/L, we can use this second expression for the concentration as a conversion factor:

$$0.108 \cancel{L\,NaCl} \times \frac{0.887\,mol\,NaCl}{\cancel{L\,NaCl}} = 0.0958\,mol\,NaCl$$

(There is an understood 1 in the denominator of the conversion factor.) If we used the definition approach, we get the same answer, but now we are using conversion factor skills. Like any other conversion factor that relates two different types of units, the reciprocal of the concentration can be also used as a conversion factor.

Example 6.9.1

Using concentration as a conversion factor, how many liters of 2.35 M $CuSO_4$ are needed to obtain 4.88 mol of $CuSO_4$?

Solution

This is a one-step conversion, but the concentration must be written as the reciprocal for the units to work out:

$$4.88 \cancel{mol\,CuSO_4} \times \frac{1\,L}{2.35\,\cancel{mol}} = 2.08\,L\,of\,solution$$

Exercise 6.9.1

Using concentration as a conversion factor, how many liters of 0.0444 M CH_2O are needed to obtain 0.0773 mol of CH_2O?

Answer: available at the end of this section.

Of course, once quantities in moles are available, another conversion can give the mass of the substance, using molar mass as a conversion factor.

Example 6.9.2

What mass of solute is present in 0.765 L of 1.93 M NaOH?

Solution

This is a two-step conversion, first using concentration as a conversion factor to determine the number of moles and then the molar mass of NaOH (40.0 g/mol) to convert to mass:

$$0.765 \cancel{L} \times \frac{1.93 \cancel{mol\ NaOH}}{\cancel{L\ solution}} \times \frac{40.0 g\ NaOH}{1 \cancel{mol\ NaOH}} = 59.1\ g\ NaOH$$

Exercise 6.9.2

What mass of solute is present in 1.08 L of 0.0578 M H_2SO_4?

Answer available at the end of this section.

More complex stoichiometry problems using balanced chemical reactions can also use concentrations as conversion factors. For example, suppose the following equation represents a chemical reaction:

$$2\,AgNO_3\,(aq) + CaCl_2\,(aq) \longrightarrow 2\,AgCl\,(s) + Ca(NO_3)_2\,(aq)$$

If we wanted to know what volume of 0.555 M $CaCl_2$ would react with 1.25 mol of $AgNO_3$, we first use the balanced chemical equation to determine the number of moles of $CaCl_2$ that would react and then use concentration to convert to liters of solution:

$$1.25\ \cancel{mol\ AgNO_3} \times \frac{1\ \cancel{mol\ CaCl_2}}{2\ \cancel{mol\ AgNO_3}} \times \frac{1 L\ solution}{0.555\ \cancel{mol\ CaCl_2}} = 1.13\ L\ CaCl_2$$

This can be extended by starting with the mass of one reactant, instead of moles of a reactant.

Example 6.9.3

What volume of 0.0995 M $Al(NO_3)_3$ will react with 3.66 g of Ag according to the following chemical equation?

$$3\,Ag\,(s) + Al(NO_3)_3\,(aq) \longrightarrow 3\,AgNO_3 + Al\,(s)$$

Solution

Here, we first must convert the mass of Ag to moles before using the balanced chemical equation and then the definition of molarity as a conversion factor:

$$3.66\ \cancel{g\ Ag} \times \frac{1\ \cancel{mol\ Ag}}{107.97\ \cancel{g\ Ag}} \times \frac{1\ \cancel{mol\ Al(NO_3)_3}}{3\ \cancel{mol\ Ag}} \times \frac{1 L\ solution}{0.0995\ \cancel{mol\ Al(NO_3)_3}} = 0.114\ L$$

The strikeouts show how the units cancel.

Exercise 6.9.3

What volume of 0.512 M NaOH will react with 17.9 g of $H_2C_2O_4(s)$ according to the following chemical equation?

$$H_2C_2O_4\,(s) + 2\,NaOH\,(aq) \longrightarrow Na_2C_2O_4\,(aq) + 2\,H_2O(\ell)$$

Answer: available at the end of this section.

We can extend our skills even further by recognizing that we can relate quantities of one solution to quantities of another solution. Knowing the volume and concentration of a solution containing one reactant, we can determine how much of another solution of another reactant will be needed using the balanced chemical equation.

Example 6.9.4

A student takes a precisely measured sample, called an *aliquot*, of 10.00 mL of a solution of FeCl3. The student carefully adds 0.1074 M Na2C2O4 until all the Fe^{3+}(aq) has precipitated as Fe2(C2O4)3(s). Using a precisely measured tube called a burette, the student finds that 9.04 mL of the Na2C2O4 solution was added to completely precipitate the Fe^{3+}(aq). What was the concentration of the FeCl3 in the original solution? (A precisely measured experiment like this, which is meant to determine the amount of a substance in a sample, is called a *titration*.) The balanced chemical equation is as follows:

$$2\,\mathrm{FeCl_3\,(aq)} + 3\,\mathrm{Na_2C_2O_4\,(aq)} \longrightarrow \mathrm{Fe_2(C_2O_4)_3\,(s)} + 6\,\mathrm{NaCl\,(aq)}$$

Solution

First we need to determine the number of moles of Na2C2O4 that reacted. We will convert the volume to liters and then use the concentration of the solution as a conversion factor:

$$9.04\,\cancel{mL} \times \frac{1\,\cancel{L}}{1000\,\cancel{mL}} \times \frac{0.1074 mol\,Na_2C_2O_4}{\cancel{L}} = 0.000971\,mol\,Na_2C_2O_4$$

Now we will use the balanced chemical equation to determine the number of moles of Fe^{3+}(aq) that were present in the initial aliquot:

$$0.000971\,\cancel{mol\,Na_2C_2O_4} \times \frac{2 mol\,FeCl_3}{3\,\cancel{mol\,Na_2C_2O_4}} = 0.000647 mol\,FeCl_3$$

Then we determine the concentration of FeCl3 in the original solution. Converting 10.00 mL into liters (0.01000 L), we use the definition of molarity directly:

$$M = \frac{mol}{L} = \frac{0.000647 mol\,FeCl_3}{0.01000 L} = 0.0647 M\,FeCl_3$$

Exercise 6.9.4

A student titrates 25.00 mL of H3PO4 with 0.0987 M KOH. She uses 54.06 mL to complete the chemical reaction. What is the concentration of H3PO4?

$$\mathrm{H_3PO_4\,(aq)} + 3\,\mathrm{KOH\,(aq)} \longrightarrow \mathrm{K_3PO_4\,(aq)} + 3\,\mathrm{H_2O}$$

Answer: available at the end of this section.

We have used molarity exclusively as the concentration of interest, but that will not always be the case. The next example shows a different concentration unit being used.

Example 6.9.5

H2O2 is used to determine the amount of Mn according to this balanced chemical equation:

$$2\,\mathrm{MnO_4^-\,(aq)} + 5\,\mathrm{H_2O_2\,(aq)} + 6\,\mathrm{H^+\,(aq)} \longrightarrow 2\,\mathrm{Mn_2^+\,(aq)} + 5\,\mathrm{O_2\,(g)} + 8\,\mathrm{H_2O(\ell)}$$

What mass of 3.00% m/m H2O2 solution is needed to react with 0.355 mol of MnO_4^-(aq)?

Solution

Because we are given an initial amount in moles, all we need to do is use the balanced chemical equation to determine the number of moles of H_2O_2 and then convert to find the mass of H_2O_2. Knowing that the H_2O_2 solution is 3.00% by mass, we can determine the mass of solution needed:

$$0.355 \cancel{mol\ MnO_4^-} \times \frac{5\ \cancel{mol\ H_2O_2}}{2\ \cancel{mol MnO_4^-}} \times \frac{34.02\ \cancel{g\ H_2O_2}}{\cancel{mol\ H_2O_2}} \times \frac{100g\ solution}{3\ \cancel{g\ H_2O_2}} = 1006g\ sol$$

The first conversion factor comes from the balanced chemical equation, the second conversion factor is the molar mass of H_2O_2, and the third conversion factor comes from the definition of percentage concentration by mass.

Exercise 6.9.5

Use the balanced chemical reaction for MnO_4^- and H_2O_2 to determine what mass of O_2 is produced if 258 g of 3.00% m/m H_2O_2 is reacted with MnO_4^-.

Answer available at the end of this section.

Summary

Know how to apply concentration units as conversion factors.

Answers

Exercise 6.9.1
1.74 L

Exercise 6.9.2
6.12 g

Exercise 6.9.3
0.777 L

Exercise 6.9.4
0.0711 M



Figure 6.9.1 Titration © Thinkstock. When a student performs a titration, a measured amount of one solution is added to another reactant.

Exercise 6.9.5
7.28 g

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6.10: Titrations

Learning Objectives

- Describe a titration experiment.
- Explain what an indicator does.
- Perform a titration calculation correctly.

The reaction of an acid with a base to make a salt and water is a common reaction in the laboratory, partly because many compounds can act as acids or bases. Another reason that acid-base reactions are so prevalent is because they are often used to determine quantitative amounts of one or the other. Performing chemical reactions quantitatively to determine the exact amount of a reagent is called a **titration**. A titration can be performed with almost any chemical reaction for which the balanced chemical equation is known. Here, we will consider titrations that involve acid-base reactions.

In a titration, one reagent has a known concentration or amount, while the other reagent has an unknown concentration or amount. Typically, the known reagent (the **titrant**) is added to the unknown quantity and is dissolved in solution. The unknown amount of substance (the **analyte**) may or may not be dissolved in solution (but usually is). The titrant is added to the analyte using a precisely calibrated volumetric delivery tube called a burette (also spelled buret; see Figure 6.10.1). The burette has markings to determine how much volume of solution has been added to the analyte. When the reaction is complete, it is said to be at the **equivalence point**; the number of moles of titrant can be calculated from the concentration and the volume, and the balanced chemical equation can be used to determine the number of moles (and then concentration or mass) of the unknown reactant.



Figure 6.10.1 Equipment for Titrations. A burette is a type of liquid dispensing system that can accurately indicate the volume of liquid dispensed.

For example, suppose 25.66 mL (or 0.02566 L) of 0.1078 M HCl was used to titrate an unknown sample of NaOH. What mass of NaOH was in the sample? We can calculate the number of moles of HCl reacted:

$$\text{\# mol HCl} = (0.02566\ \text{L})(0.1078\ \text{M}) = 0.002766\ \text{mol HCl}$$

We also have the balanced chemical reaction between HCl and NaOH:

$$\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$$

So we can construct a conversion factor to convert to number of moles of NaOH reacted:

$$0.002766\,\cancel{mol\ HCl} \times \frac{1\,mol\,NaOH}{1\,\cancel{mol\ HCl}} = 0.002766\,mol\,NaOH$$

Then we convert this amount to mass, using the molar mass of NaOH (40.00 g/mol):

$$0.002766\,\cancel{mol\ HCl} \times \frac{40.00\,g\,NaOH}{1\,\cancel{mol\ HCl}} = 0.1106\,g\,NaOH$$

This type of calculation is performed as part of a titration.

Example 6.10.1

What mass of Ca(OH)2 is present in a sample if it is titrated to its equivalence point with 44.02 mL of 0.0885 M HNO3? The balanced chemical equation is as follows:

$$2HNO_3 + Ca(OH)_2 \rightarrow Ca(NO_3)_2 + 2H_2O$$

Solution

In liters, the volume is 0.04402 L. We calculate the number of moles of titrant:

$$\#\text{ moles } HNO_3 = (0.04402\text{ L})(0.0885\text{ M}) = 0.00390\text{ mol } HNO_3$$

Using the balanced chemical equation, we can determine the number of moles of Ca(OH)2 present in the analyte:

$$0.00390\ \cancel{mol\ HNO_3} \times \frac{1\,mol\,Ca(OH)_2}{2\ \cancel{mol\ HNO_3}} = 0.00195\,mol\,Ca(OH)_2$$

Then we convert this to a mass using the molar mass of Ca(OH)2:

$$0.00195\ \cancel{mol\,Ca(OH)_2} \times \frac{74.1\,g\,Ca(OH)_2}{\cancel{mol\,Ca(OH)_2}} = 0.144\,g\,Ca(OH)_2$$

Exercise 6.10.1

What mass of H2C2O4 is present in a sample if it is titrated to its equivalence point with 18.09 mL of 0.2235 M NaOH? The balanced chemical reaction is as follows:

$$H_2C_2O_4 + 2\,NaOH \longrightarrow Na_2C_2O_4 + 2\,H_2O$$

Answer available at the end of this section.

How does one know if a reaction is at its equivalence point? Usually, the person performing the titration adds a small amount of an **indicator**, a substance that changes color depending on the acidity or basicity of the solution. Because different indicators change colors at different levels of acidity, choosing the correct one is important in performing an accurate titration.

Summary

A titration is the quantitative reaction of an acid and a base. Indicators are used to show that all of the analyte has reacted with the titrant.

Answers

Exercise 6.10.1

0.182 g

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6.11: Stoichiometry with Reaction Enthalpies

Learning Objective

- Perform stoichiometry calculations using energy changes from thermochemical equations.

Previously, we related quantities of one substance to another in a chemical equation by performing calculations that used the balanced chemical equation; the balanced chemical equation provided equivalents that we used to construct conversion factors. For example, in the balanced chemical equation

$$2\,H_2\,(g) + O_2\,(g) \longrightarrow 2\,H_2O(\ell)$$

we recognized the equivalents

$$2\,\text{mol}\ H_2 \Leftrightarrow\ 1\,\text{mol}\ O_2 \Leftrightarrow\ 2\,\text{mol}\ H_2O$$

where $\Leftrightarrow$ is the mathematical symbol for "is equivalent to." In our thermochemical equation, however, we have another quantity-energy change:

\[\ce{2H2(g) + O2(g) → 2H2O(ℓ)} \quad \Δ*H* = –570\, kJ \nonumber \]

This new quantity allows us to add another equivalent to our list:

$$2\,\text{mol}\ H_2 \Leftrightarrow\ 1\,\text{mol}\ O_2 \Leftrightarrow\ 2\,\text{mol}\ H_2O \Leftrightarrow\ -570\ \ k\,\text{J}$$

That is, we can now add an energy amount to the equivalents—the enthalpy change of a balanced chemical reaction. This equivalence can also be used to construct conversion factors so that we can relate enthalpy change to amounts of substances reacted or produced.

Note that these equivalents address a concern. When an amount of energy is listed for a balanced chemical reaction, what amount(s) of reactants or products does it refer to? The answer is that it relates to the number of moles of the substance, as indicated by its coefficient in the balanced chemical reaction. Thus, 2 mol of H_2 are related to –570 kJ, while 1 mol of O_2 is related to –570 kJ. This is why the unit on the energy change is kJ, not kJ/mol.

For example, consider the thermochemical equation

$$H_2\,(g) + Cl_2\,(g) \longrightarrow 2\,HCl\,(g) \quad \Delta H = -184.6\,kJ$$

The equivalencies for this thermochemical equation are

$$1\,\text{mol}\ H_2 \Leftrightarrow\ 1\,\text{mol}\ Cl_2 \Leftrightarrow\ 2\,\text{mol}\ HCl \Leftrightarrow\ -184.6\,\text{kJ}$$

Suppose we are asked how much energy is given off when 8.22 mol of H_2 react. We would construct a conversion factor between the number of moles of H_2 and the energy given off, –184.6 kJ:

$$8.22\ \cancel{mol\ H_2} \times \frac{-184.6\,kJ}{1\ \cancel{mol\ H_2}} = -1520\,kJ$$

The negative sign means that this much energy is given off.

Example 6.11.1

Given the thermochemical equation

$$N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g) \quad \Delta H = -91.8\,kJ$$

how much energy is given off when 222.4 g of N_2 reacts?

Solution

The balanced thermochemical equation relates the energy change to moles, not grams, so we first convert the amount of N_2 to moles and then use the thermochemical equation to determine the energy change:

$$222.4\ \cancel{g\ N_2} \times \frac{1\ \cancel{mol\ N_2}}{28.00\ \cancel{g\ N_2}} \times \frac{-91.8\ kJ}{1\ \cancel{mol\ N_2}} = -729\ kJ$$

Exercise 6.11.1

Given the thermochemical equation

$$\mathrm{N_2\,(g) + 3\,H_2\,(g) \longrightarrow 2\,NH_3\,(g)} \quad \Delta H = -91.8\ kJ$$

how much heat is given off when 1.00 g of H_2 reacts?

Answer available at the end of this section.

Like any stoichiometric quantity, we can start with energy and determine an amount, rather than the other way around.

Example 6.11.1

Given the thermochemical equation

$$\mathrm{N_2\,(g) + O_2\,(g) \longrightarrow 2\,NO\,(g)} \quad \Delta H = 180.6\ kJ$$

if 558 kJ of energy are supplied, what mass of NO can be made?

Solution

This time, we start with an amount of energy

$$558\ \cancel{kJ} \times \frac{2\ \cancel{mol\ NO}}{180.6\ \cancel{kJ}} \times \frac{30.0\ g\ NO}{1\ \cancel{mol\ NO}} = 185\ g\ NO$$

Exercise 6.11.1

How many grams of N_2 will react if 100.0 kJ of energy are supplied?

$$\mathrm{N_2\,(g) + O_2\,(g) \longrightarrow 2\,NO\,(g)} \quad \Delta H = 180.6\ kJ$$

Answer available at the end of this section.

Chemistry is Everywhere: Welding with Chemical Reactions

One very energetic reaction is called the **thermite reaction**. Its classic reactants are aluminum metal and iron(III) oxide; the reaction produces iron metal and aluminum oxide:

$$\mathrm{2\,Al\,(s) + Fe_2O_3\,(s) \longrightarrow Al_2O_3\,(s) + 2\,Fe\,(s)} \quad \Delta H = -850.2\ kJ$$

When properly done, the reaction gives off so much energy that the iron product comes off as a *liquid*. (Iron normally melts at 1,536°C.) If carefully directed, the liquid iron can fill spaces between two or more metal parts and, after it quickly cools, can weld the metal parts together.

Thermite reactions are used for this purpose even today. For civilian purposes, they are used to reweld broken locomotive axles that cannot be easily removed for repair. They are also used to weld railroad tracks together. Thermite reactions can also be used to separate thin pieces of metal if, for whatever reason, a torch doesn't work.



Figure 6.11.1 Thermite mixture. A small clay pot contains a thermite mixture. It is reacting at high temperature in the photo and will eventually produce molten metal to join the railroad tracks below it. (Public Domain; Skatebiker via Wikipedia)

Thermite reactions are also used for military purposes. Thermite mixtures are frequently used with additional components as incendiary devices—devices that start fires. Thermite reactions are also useful in disabling enemy weapons—a piece of artillery does not work when it has a hole melted into its barrel as a result of a thermite reaction!

Key Takeaway

- The energy change of a chemical reaction can be used in stoichiometry calculations.

Answers

Exercise 6.11.1

–15.1 kJ

Exercise 6.11.1

15.5 g

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6.12: Theoretical and Percent Yields

Learning Objective

- Define and determine theoretical yields, actual yields, and percent yields.

In all the previous calculations we have performed involving balanced chemical equations, we made two assumptions:

1. The reaction goes exactly as written.
2. The reaction proceeds completely.

In reality, such things as side reactions occur that make some chemical reactions rather messy. For example, in the actual combustion of some carbon-containing compounds, such as methane, some CO is produced as well as CO_2. However, we will continue to ignore side reactions, unless otherwise noted. The second assumption, that the reaction proceeds completely, is more troublesome. Many chemical reactions do not proceed to completion as written, for a variety of reasons (some of which we will consider in Chapter 13). When we calculate an amount of product assuming that all the reactant reacts, we calculate the **theoretical yield**, an amount that is theoretically produced as calculated using the balanced chemical reaction.

In many cases, however, this is not what really happens. In many cases, less—sometimes, much less—of a product is made during the course of a chemical reaction. The amount that is actually produced in a reaction is called the **actual yield**. By definition, the actual yield is less than or equal to the theoretical yield. If it is not, then an error has been made.

Both theoretical yields and actual yields are expressed in units of moles or grams. It is also common to see something called a percent yield. The **percent yield** is a comparison between the actual yield and the theoretical yield and is defined as

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\% \tag{6.12.1}$$

It does not matter whether the actual and theoretical yields are expressed in moles or grams, as long as they are expressed in the same units. However, the percent yield always has units of percent. Proper percent yields are between 0% and 100%. Again, if percent yield is greater than 100%, an error has been made.

Example 6.12.1

A worker reacts 30.5 g of Zn with nitric acid and evaporates the remaining water to obtain 65.2 g of $Zn(NO_3)_2$. What are the theoretical yield, the actual yield, and the percent yield?

$$\mathrm{Zn\,(s) + 2\,HNO_3\,(aq) \longrightarrow Zn(NO_3)_2\,(aq) + H_2\,(g)}$$

Solution

A mass-mass calculation can be performed to determine the theoretical yield. We need the molar masses of Zn (65.39 g/mol) and $Zn(NO_3)_2$ (189.41 g/mol). In three steps, the mass-mass calculation is:

$$30.5\ \cancel{g\ Zn} \times \frac{1\ \cancel{mol\ Zn}}{65.39\ \cancel{g\ Zn}} \times \frac{1\ \cancel{mol\ Zn(NO_3)_2}}{1\ \cancel{mol\ Zn}} \times \frac{189.41\ g\ Zn(NO_3)_2}{1\ \cancel{mol\ Zn(NO_3)_2}} = 88.3\ g\ Zn(NO_3)_2$$

Thus, the theoretical yield is 88.3 g of $Zn(NO_3)_2$. The actual yield is the amount that was actually made, which was 65.2 g of $Zn(NO_3)_2$. To calculate the percent yield, we take the actual yield and divide it by the theoretical yield and multiply by 100 (Equation 6.12.1):

$$\frac{65.2\ g\ Zn(NO_3)_2}{88.3\ g\ Zn(NO_3)_2} \times 100\% = 73.8\%$$

The worker achieved almost three-fourths of the possible yield.

Exercise 6.12.1

A synthesis produced 2.05 g of NH_3 from 16.5 g of N_2. What is the theoretical yield and the percent yield?

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

*Technically, this is a reversible reaction (with double arrows), but for this exercise consider it irreversible (single arrow).

Answer

theoretical yield = 20.1 g; percent yield = 10.2%

Chemistry is Everywhere: Actual Yields in Drug Synthesis and Purification

Many drugs are the product of several steps of chemical synthesis. Each step typically occurs with less than 100% yield, so the overall percent yield might be very small. The general rule is that the overall percent yield is the product of the percent yields of the individual synthesis steps. For a drug synthesis that has many steps, the overall percent yield can be very tiny, which is one factor in the huge cost of some drugs. For example, if a 10-step synthesis has a percent yield of 90% for each step, the overall yield for the entire synthesis is only 35%. Many scientists work every day trying to improve percent yields of the steps in the synthesis to decrease costs, improve profits, and minimize waste.

Even purifications of complex molecules into drug-quality purity are subject to percent yields. Consider the purification of impure albuterol. Albuterol ($C_{13}H_{21}NO_2$; accompanying figure) is an inhaled drug used to treat asthma, bronchitis, and other obstructive pulmonary diseases. It is synthesized from norepinephrine, a naturally occurring hormone and neurotransmitter. Its initial synthesis makes very impure albuterol that is purified in five chemical steps. The details of the steps do not concern us; only the percent yields do:

A child using an albuterol inhaler, a container of albuterol medication, and a molecular model of albuterol are shown in three combined images.

impure albuterol → intermediate A	percent yield = 70%
intermediate A → intermediate B	percent yield = 100%
intermediate B → intermediate C	percent yield = 40%
intermediate C → intermediate D	percent yield = 72%
intermediate D → purified albuterol	percent yield = 35%
overall percent yield = 70% × 100% × 40% × 72% × 35% = 7.5%	

That is, only about *one-fourteenth* of the original material was turned into the purified drug. This demonstrates one reason why some drugs are so expensive—a lot of material is lost in making a high-purity pharmaceutical.



Figure 6.12.1 A child using an albuterol inhaler, the container of albuterol medication, and a molecular model of the albuterol molecule. Source: Photo on far left © Thinkstock. Photo in center courtesy of Intropin, http://commons.wikimedia.org/wiki/Fi...te_%281%29.JPG.

Summary

Theoretical yield is the calculated yield using the balanced chemical reaction. Actual yield is what is actually obtained in a chemical reaction. Percent yield is a comparison of the actual yield with the theoretical yield.

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6.13: Limiting and Excess Reactants

Learning Objectives

- Identify a limiting reagent from a set of reactants.
- Calculate how much product will be produced from the limiting reagent.
- Calculate how much reactant(s) remains when the reaction is complete.

In addition to the assumption that reactions proceed all the way to completion, one additional assumption we have made about chemical reactions is that all the reactants are present in the proper quantities to react to products; this is not always the case. In Figure 6.13.2 we are taking hydrogen atoms and oxygen atoms (left) to make water molecules (right). However, there are not enough oxygen atoms to use up all the hydrogen atoms. We run out of oxygen atoms and cannot make any more water molecules, so the process stops when we run out of oxygen atoms.

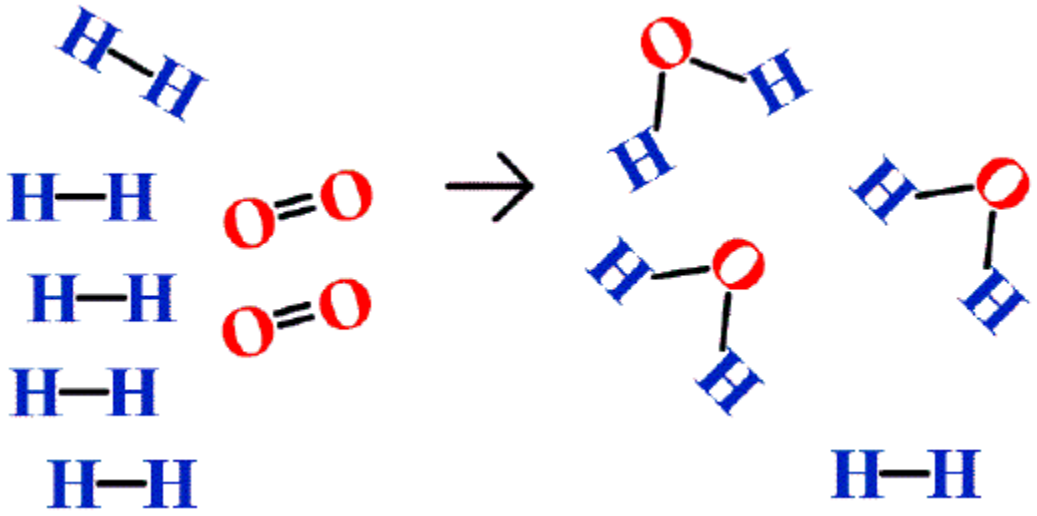


Figure 6.13.2: Making Water. In this scenario for making water molecules, we run out of O atoms before we use up all the H atoms. Similar situations exist for many chemical reactions when one reactant runs out before the other.

A similar situation exists for many chemical reactions: you usually run out of one reactant before all of the other reactant has reacted. The reactant you run out of is called the limiting reagent; the other reactant or reactants are considered to be *in excess*. A crucial skill in evaluating the conditions of a chemical process is to determine which reactant is the limiting reagent and which is in excess.

The key to recognizing which reactant is the limiting reagent is based on a mole-mass or mass-mass calculation: whichever reactant gives the *lesser* amount of product is the limiting reagent. What we need to do is determine an amount of one product (either moles or mass), assuming all of each reactant reacts. Whichever reactant gives the least amount of that particular product is the limiting reagent. It does not matter which product we use, as long as we use the same one each time. It does not matter whether we determine the number of moles or grams of that product; however, we will see shortly that knowing the final mass of product can be useful.

For example, consider this reaction:

$$4As(s) + 3O_2(g) \rightarrow 2As_2O_3(s)$$

Suppose we start a reaction with 50.0 g of As and 50.0 g of O2. Which one is the limiting reagent? We need to perform two mole-mass calculations, each assuming that each reactant reacts completely. Then we compare the amount of the product produced by each and determine which is less.

The calculations are as follows:

$$50.0\ \cancel{g\,As} \times \frac{1\ \cancel{mol\,As}}{74.92\ \cancel{g\,As}} \times \frac{2\,mol\,As_2O_3}{4\ \cancel{mol\,As}} = 0.334\,mol\,As_2O_3$$

$$50.0\ \cancel{g\,O_2} \times \frac{1\ \cancel{mol\,O_2}}{32.00\ \cancel{g\,O_2}} \times \frac{2\,mol\,As_2O_3}{3\ \cancel{mol\,O_2}} = 1.04\,mol\,As_2O_3$$

Comparing these two answers, it is clear that 0.334 mol of As_2O_3 is less than 1.04 mol of As_2O_3, so arsenic is the limiting reagent. If this reaction is performed under these initial conditions, the arsenic will run out before the oxygen runs out. We say that the oxygen is "in excess."

Identifying the limiting reagent, then, is straightforward. However, there are usually two associated questions: (1) what mass of product (or products) is then actually formed? and (2) what mass of what reactant is left over? The first question is straightforward to answer: simply perform a conversion from the number of moles of product formed to its mass, using its molar mass. For As_2O_3, the molar mass is 197.84 g/mol; knowing that we will form 0.334 mol of As_2O_3 under the given conditions, we will get

$$0.334\ \cancel{mol\,As_2O_3} \times \frac{197.84\,g\,As_2}{1\ \cancel{mol\,As_2O_3}} = 66.1\,g\,As_2O_3$$

The second question is somewhat more convoluted to answer. First, we must do a mass-mass calculation relating the limiting reagent (here, As) to the other reagent (O_2). Once we determine the mass of O_2 that reacted, we subtract that from the original amount to determine the amount left over. According to the mass-mass calculation,

$$50.0\ \cancel{g\,As} \times \frac{1\ \cancel{mol\,As}}{74.92\ \cancel{g\,As}} \times \frac{3\ \cancel{mol\,O_2}}{4\ \cancel{mol\,As}} \times \frac{32.00\,g\,O_2}{1\ \cancel{mol\,O_2}} = 16.0\,g\,O_2\ reacted$$

Because we reacted 16.0 g of our original O_2, we subtract that from the original amount, 50.0 g, to get the mass of O_2 remaining:

50.0 g O_2 – 16.0 g O_2 reacted = 34.0 g O_2 left over

You must remember to perform this final subtraction to determine the amount remaining; a common error is to report the 16.0 g as the amount remaining.

Example 6.13.1

A 5.00 g quantity of Rb is combined with 3.44 g of $MgCl_2$ according to this chemical reaction:

$$2Rb(s) + MgCl_2(s) \rightarrow Mg(s) + 2RbCl(s)$$

What mass of Mg is formed, and what mass of what reactant is left over?

Solution

Because the question asks what mass of magnesium is formed, we can perform two mass-mass calculations and determine which amount is less.

$$5.00\ \cancel{g\,Rb} \times \frac{1\ \cancel{mol\,Rb}}{85.47\ \cancel{g\,Rb}} \times \frac{1\ \cancel{mol\,Mg}}{2\ \cancel{mol\,Rb}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 0.711\,g\,Mg$$

$$3.44\ \cancel{g\,MgCl_2} \times \frac{1\ \cancel{mol\,MgCl_2}}{95.21\ \cancel{g\,MgCl_2}} \times \frac{1\ \cancel{mol\,Mg}}{1\ \cancel{mol\,MgCl_2}} \times \frac{24.31\,g\,Mg}{1\ \cancel{mol\,Mg}} = 0.878\,g\,Mg$$

The 0.711 g of Mg is the lesser quantity, so the associated reactant—5.00 g of Rb—is the limiting reagent. To determine how much of the other reactant is left, we have to do one more mass-mass calculation to determine what mass of MgCl2 reacted with the 5.00 g of Rb, and then subtract the amount reacted from the original amount.

$$5.00\ \cancel{g\ Rb} \times \frac{1\ \cancel{mol\ Rb}}{85.47\ \cancel{g\ Rb}} \times \frac{1\ \cancel{mol\ MgCl_2}}{2\ \cancel{mol\ Rb}} \times \frac{95.21\ g\ Mg}{1\ \cancel{mol\ MgCl_2}} = 2.78\ g\ MgCl_2\ \ reacted$$

Because we started with 3.44 g of MgCl2, we have

3.44 g MgCl2 – 2.78 g MgCl2 reacted = 0.66 g MgCl2 left

Exercise 6.13.1

Given the initial amounts listed, what is the limiting reagent, and what is the mass of the leftover reagent?

$$\underbrace{22.7\ g}_{MgO(s)} + \underbrace{17.9\ g}_{H_2S} \rightarrow MgS(s) + H_2O(l)$$

Answer available at the end of this section.

Summary

The limiting reagent is the reactant that produces the least amount of product. Mass-mass calculations can determine how much product is produced and how much of the other reactants remain.

Answers

Exercise 6.13.1

H2S is the limiting reagent; 1.5 g of MgO are left over.

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7: CHEMISTRY IN LIFE (UNDER CONSTRUCTION)

7.1: Reaction Kinetics
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7.3: Acids and Bases
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7.1: Reaction Kinetics

Learning Outcomes

- Define reaction rate.
- Define catalyst and explain its behavior in a chemical reaction.
- Describe how temperatures, concentration or pressure of reactant, and a catalyst affect the reaction rate.
- Explain the concept of activation energy.
- Label a diagram with reactants, products, enthalpy of forward and reverse reactions, activation energy of forward and reverse reactions, and activated complex.
- Calculate the enthalpy and activation energy with data on a reaction diagram.

Chemical kinetics is the study of the rates of chemical reactions. In this lesson, you will learn how to express the rate of a chemical reaction and about various factors that influence reaction rates.

Expressing Reaction Rate

Chemical reactions vary widely in the speeds with which they occur. Some reactions occur very quickly. It a lighted match is brought into contact with lighter fluid or another flammable liquid, it erupts into flame instantly and burns fast. Other reactions occur very slowly. A container of milk in the refrigerator will be good to drink for weeks before it begins to turn sour. Millions of years were required for dead plants under Earth's surface to accumulate and eventually turn into fossil fuels such as coal and oil.

Chemists need to be concerned with the rates at which chemical reactions occur. Rate is another word for speed. If a sprinter takes 11.0 seconds (s) to run a 100 meter (\left(m \right)\) dash, his rate or speed is given by the distance traveled divided by the time (see figure below).

$$\text{speed} = \frac{\text{distance}}{\text{time}} = \frac{100\ \text{m}}{11.0\ \text{s}} = 9.09\ \text{m/s} \tag{7.1.1}$$

The sprinter's average running rate for the race is $9.09\ \text{m/s}$. We say that it is his average rate because he did not run at that speed for the entire race. At the very beginning of the race, while coming from a standstill, his rate must be slower until he is able to get up to his top speed. His top speed must the be greater than $9.09\ \text{m/s}$ so that taken over the entire race, the average ends up at $9.09\ \text{m/s}$.

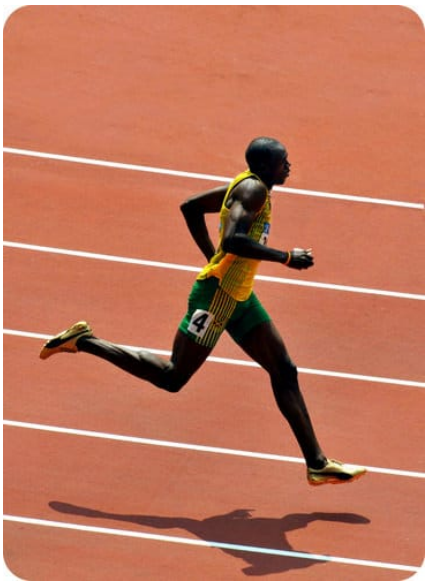


Figure 7.1.1: Usain Bolt set the world record for the 100 meter dash in 2009 with a time of 9.58 seconds. His average running rate over the course of this race was $10.4\ \mathrm{m/s}$, or $23.4\ \mathrm{mph}$.

Chemical reactions can't be measured in units of meters per second, as that would not make any sense. A **reaction rate** *is the change in concentration of a reactant or product with time*. Suppose that a simple reaction were to take place in which a 1.00 molar (M) aqueous solution of substance A was converted to substance B.

$$\mathrm{A}\,(aq) \rightarrow \mathrm{B}\,(aq) \tag{7.1.2}$$

Suppose that after 20.0 seconds, the concentration of A had dropped from $1.00\ \mathrm{M}$ to $0.72\ \mathrm{M}$ as it was being converted to substance B. We can express the rate of this reaction as the change in concentration of A divided by the time.

$$\text{rate} = -\frac{\Delta\,[\mathrm{A}]}{\Delta t} = -\frac{[\mathrm{A}]_{\text{final}} - [\mathrm{A}]_{\text{initial}}}{\Delta t} \tag{7.1.3}$$

A bracket around a symbol or formula means the concentration in molarity of that substance. The change in concentration of A is its final concentration minus its initial concentration. Because the concentration of A is decreasing over time, the negative sign is used. Thus, the rate for the reaction is positive, and the units are molarity per second or $\mathrm{M/s}$.

$$\text{rate} = -\frac{0.72\ \mathrm{M} - 1.00\ \mathrm{M}}{20.0\ \mathrm{s}} = -\frac{-0.28\ \mathrm{M}}{20.0\ \mathrm{s}} = 0.041\ \mathrm{M/s} \tag{7.1.4}$$

Over the first 20.0 seconds of this reaction, the molarity of A decreases by an average rate of $0.041\ \mathrm{M}$ every second. In summary, the rate of a chemical reaction is measured by the change in concentration over time for a reactant or product. The unit of measurement for a reaction rate is molarity per second $(\mathrm{M/s})$.

Collision Theory

The behavior of the reactant atoms, molecules, or ions is responsible for the rates of a given chemical reaction. **Collision theory** *is a set of principles based around the idea that reactant particles form products when they collide with one another, but only when those collisions have enough kinetic energy and the correct orientation to cause a reaction*. Particles that lack the necessary kinetic energy may collide, but the particles will simply bounce off one another unchanged. The figure below illustrates the difference. In the first collision, the particles bounce off one another, and no rearrangement of atoms has occurred. The second collision occurs with greater kinetic energy, and so the bond between the two red atoms breaks. One red atom bonds with the other molecule as one product, while the single red atom is the other produce. The first collision is called an ineffective collision, while the second collision is called an effective collision.

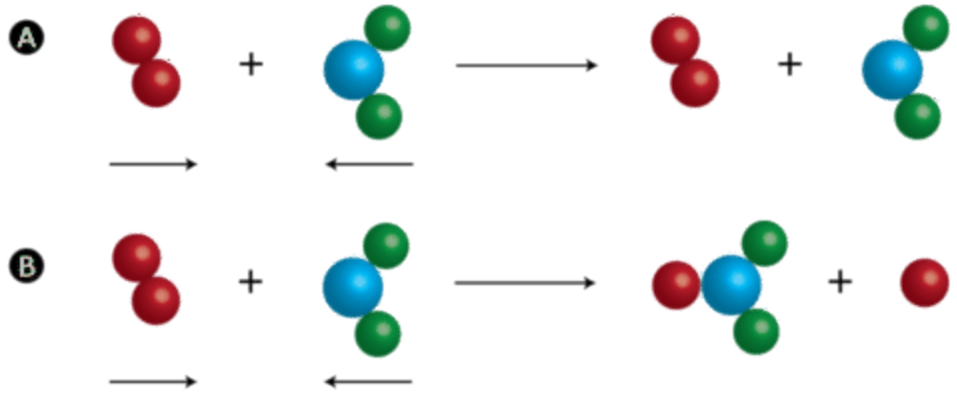


Figure 7.1.2: (A) An ineffective collision is one that does not result in product formation. (B) An effective collision is one in which chemical bonds are broken, and a product is formed.

Supplying reactant particles with energy causes the bonds between the atoms to vibrate with a greater frequency. This increase in vibrational energy makes a chemical bond more likely to break and a chemical reaction more likely to occur when those particles collide with other particles. Additionally, more energetic particles have more forceful collisions, which also increases the likelihood that a rearrangement of atoms will take place. The **activation energy** *for a reaction is the minimum energy that colliding particles must have in order to undergo a reaction*. Some reactions occur readily at room temperature because most of the reacting particles already have the requisite activation energy at that temperature. Other reactions only occur when heated because the particles do not have enough energy to react unless more is provided by an external source of heat.

Potential Energy Diagrams

Then energy changes that occur during a chemical reaction can be shown in a diagram called a potential energy diagram, sometimes called a reaction progress curve. A **potential energy diagram** *shows the change in the potential energy of a system as reactants are converted into products*. The figure below shows basic potential energy diagrams for an endothermic (left) and an exothermic (right) reaction. Recall that the enthalpy change (ΔH) is positive for an endothermic reaction and negative for an exothermic reaction. This can be seen in the potential energy diagrams. The total potential energy of the system increases for the endothermic reaction as the system absorbs energy from the surroundings. The total potential energy of the system decreases for the exothermic reaction as the system releases energy to the surroundings.

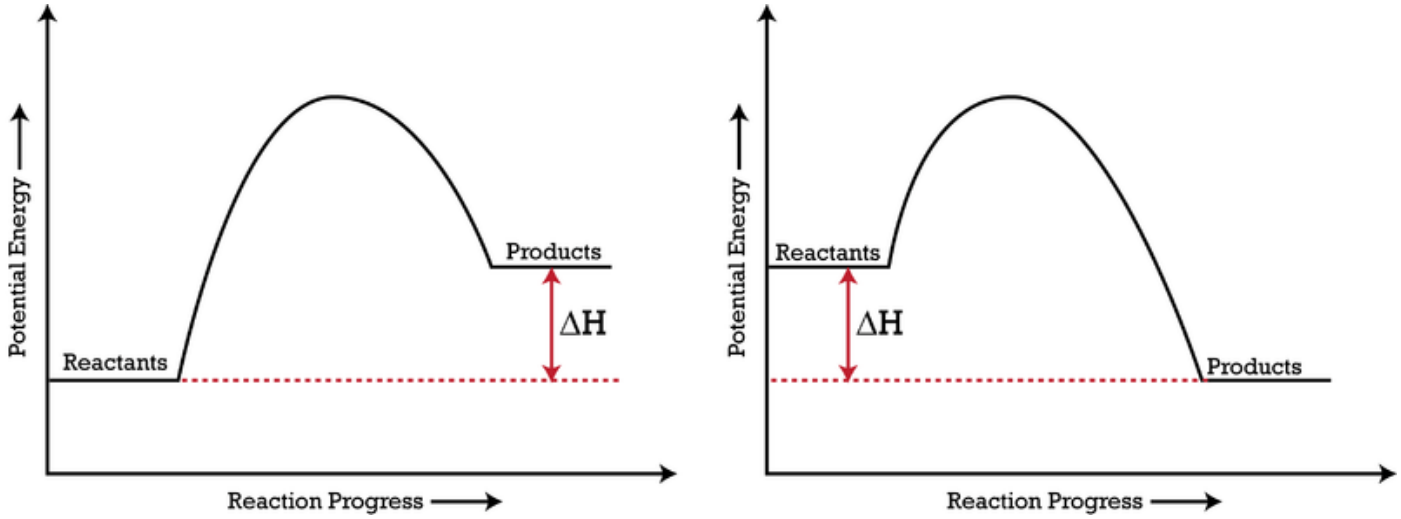


Figure 7.1.3: A potential energy diagram shows the total potential energy of a reacting system as the reaction proceeds. (Left) In an endothermic reaction, the energy of the products is greater than the energy of the reactants, and ΔH is positive. (Right) In an exothermic reaction, the energy of the products is lower than the energy of the reactants, and ΔH is negative.

The activation energy for a reaction is illustrated in the potential energy diagram by the height of the hill between the reactants and the products. For this reason, the activation energy of a reaction is sometimes referred to as the activation energy barrier. Reacting particles must have enough energy so that when they collide, they can overcome this barrier (see figure below).

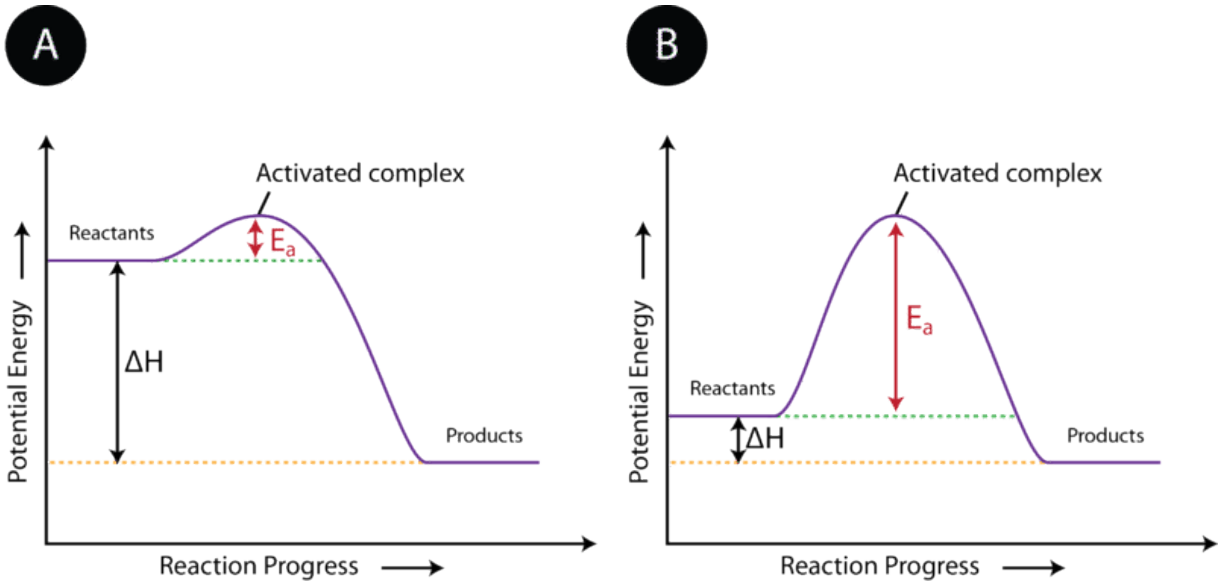


Figure 7.1.4: The activation energy (E_a) of a reaction is the barrier that must be overcome in order for the reactants to become products. (A) The activation energy is low, meaning that the reaction is likely to be fast. (B) The activation energy is high, meaning that the reaction is likely to be slow.

As discussed earlier, reactant particles sometimes collide with one another and yet remain unchanged by the collision. Other times, the collision leads to the formation of products. The state of the particles that is in between the reactants and products is called the activated complex. An **activated complex** *is an unstable arrangement of atoms that exists momentarily at the peak of the activation energy barrier*. Because of its high energy, the activated complex exists only for an extremely short period of time (about 10^{-13} s). The activated complex is equally likely to either reform the original reactants or go on to form the products. The figure below shows the formation of a possible activated complex between colliding hydrogen and oxygen molecules. Because of their unstable nature and brief existence, very little is known about the exact structures of most activated complexes.

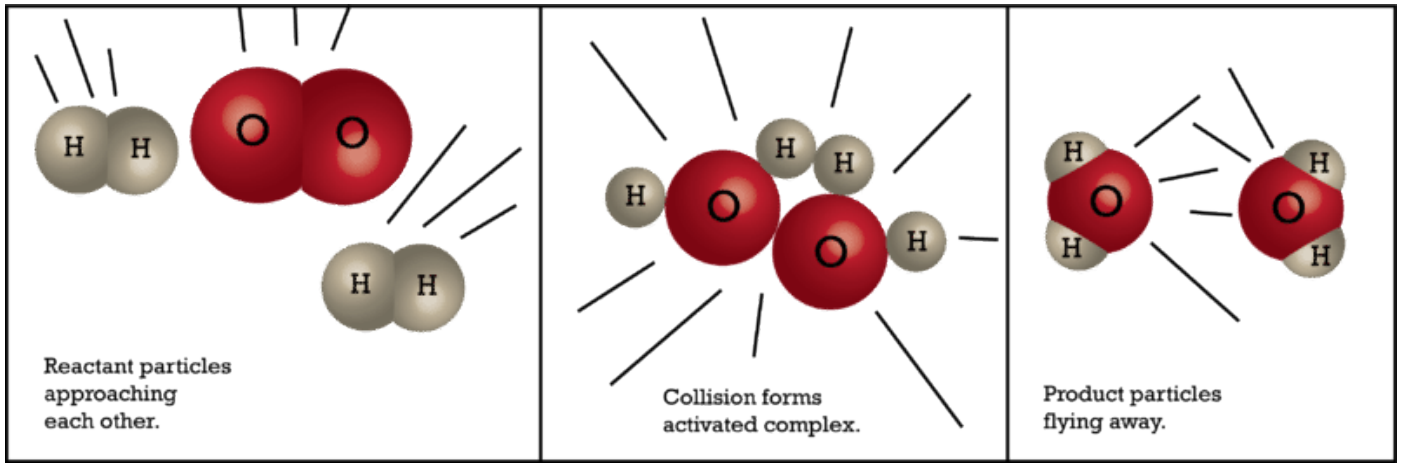


Figure 7.1.5: An activated complex is a short-lived state in which the colliding particles are at the peak of the potential energy curve.

Factors Affecting Reaction Rates

By their nature, some reactions occur very quickly, while others are very slow. However, certain changes in the reaction conditions can have an effect on the rate of a given chemical reaction. Collision theory can be utilized to explain these rate effects.

Concentration

Increasing the concentration of one or more of the reacting substances generally increases the reaction rate. When more particles are present in a given amount of space, a greater number of collisions will naturally occur between those particles. Since the rate

of a reaction is dependent on the frequency of collisions between the reactants, the rate increases as the concentration increases.

Pressure

When the pressure of a gas is increased, its particles are forced closer together, decreasing the amount of empty space between them. Therefore, an increase in the pressure of a gas is also an increase in the concentration of the gas. For gaseous reactions, an increase in pressure increases the rate of reaction for the same reasons as described above for an increase in concentration. Higher gas pressure leads to a greater frequency of collisions between reacting particles.

Surface Area

A large log placed in a fire will burn relatively slowly. If the same mass of wood were added to the fire in the form of small twigs, they would burn much more quickly. This is because the twigs provide a greater surface are than the log does. An increase in the surface area of a reactant increases the rate of a reaction. Surface area is larger when a given amount of a solid is present as smaller particles. A powdered reactant has a greater surface area than the same reactant as a solid chunk. In order to increase the surface area of a substance, it may be ground into smaller particles or dissolved into a liquid. In solution, the dissolved particles are separated from each other and will react more quickly with other reactants. The figure below shows the unfortunate result of high surface area in an unwanted combustion reaction. Small particles of grain dust are very susceptible to rapid reactions with oxygen, which can result in violent explosions and quick-burning fires.



Figure 7.1.6: This grain elevator in Kansas exploded in 1998. The tiny size of the reacting particles (grain dust) caused the reaction with oxygen in the air to be violently explosive.

Temperature

Raising the temperature of a chemical reaction results in a higher reaction rate. When the reactant particles are heated, they move faster and faster, resulting in a greater frequency of collisions. An even more important effect of the temperature increase is that the collisions occur with a greater force, which means the reactants are more likely to surmount the activation energy barrier and go on to form products. Increasing the temperature of a reaction increases not only the frequency of collisions, but also the percentage of those collisions that are effective, resulting in an increased reaction rate.

Paper is certainly a highly combustible material, but paper does not burn at room temperature because the activation energy for the reaction is too high. The vast majority of collisions between oxygen molecules and the paper are ineffective. However, when the paper is heated by the flame from a match, it reaches a point where the molecules now have enough energy to react. The reaction is very exothermic, so the heat released by the initial reaction will provide enough energy to allow the reaction to continue, even if the match is removed. The paper continues to burn rapidly until it is gone.

Catalysts

The rates of some chemical reactions can be increased dramatically by introducing certain other substances into the reaction mixture. Hydrogen peroxide is used as a disinfectant for scrapes and cuts, and it can be found in many medicine cabinets as a 3% aqueous solution. Hydrogen peroxide naturally decomposes to produce water and oxygen gas, but the reaction is very slow. A bottle of hydrogen peroxide will last for several years before it needs to be replaced. However, the addition of just a small amount of manganese (IV) oxide to hydrogen peroxide will cause it to decompose completely in just a matter of minutes. A **catalyst** *is*

a substance that increases the rate of a chemical reaction without being used up in the reaction. It accomplishes this task by providing an alternate reaction pathway that has a lower activation energy barrier. After the reaction occurs, a catalyst returns to its original state, so catalysts can be used over and over again. Because it is neither a reactant nor a product, a catalyst is shown in a chemical equation by being written above the yield arrow.

$$2\mathrm{H_2O_2}\,(aq) \overset{\mathrm{MnO_2}}{\rightarrow} 2\mathrm{H_2O}\,(l) + \mathrm{O_2}\,(g) \tag{7.1.5}$$

A catalyst works by changing the mechanism of the reaction, which can be though of as the specific set of smaller steps by which the reactants become products. The important point is that the use of a catalyst lowers the overall activation energy of the reaction (see figure below). With a lower activation energy barrier, a greater percentage of reactant molecules are able to have effective collisions, and the reaction rate increases.

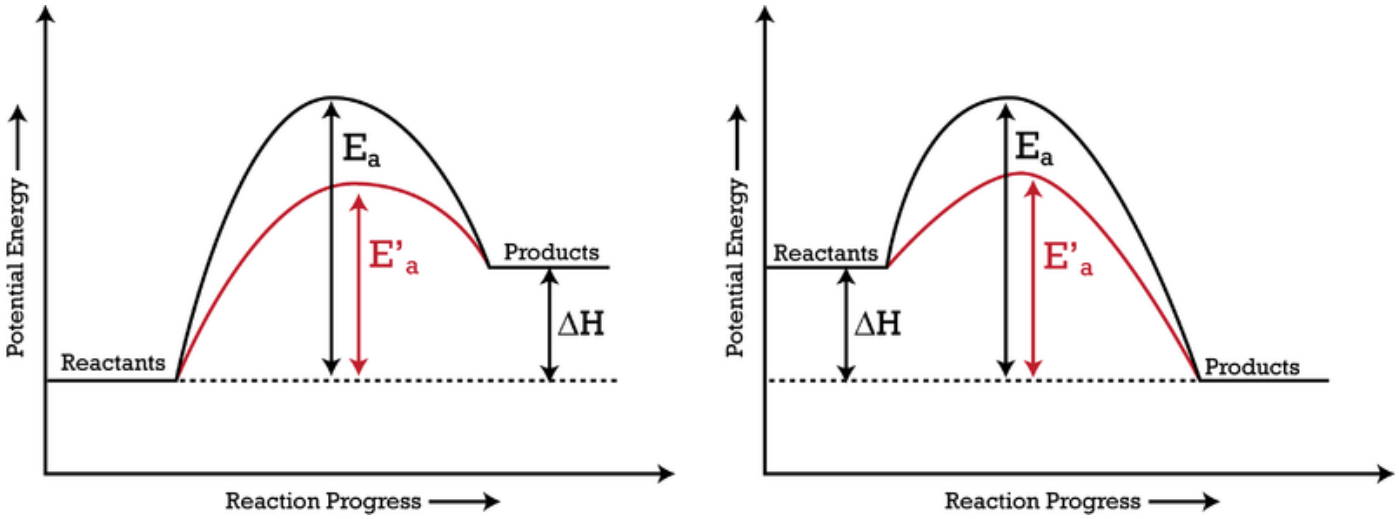


Figure 7.1.7: The addition of a catalyst to a reaction lowers the activation energy, increasing the rate of the reaction. The activation energy of the uncatalyzed reaction is shown by E_a, while the catalyzed reaction is shown by E_a'. The heat of reaction (ΔH) is unchanged by the presence of the catalyst.

Catalysts are extremely important parts of many chemical reactions. Enzymes in your body act as nature's catalysts, allowing important biochemical reactions to occur at reasonable rates. Chemical companies constantly search for new and better catalysts to make reactions go faster and thus make the company more profitable.

Contributors and Attributions

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7.2: Reaction Equilibrium

Learning Objectives

- Define *chemical equilibrium*.
- Recognize chemical equilibrium as a dynamic process.

Consider the following reaction occurring in a closed container (so that no material can go in or out):

$$H2 + I2 \rightarrow 2HI$$

This is simply the reaction between elemental hydrogen and elemental iodine to make hydrogen iodide. The way the equation is written, we are led to believe that the reaction goes to completion, that all the $H2$ and the $I2$ react to make HI.

However, this is not the case. The reverse chemical reaction is also taking place:

$$2HI \rightarrow H2 + I2$$

It acts to undo what the first reaction does. Eventually, the reverse reaction proceeds so quickly that it matches the speed of the forward reaction. When that happens, any continued overall reaction stops: the reaction has reached **chemical equilibrium** (sometimes just spoken as *equilibrium*; plural *equilibria*), the point at which the forward and reverse processes balance each other's progress.

Because two opposing processes are occurring at once, it is conventional to represent an equilibrium using a double arrow, like this:

$$H_2 + I_2 \rightleftharpoons 2HI$$

The double arrow implies that the reaction is going in both directions. Note that the reaction must still be balanced.

Example 7.2.1

Write the equilibrium equation that exists between calcium carbonate as a reactant and calcium oxide and carbon dioxide as products.

Solution

As this is an equilibrium situation, a double arrow is used. The equilibrium equation is written as follows:

$$CaCO_3 + \rightleftharpoons CaO + CO_2$$

Exercise 7.2.1

Write the equilibrium equation between elemental hydrogen and elemental oxygen as reactants and water as the product.

Answer available at the end of this section.

One thing to note about equilibrium is that the reactions do not stop; both the forward reaction and the reverse reaction continue to occur. They both occur at the same rate, so any overall change by one reaction is canceled by the reverse reaction. We say that chemical equilibrium is *dynamic*, rather than static. Also, because both reactions are occurring simultaneously, the equilibrium can be written backward. For example, representing an equilibrium as

$$H_2 + I_2 \rightleftharpoons 2HI$$

is the same thing as representing the same equilibrium as

$$2HI \rightleftharpoons H_2 + I_2$$

The reaction must be at equilibrium for this to be the case, however.

Key Takeaways

- Chemical reactions eventually reach equilibrium, a point at which forward and reverse reactions balance each other's progress.
- Chemical equilibria are dynamic: the chemical reactions are always occurring; they just cancel each other's progress.

Learning Objectives

- Explain the importance of the equilibrium constant.
- Construct an equilibrium constant expression for a chemical reaction.

In the mid 1860s, Norwegian scientists C. M. Guldberg and P. Waage noted a peculiar relationship between the amounts of reactants and products in an equilibrium. No matter how many reactants they started with, a certain ratio of reactants and products was achieved at equilibrium. Today, we call this observation the **law of mass action**. It relates the amounts of reactants and products at equilibrium for a chemical reaction. For a general chemical reaction occurring in solution,

$$aA + bB \rightleftharpoons cC + dD$$

the **equilibrium constant**, also known as K_{eq}, is defined by the following expression:

$$K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

where [A] is the molar concentration of species A at equilibrium, and so forth. The coefficients *a*, *b*, *c*, and *d* in the chemical equation become exponents in the expression for K_{eq}. The K_{eq} is a characteristic numerical value for a given reaction at a given temperature; that is, each chemical reaction has its own characteristic K_{eq}. The concentration of each reactant and product in a chemical reaction at equilibrium is *related*; the concentrations cannot be random values, but they depend on each other. The numerator of the expression for K_{eq} has the concentrations of every product (however many products there are), while the denominator of the expression for K_{eq} has the concentrations of every reactant, leading to the common *products over reactants* definition for the K_{eq}.

Let us consider a simple example. Suppose we have this equilibrium:

$$A \rightleftharpoons B$$

There is one reactant, one product, and the coefficients on each are just 1 (assumed, not written). The K_{eq} expression for this equilibrium is

$$K_{eq} = \frac{[B]}{[A]}$$

(Exponents of 1 on each concentration are understood.) Suppose the numerical value of K_{eq} for this chemical reaction is 2.0. If [B] = 4.0 M, then [A] must equal 2.0 M so that the value of the fraction equals 2.0:

$$K_{eq} = \frac{[B]}{[A]} = \frac{4.0}{2.0} = 2.0$$

By convention, the units are understood to be M and are omitted from the K_{eq} expression. Suppose [B] were 6.0 M. For the K_{eq} value to remain constant (it is, after all, called the equilibrium *constant*), then [A] would have to be 3.0 M at equilibrium:

$$K_{eq} = \frac{[B]}{[A]} = \frac{6.0}{3.0} = 2.0$$

If [A] were *not* equal to 3.0 M, the reaction would not be at equilibrium, and a net reaction would occur until that ratio was indeed 2.0. At that point, the reaction is at equilibrium, and any net change would cease. (Recall, however, that the forward and reverse reactions do not stop because chemical equilibrium is dynamic.)

The issue is the same with more complex expressions for the K_{eq}; only the mathematics become more complex. Generally speaking, given a value for the K_{eq} and all but one concentration at equilibrium, the missing concentration can be calculated.

Example 7.2.1

Given the following reaction:

$$\mathrm{H_2 + I_2 \rightleftharpoons 2\,HI} \tag{7.2.1}$$

If the equilibrium [HI] is 0.75 M and the equilibrium [H_2] is 0.20 M, what is the equilibrium [I_2] if the K_{eq} is 0.40?

Solution

We start by writing the K_{eq} expression. Using the *products over reactants* approach, the K_{eq} expression is as follows:

$$K_{eq} = \frac{[HI]^2}{[H_2][I_2]}$$

Note that [HI] is squared because of the coefficient 2 in the balanced chemical equation. Substituting for the equilibrium [H_2] and [HI] and for the given value of K_{eq}:

$$0.40 = \frac{(0.75)^2}{(0.20)[I_2]}$$

To solve for [I_2], we have to do some algebraic rearrangement: divide the 0.40 into both sides of the equation and multiply both sides of the equation by [I_2]. This brings [I_2] into the numerator of the left side and the 0.40 into the denominator of the right side:

$$[I_2] = \frac{(0.75)^2}{(0.20)(0.40)}$$

Solving,

[I_2] = 7.0 M

The concentration unit is assumed to be molarity. This value for [I_2] can be easily verified by substituting 0.75, 0.20, and 7.0 into the expression for K_{eq} and evaluating: you should get 0.40, the numerical value of K_{eq} (and you do).

Exercise 7.2.1

Given the following reaction:

$$\mathrm{H_2 + I_2 \rightleftharpoons 2\,HI} \tag{7.2.2}$$

If the equilibrium [HI] is 0.060 M and the equilibrium [I_2] is 0.90 M, what is the equilibrium [H_2] if the K_{eq} is 0.40?

Answer available at the end of this section.

In some types of equilibrium problems, square roots, cube roots, or even higher roots need to be analyzed to determine a final

answer. Make sure you know how to perform such operations on your calculator; if you do not know, ask your instructor for assistance.

Example 7.2.2

The following reaction is at equilibrium:

$$\mathrm{N_2} + 3\,\mathrm{H_2} \rightleftharpoons 2\,\mathrm{NH_3} \tag{7.2.3}$$

The K_{eq} at a particular temperature is 13.7. If the equilibrium [N_2] is 1.88 M and the equilibrium [NH_3] is 6.62 M, what is the equilibrium [H_2]?

Solution

We start by writing the K_{eq} expression from the balanced chemical equation:

$$K_{eq} = \frac{[NH_3]^2}{[N_2][H_2]^3}$$

Substituting for the known equilibrium concentrations and the K_{eq}, this becomes

$$13.7 = \frac{(6.62)^2}{(1.88)[H_2]^3}$$

Rearranging algebraically and then evaluating the numerical expression, we get

$$[H_2]^3 = \frac{(6.62)^2}{(1.88)(13.7)} = 1.7015219754$$

To solve for [H_2], we need to take the cube root of the equation. Performing this operation, we get

[H_2] = 1.19 M

You should verify that this is correct using your own calculator to confirm that you know how to do a cube root correctly.

Exercise 7.2.2

The following reaction is at equilibrium:

$$\mathrm{N_2} + 3\,\mathrm{H_2} \rightleftharpoons 2\,\mathrm{NH_3} \tag{7.2.4}$$

The K_{eq} at a particular temperature is 13.7. If the equilibrium [N_2] is 0.055 M and the equilibrium [H_2] is 1.62 M, what is the equilibrium [NH_3]?

Answer available at the end of this section.

The K_{eq} was defined earlier in terms of concentrations. For gas-phase reactions, the K_{eq} can also be defined in terms of the partial pressures of the reactants and products, P_i. For the gas-phase reaction

$$aA(g) + bB(g) \rightleftharpoons cC(g) + dD(g)$$

the pressure-based equilibrium constant, K_P, is defined as follows:

$$K_P = \frac{P_C^c P_D^d}{P_A^a P_B^b}$$

where P_A is the partial pressure of substance A at equilibrium in atmospheres, and so forth. As with the concentration-based equilibrium constant, the units are omitted when substituting into the expression for K_P.

Example 7.2.3

What is the K_P for this reaction, given the equilibrium partial pressures of 0.664 atm for NO_2 and 1.09 for N_2O_4?

$$2\,NO_2\,(g) \rightleftharpoons N_2O_4\,(g) \tag{7.2.5}$$

Solution

Write the K_P expression for this reaction:

$$K_P = \frac{P_{N_2O_4}}{P_{NO_2}^2}$$

Then substitute the equilibrium partial pressures into the expression and evaluate:

$$K_P = \frac{(1.09)}{(0.664)^2} = 2.47$$

Exercise 7.2.3

What is the K_P for this reaction, given the equilibrium partial pressures of 0.44 atm for H_2, 0.22 atm for Cl_2, and 2.98 atm for HCl?

$$H_2 + Cl_2 \rightleftharpoons 2\,HCl \tag{7.2.6}$$

Answer available at the end of this section.

There is a simple relationship between K_{eq} (based on concentration units) and K_P (based on pressure units):

$$K_P = K_{eq} \cdot (RT)^{\Delta n}$$

where R is the ideal gas law constant (in units of L·atm/mol·K), T is the absolute temperature, and Δn is the change in the number of moles of gas in the balanced chemical equation, defined as $n_{gas,prods} - n_{gas,rcts}$. Note that this equation implies that if the number of moles of gas are the same in reactants and products, $K_{eq} = K_P$.

Example 7.2.4

What is the K_P at 25°C for this reaction if the K_{eq} is 4.2×10^{-2}?

$$N_2 + 3\,H_2 \rightleftharpoons 2\,NH_3 \tag{7.2.7}$$

Solution

Before we use the relevant equation, we need to do two things: convert the temperature to kelvins and determine Δn. Converting the temperature is easy:

$T = 25 + 273 = 298$ K

To determine the change in the number of moles of gas, take the number of moles of gaseous products and subtract the number of moles of gaseous reactants. There are 2 mol of gas as product and 4 mol of gas of reactant:

$\Delta n = 2 - 4 = -2$ mol

Note that Δn is negative. Now we can substitute into our equation, using R = 0.08205 L·atm/mol·K. The units are omitted for clarity:

$K_P = (4.2 \times 10^{-2})(0.08205)(298)^{-2}$

Solving,

$K_P = 7.0 \times 10^{-5}$

Exercise 7.2.4

What is the K_P at 25°C for this reaction if the K_{eq} is 98.3?,-

$$\mathrm{I_2\,(g) \rightleftharpoons 2\,I\,(g)} \tag{7.2.8}$$

Answer available at the end of this section.

Finally, we recognize that many chemical reactions involve substances in the solid or liquid phases. For example, a particular chemical reaction is represented as follows:

$$\mathrm{2\,NaHCO_3\,(s) \rightleftharpoons Na_2CO_3\,(s) + CO_2\,(g) + H_2O\,(l)} \tag{7.2.9}$$

This chemical equation includes all three phases of matter. This kind of equilibrium is called a **heterogeneous equilibrium** because there is more than one phase present.

The rule for heterogeneous equilibria is as follows: *Do not include the concentrations of pure solids and pure liquids in* K_{eq} *expressions.* Only partial pressures for gas-phase substances or concentrations in solutions are included in the expressions of equilibrium constants. As such, the equilibrium constant expression for this reaction would simply be

$$K_P = P_{CO_2}$$

because the two solids and one liquid would not appear in the expression.

Key Takeaways

- Every chemical equilibrium can be characterized by an equilibrium constant, known as K_{eq}.
- The K_{eq} and K_P expressions are formulated as amounts of products divided by amounts of reactants; each amount (either a concentration or a pressure) is raised to the power of its coefficient in the balanced chemical equation.
- Solids and liquids do not appear in the expression for the equilibrium constant.

Learning Objectives

- Define *Le Chatelier's principle*.
- Predict the direction of shift for an equilibrium under stress.

Once equilibrium is established, the reaction is over, right? Not exactly. An experimenter has some ability to affect the equilibrium.

Chemical equilibria can be shifted by changing the conditions that the system experiences. We say that we "stress" the equilibrium. When we stress the equilibrium, the chemical reaction is no longer at equilibrium, and the reaction starts to move back toward equilibrium in such a way as to decrease the stress. The formal statement is called **Le Chatelier's principle**: If an equilibrium is stressed, then the reaction shifts to reduce the stress.

There are several ways to stress an equilibrium. One way is to add or remove a product or a reactant in a chemical reaction at equilibrium. When additional reactant is added, the equilibrium shifts to reduce this stress: it makes more product. When

additional product is added, the equilibrium shifts to reactants to reduce the stress. If reactant or product is removed, the equilibrium shifts to make more reactant or product, respectively, to make up for the loss.

Example 7.2.1

Given this reaction at equilibrium:

$$N_2 + 3H_2 \rightleftharpoons 2NH_3$$

In which direction—toward reactants or toward products-—does the reaction shift if the equilibrium is stressed by each change?

1. H_2 is added.
2. NH_3 is added.
3. NH_3 is removed.

Solution

1. If H_2 is added, there is now more reactant, so the reaction will shift toward products to reduce the added H_2.
2. If NH_3 is added, there is now more product, so the reaction will shift toward reactants to reduce the added NH_3.
3. If NH_3 is removed, there is now less product, so the reaction will shift toward products to replace the product removed.

Exercise 7.2.1

Given this reaction at equilibrium:

$$CO(g) + Br_2(g) \rightleftharpoons COBr_2(g)$$

In which direction—toward reactants or toward products—does the reaction shift if the equilibrium is stressed by each change?

1. Br_2 is removed.
2. $COBr_2$ is added.

Answers

1. toward reactants
2. toward reactants

It is worth noting that when reactants or products are added or removed, *the value of the* K_{eq} *does not change*. The chemical reaction simply shifts, in a predictable fashion, to reestablish concentrations so that the K_{eq} expression reverts to the correct value.

How does an equilibrium react to a change in pressure? Pressure changes do not markedly affect the solid or liquid phases. However, pressure strongly impacts the gas phase. Le Chatelier's principle implies that a pressure increase shifts an equilibrium to the side of the reaction with the fewer number of moles of gas, while a pressure decrease shifts an equilibrium to the side of the reaction with the greater number of moles of gas. If the number of moles of gas is the same on both sides of the reaction, pressure has no effect.

Example 7.2.2

What is the effect on this equilibrium if pressure is increased?

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$$

Solution

According to Le Chatelier's principle, if pressure is increased, then the equilibrium shifts to the side with the fewer number of moles of gas. This particular reaction shows a total of 4 mol of gas as reactants and 2 mol of gas as products, so the reaction shifts toward the products side.

Exercise 7.2.2

What is the effect on this equilibrium if pressure is decreased?

$$3O_2(g) \rightleftharpoons 2O_3(g)$$

Answer available at the end of this section.

What is the effect of temperature changes on an equilibrium? It depends on whether the reaction is endothermic or exothermic. Recall that *endothermic* means that energy is absorbed by a chemical reaction, while *exothermic* means that energy is given off by the reaction. As such, energy can be thought of as a reactant or a product, respectively, of a reaction:

endothermic: energy + reactants → products

exothermic: reactants → products + energy

Because temperature is a measure of the energy of the system, increasing temperature can be thought of as adding energy. The reaction will react as if a reactant or a product is being added and will act accordingly by shifting to the other side. For example, if the temperature is increased for an endothermic reaction, essentially a reactant is being added, so the equilibrium shifts toward products. Decreasing the temperature is equivalent to decreasing a reactant (for endothermic reactions) or a product (for exothermic reactions), and the equilibrium shifts accordingly.

Example 7.2.3

Predict the effect of increasing the temperature on this equilibrium.

$$PCl_3 + Cl_2 \rightleftharpoons PCl_5 + 60kJ$$

Solution

Because energy is listed as a product, it is being produced, so the reaction is exothermic. If the temperature is increasing, a product is being added to the equilibrium, so the equilibrium shifts to minimize the addition of extra product: it shifts back toward reactants.

Exercise 7.2.3

Predict the effect of decreasing the temperature on this equilibrium.

$$N_2O_4 + 57kJ \rightleftharpoons 2NO_2$$

Answer available at the end of this section.

In the case of temperature, the value of the equilibrium has changed because the K_{eq} is dependent on temperature. That is why equilibria shift with changes in temperature.

A **catalyst** is a substance that increases the speed of a reaction. Overall, a catalyst is not a reactant and is not used up, but it still affects how fast a reaction proceeds. However, a catalyst does not affect the extent or position of a reaction at equilibrium. It helps a reaction achieve equilibrium faster.

Chemistry is Everywhere: Equilibria in the Garden

Hydrangeas are common flowering plants around the world. Although many hydrangeas are white, there is one common species (*Hydrangea macrophylla*) whose flowers can be either red or blue, as shown in the accompanying figure. How is it that a plant can have different colored flowers like this?



Figure 7.2.1 Garden Equilibria © Thinkstock. This species of hydrangea has flowers that can be either red or blue. Why the color difference?

Interestingly, the color of the flowers is due to the acidity of the soil that the hydrangea is planted in. An astute gardener can adjust the pH of the soil and actually change the color of the flowers. However, it is not the H^+ or OH^- ions that affect the color of the flowers. Rather, it is the presence of aluminum that causes the color change.

The solubility of aluminum in soil, and the ability of plants to absorb it, is dependent upon the acidity of the soil. If the soil is relatively acidic, the aluminum is more soluble, and plants can absorb it more easily. Under these conditions, hydrangea flowers are blue, as Al ions interact with anthocyanin pigments in the plant. In more basic soils, aluminum is less soluble, and under these conditions the hydrangea flowers are red. Gardeners who change the pH of their soils to change the color of their hydrangea flowers are therefore employing Le Chatelier's principle: the amount of acid in the soil changes the equilibrium of aluminum solubility, which in turn affects the color of the flowers.

Key Takeaways

- Le Chatelier's principle addresses how an equilibrium shifts when the conditions of an equilibrium are changed.
- The direction of shift can be predicted for changes in concentrations, temperature, or pressure.
- Catalysts do not affect the position of an equilibrium; they help reactions achieve equilibrium faster.

Exercise 7.2.1

1. Define *Le Chatelier's principle*.
2. What is meant by a stress? What are some of the ways an equilibrium can be stressed?
3. Given this equilibrium, predict the direction of shift for each stress.

$$H_2(g) + I_2(s) + 53kJ \rightleftharpoons 2HI(g)$$

   a. decreased temperature
   b. increased pressure

c. removal of HI

4. Given this equilibrium, predict the direction of shift for each stress.

$$H_2(g) + F_2(g) \rightleftharpoons 2HF(g) + 546kJ$$

   a. increased temperature
   b. addition of H_2
   c. decreased pressure

5. Given this equilibrium, predict the direction of shift for each stress.

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g) + 196kJ$$

   a. removal of SO_3
   b. addition of O_2
   c. decreased temperature

6. Given this equilibrium, predict the direction of shift for each stress.

$$CO_2(g) + C(s) + 171kJ \rightleftharpoons 2CO(g)$$

   a. addition of CO
   b. increased pressure
   c. addition of a catalyst

7. The synthesis of NH_3 uses this chemical reaction.

$$N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + 92kJ$$

   Identify three stresses that can be imposed on the equilibrium to maximize the amount of NH_3.

8. The synthesis of $CaCO_3$ uses this chemical reaction.

$$CaO(s) + CO_2(g) \rightleftharpoons CaCO_3(s) + 180kJ$$

   Identify three stresses that can be imposed on the equilibrium to maximize the amount of $CaCO_3$.

Answers

1. When an equilibrium is stressed, the equilibrium shifts to minimize that stress.
2.
3. a. toward reactants

- toward reactants
- toward products
-
- a. toward products
  b. toward products
  c. toward products
-
- increased pressure, decreased temperature, removal of NH_3

Learning Objective

- Identify several special chemical equilibria and construct their K_a expressions.

In one sense, all chemical equilibria are treated the same. However, there are several classes of reactions that are noteworthy because of either the identities of the reactants and products, or the form of the K_{eq} expression.

Weak Acids and Bases

In Chapter 12 - Acids and Bases, we noted how some acids and bases are strong and some are weak. If an acid or base is strong,

it is ionized 100% in H_2O. HCl(aq) is an example of a strong acid:

$$HCl(aq) \overset{100\%}{\rightarrow} H^+(aq) + Cl^-(aq)$$

However, if an acid or base is weak, it may dissolve in H_2O, but does not ionize completely. This means that there is an equilibrium between the unionized acid or base and the ionized form. $HC_2H_3O_2$ is an example of a weak acid:

$$HC_2H_3O_2(aq) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

$HC_2H_3O_2$ is soluble in H_2O (in fact, it is the acid in vinegar), so the reactant concentration will appear in the equilibrium constant expression. But not all the molecules separate into ions. This is the case for all weak acids and bases.

An **acid dissociation constant**, K_a, is the equilibrium constant for the dissociation of a weak acid into ions. Note the *a* subscript on the *K*; it implies that the substance is acting as an acid. The larger K_a is, the stronger the acid is. Table 7.2.1 - Acid Dissociation Constants for Some Weak Acids, lists several acid dissociation constants. Keep in mind that they are just equilibrium constants.

Table, two columns and 7 rows. The first column on the right has different acids in the rows underneath. The second column on the right side has the acid dissociation content for the corresponding acid in the rows underneath.

Acid	K_a
$HC_2H_3O_2$	1.8×10^{-5}
$HClO_2$	1.1×10^{-2}
$H_2PO_4^-$	6.2×10^{-8}
HCN	6.2×10^{-10}
HF	6.3×10^{-4}
HNO_2	5.6×10^{-4}
H_3PO_4	7.5×10^{-3}

Table 7.2.1 *Acid Dissociation Constants for Some Weak Acids*

Note also that the acid dissociation constant refers to *one* H^+ ion coming off of the initial reactant. Thus the acid dissociation constant for H_3PO_4 refers to this equilibrium:

$$H_3PO_4(aq) \rightleftharpoons H^+(aq) + H_2PO_4^-(aq) \quad K_a = 7.5 \times 10^{-3}$$

The $H_2PO_4^-$ ion, called the dihydrogen phosphate ion, is also a weak acid with its own acid dissociation constant:

$$H_2PO_4^-(aq) \rightleftharpoons H^+(aq) + HPO_4^{2-}(aq) \quad K_a = 6.2 \times 10^{-8}$$

Thus, for so-called *polyprotic* acids, each H^+ ion comes off in sequence; each H^+ ion that ionizes does so with its own characteristic K_a.

Example 7.2.1

Write the equilibrium equation and the K_a expression for HSO_4^- acting as a weak acid.

Solution

HSO_4^- acts as a weak acid by separating into an H^+ ion and an SO_4^{2-} ion:

$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

The K_a is written just like any other equilibrium constant, in terms of the concentrations of products divided by concentrations of reactants:

$$K_a = \frac{[H^+][SO_4^{2-}]}{[HSO_4^-]}$$

Exercise 7.2.1

Write the equilibrium equation and the K_a expression for HPO_4^{2-} acting as a weak acid.

Answer available at the end of this section.

The K_a is used in equilibrium constant problems just like other equilibrium constants are. However, in some cases, we can simplify the mathematics if the numerical value of the K_a is small, much smaller than the concentration of the acid itself. Example 11 illustrates this.

Example 7.2.2

What is the pH of a 1.00 M solution of $HC_2H_3O_2$? The K_a of $HC_2H_3O_2$ is 1.8×10^{-5}.

Solution

This is a two-part problem. We need to determine $[H^+]$ and then use the definition of pH to determine the pH of the solution. For the first part, we can use an ICE chart:

Solutions to Example 13.6.2

	$HC_2H_3O_2(aq)$	⇄	$H^+(g)$	+	$C_2H_3O_2^-(g)$
I	1.00		0		0
C	$-x$		$+x$		$+x$
E	$1.00 - x$		$+x$		$+x$

We now construct the K_a expression, substituting the concentrations from the equilibrium row in the ICE chart:

$$K_a = \frac{[H^+][C_2H_3O_2^-]}{[HC_2H_3O_2]} = \frac{(x)(x)}{(1.00 - x)} = 1.8 \times 10^{-5}$$

Here is where a useful approximation comes in: at 1.8×10^{-5}, $HC_2H_3O_2$ will not ionize very much, so we expect that the value of x will be small. It should be so small that in the denominator of the fraction, the term $(1.00 - x)$ will likely be very close to 1.00. As such, we would introduce very little error if we simply neglect the x in that term, making it equal to 1.00:

$(1.00 - x) \approx 1.00$ for small values of x

This simplifies the mathematical expression we need to solve:

$$\frac{(x)(x)}{1.00} = 1.8 \times 10^{-5}$$

This is much easier to solve than a more complete quadratic equation. The new equation to solve becomes

$$x^2 = 1.8 \times 10^{-5}$$

Taking the square root of both sides,

$$x = 4.2 \times 10^{-3}$$

Because x is the equilibrium concentration of H^+ and $C_2H_3O_2^-$, we thus have

$$[H^+] = 4.2 \times 10^{-3}\ M$$

Notice that we are justified by neglecting the x in the denominator; it truly is small compared to 1.00. Now we can determine the pH of the solution:

$$pH = -\log[H^+] = -\log(4.2 \times 10^{-3}) = 2.38$$

Exercise 7.2.2

What is the pH of a 0.500 M solution of HCN? The K_a of HCN is 6.2×10^{-10}.

Answer available at the end of this section.

Weak bases also have dissociation constants, labeled K_b (the b subscript stands for base). However, values of K_b are rarely tabulated because there is a simple relationship between the K_b of a base and the K_a of its conjugate acid:

$$K_a \times K_b = 1.0 \times 10^{-14}$$

Thus it is simple to calculate the K_b of a base from the K_a of its conjugate acid.

Example 7.2.3

What is the value of K_b for $C_2H_3O_2^-$, which can accept a proton and act as a base?

Solution

To determine the K_b for $C_2H_3O_2^-$, we need to know the K_a of its conjugate acid. The conjugate acid of $C_2H_3O_2^-$ is $HC_2H_3O_2$. The K_a for $HC_2H_3O_2$ is in Table 7.2.1 "Acid Dissociation Constants for Some Weak Acids" and is 1.8×10^{-5}. Using the mathematical relationship between K_a and K_b:

$$(1.8 \times 10^{-5})K_b = 1.0 \times 10^{-14}$$

Solving,

$$K_b = \frac{1.0 \times 10^{-14}}{1.8 \times 10^{-5}} = 5.6 \times 10^{-10}$$

Exercise 7.2.3

What is the value of K_b for PO_4^{3-}, which can accept a proton and act as a base? The K_a for HPO_4^{2-} is 2.2×10^{-13}.

Answer available at the end of this section.

Autoionization of water

In Chapter 12, "Acids and Bases," we introduced the autoionization of water—the idea that water can act as a proton donor and proton acceptor simultaneously. Because water is not a strong acid (Table 12.5.1 - Strong Acids and Bases), it must be a weak acid, which means that its behavior as an acid must be described as an equilibrium. That equilibrium is as follows:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

The equilibrium constant includes [H_3O^+] and [OH^-] but not [$H_2O(\ell)$] because it is a pure liquid. Hence the expression *does not have any terms in its denominator*:

$K = [H_3O^+][OH^-] \equiv K_W = 1.0 \times 10^{-14}$

This is the same K_W that was introduced in Chapter 12 and the same 1.0×10^{-14} that appears in the relationship between the K_a and the K_b of a conjugate acid-base pair. In fact, we can rewrite this relationship as follows:

$K_a \times K_b = K_W$

Insoluble Compounds

In Chapter 4, section 4.3: "Types of Chemical Reactions - Single and Double Displacement Reactions," the concept of soluble and insoluble compounds was introduced. Solubility rules were presented that allow a person to predict whether certain simple ionic compounds will or will not dissolve.

Describing a substance as soluble or insoluble is a bit misleading because virtually all substances are soluble; they are just soluble to different extents. In particular, for ionic compounds, what we typically describe as an *insoluble* compound can actually be ever so slightly soluble; an equilibrium is quickly established between the solid compound and the ions that do form in solution. Thus, the hypothetical compound MX does in fact dissolve but only very slightly. That means we can write an equilibrium for it:

$$MX(s) \rightleftharpoons M^+(aq) + X^-(aq)$$

The equilibrium constant for a compound normally considered insoluble is called a **solubility product constant** and is labeled K_{sp} (with the subscript *sp*, meaning "solubility product"). Because the reactant is a solid, its concentration does not appear in the K_{sp} expression, so like K_W, expressions for K_{sp} do not have denominators. For example, the chemical equation and the expression for the K_{sp} for AgCl, normally considered insoluble, are as follows:

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq) \qquad K_{sp} = [Ag^+][Cl^-]$$

Table 7.2.2 - Solubility Product Constants for Slightly Soluble Ionic Compounds, lists some values of the K_{sp} for slightly soluble ionic compounds.

Table, two columns and 8 rows. The first column on the right has different compounds in the rows underneath. The second column on the right side has the acid dissociation content for the corresponding compounds in the rows underneath.

Compound	K_{sp}
$BaSO_4$	1.1×10^{-10}

Compound	K_{sp}
$Ca(OH)_2$	5.0×10^{-6}
$Ca_3(PO_4)_2$	2.1×10^{-33}
$Mg(OH)_2$	5.6×10^{-12}
HgI_2	2.9×10^{-29}
$AgCl$	1.8×10^{-10}
AgI	8.5×10^{-17}
Ag_2SO_4	1.5×10^{-5}

Table* 7.2.2 *Solubility Product Constants for Slightly Soluble Ionic Compounds.

Example 7.2.4

Write the K_{sp} expression for $Ca_3(PO_4)_2$.

Solution

Recall that when an ionic compound dissolves, it separates into its individual ions. For $Ca_3(PO_4)_2$, the ionization reaction is as follows:

$$Ca_3(PO_4)_2(s) \rightleftharpoons 3Ca^{2+}(aq) + 2PO_4^{3-}(aq)$$

Hence the K_{sp} expression is

$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2$

Exercise 7.2.4

Write the K_{sp} expression Ag_2SO_4.

Answer available at the end of this section.

Equilibrium problems involving the K_{sp} can also be done, and they are usually more straightforward than other equilibrium problems because there is no denominator in the K_{sp} expression. Care must be taken, however, in completing the ICE chart and evaluating exponential expressions.

Example 7.2.5

What are $[Ag^+]$ and $[Cl^-]$ in a saturated solution of AgCl? The K_{sp} of AgCl is 1.8×10^{-10}.

Solution

The chemical equation for the dissolving of AgCl is

$$AgCl(s) \rightleftharpoons Ag^+(aq) + Cl^-(aq)$$

The K_{sp} expression is as follows:

$K_{sp} = [Ag^+][Cl^-]$

So the ICE chart for the equilibrium is as follows:

Solutions to Example 13.6.5

	$AgCl(s)$	$\rightleftarrows$	$Ag^+(aq)$	+	$Cl^-(aq)$
I			0		0
C	$-x$		$+x$		$+x$
E			$+x$		$+x$

Notice that we have little in the column under AgCl except the stoichiometry of the change; we do not need to know its initial or equilibrium concentrations because its concentration does not appear in the K_{sp} expression. Substituting the equilibrium values into the expression:

$(x)(x) = 1.8 \times 10^{-10}$

Solving,

$x^2 = 1.8 \times 10^{-10} x = 1.3 \times 10^{-5}$

Thus $[Ag^+]$ and $[Cl^-]$ are both 1.3×10^{-5} M.

Exercise 7.2.5

What are $[Ba^{2+}]$ and $[SO_4^{2-}]$ in a saturated solution of $BaSO_4$? The K_{sp} of $BaSO_4$ is 1.1×10^{-10}.

Answer available at the end of this section.

Example 7.2.6

What are $[Ca^{2+}]$ and $[PO_4^{3-}]$ in a saturated solution of $Ca_3(PO_4)_2$? The K_{sp} of $Ca_3(PO_4)_2$ is 2.1×10^{-33}.

Solution

This is similar to Example 14, but the ICE chart is much different because of the number of ions formed.

Solutions to Example 13.6.6

	$Ca_3(PO_4)_2(s)$	$\rightleftarrows$	$3Ca^{2+}(aq)$	+	$2PO_4^{3-}(aq)$
I			0		0
C	$-x$		$+3x$		$+2x$
E			$+3x$		$+2x$

For every unit of $Ca_3(PO_4)_2$ that dissolves, three Ca^{2+} ions and two PO_4^{3-} ions are formed. The expression for the K_{sp} is also different:

$K_{sp} = [Ca^{2+}]^3[PO_4^{3-}]^2 = 2.1 \times 10^{-33}$

Now when we substitute the unknown concentrations into the expression, we get

$(3x)^3(2x)^2 = 2.1 \times 10^{-33}$

When we raise each expression inside parentheses to the proper power, remember that the power affects everything inside the parentheses, including the number. So,

$(27x^3)(4x^2) = 2.1 \times 10^{-33}$

Simplifying,

$108x^5 = 2.1 \times 10^{-33}$

Dividing both sides of the equation by 108, we get

$x^5 = 1.9 \times 10^{-35}$

Now we take the fifth root of both sides of the equation (be sure you know how to do this on your calculator):

$x = 1.1 \times 10^{-7}$

We are not done yet. We still need to determine the concentrations of the ions. According to the ICE chart, $[Ca^{2+}]$ is $3x$, not x. So

$[Ca^{2+}] = 3x = 3 \times 1.1 \times 10^{-7} = 3.3 \times 10^{-7}$ M

$[PO_4^{3-}]$ is $2x$, so

$[PO_4^{3-}] = 2x = 2 \times 1.1 \times 10^{-7} = 2.2 \times 10^{-7}$ M

Exercise 7.2.6

What are $[Mg^{2+}]$ and $[OH^-]$ in a saturated solution of $Mg(OH)_2$? The K_{sp} of $Mg(OH)_2$ is 5.6×10^{-12}.

Answer available at the end of this section.

Food and Drink Application: Solids in Your Wine Bottle

People who drink wine from bottles (as opposed to boxes) will occasionally notice some insoluble materials in the wine, either crusting the bottle, stuck to the cork, or suspended in the liquid wine itself. The accompanying figure shows a cork encrusted with colored crystals. What are these crystals?

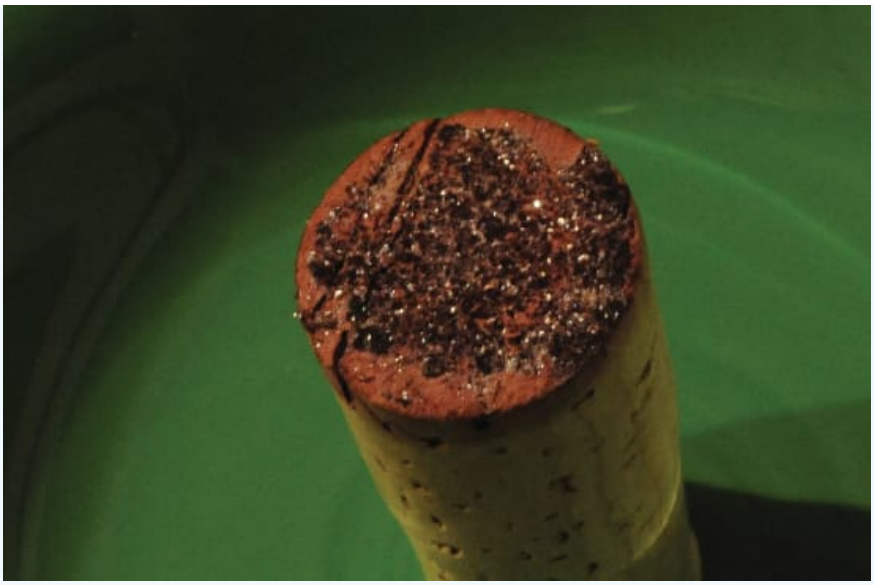


Figure 7.2.1 Wine Cork. The red crystals on the top of the wine cork are from insoluble compounds that are not soluble in the wine. Source: Photo courtesy of Paul A. Hernandez, flickr(opens in new window).

One of the acids in wine is tartaric acid ($H_2C_4H_4O_6$). Like the other acids in wine (citric and malic acids, among others), tartaric acid imparts a slight tartness to the wine. Tartaric acid is rather soluble in H_2O, dissolving over 130 g of the acid in only 100 g of H_2O. However, the potassium salt of singly ionized tartaric acid, potassium hydrogen tartrate ($KHC_4H_4O_6$; also known as potassium bitartrate and better known in the kitchen as cream of tartar), has a solubility of only 6 g per 100 g of H_2O. Thus, over time, wine stored at cool temperatures will slowly precipitate potassium hydrogen tartrate. The crystals precipitate in the wine or grow on the insides of the wine bottle and, if the bottle is stored on its side, on the bottom of the cork. The color of the crystals comes from pigments in the wine; pure potassium hydrogen tartrate is clear in its crystalline form, but in powder form it is white.

The crystals are harmless to ingest; indeed, cream of tartar is used as an ingredient in cooking. However, most wine drinkers do not like to chew their wine, so if tartrate crystals are present in a wine, the wine is usually filtered or decanted to remove the crystals. Tartrate crystals are almost exclusively in red wines; white and rose wines do not have as much tartaric acid in them.

Key Takeaway

- Equilibrium constants exist for certain groups of equilibria, such as weak acids, weak bases, the autoionization of water, and slightly soluble salts.

Answers

Exercise 7.2.1

$$2H_2 + O_2 + \rightleftharpoons 2H_2O$$

Exercise 7.2.1

0.010 M

Exercise 7.2.2

1.79 M

Exercise 7.2.3

91.7

.Exercise 7.2.4

2.40×10^3

Exercise 7.2.2

Reaction shifts toward reactants.

Exercise 7.2.3

Equilibrium shifts toward reactants.

Exercise 7.2.1

$$HPO_4^{2-}(aq) \rightleftharpoons H^+(aq) + PO_4^{3-}(aq) \qquad K_a = \frac{[H^+][PO_4^{3-}]}{[HPO_4^{2-}]}$$

Exercise 7.2.2

4.75

Exercise 7.2.3

4.5×10^{-2}

Exercise 7.2.4

$K_{sp} = [Ag^+]^2[SO_4^{2-}]$

Exercise 7.2.5

1.0×10^{-5} M

Exercise 7.2.6

$[Mg^{2+}] = 1.1 \times 10^{-4}$ M; $[OH^-] = 2.2 \times 10^{-4}$ M

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7.3: Acids and Bases

Learning Objectives

- Examine properties of acids and bases.

Many people enjoy drinking coffee. A cup first thing in the morning helps start the day. But keeping the coffee maker clean can be a problem. Lime deposits build up after a while and slow down the brewing process. The best cure for this is to put vinegar (dilute acetic acid) in the pot and run it through the brewing cycle. The vinegar dissolves the deposits and cleans the maker, which will speed up the brewing process back to its original rate. Just be sure to run water through the brewing process after the vinegar, or you will get some really horrible coffee.

Perhaps you have eaten too much pizza and felt very uncomfortable hours later. This feeling is due to excess stomach acid being produced. The discomfort can be dealt with by taking an antacid. The base in the antacid will react with the HCl in the stomach and neutralize it, taking care of that unpleasant feeling. Figure $7.3.1$ provides some examples of acids and bases.



Figure $7.3.1$ Some examples of common household acids and bases.

Acids

Acids are very common in some of the foods that we eat. Citrus fruits such as oranges and lemons contain citric acid and ascorbic acid, which is better known as vitamin C. Carbonated sodas contain phosphoric acid. Vinegar contains acetic acid. Your own stomach utilizes hydrochloric acid to digest food. Acids are a distinct class of compounds because of the properties of their aqueous solutions as outlined below:

1. Aqueous solutions of acids are electrolytes, meaning that they conduct electrical current. Some acids are strong electrolytes because they ionize completely in water, yielding a great many ions. Other acids are weak electrolytes that exist primarily in a non-ionized form when dissolved in water.
2. Acids have a sour taste. Lemons, vinegar, and sour candies all contain acids.
3. Acids change the color of certain acid-base indicators. Two common indicators are litmus and phenolphthalein. Blue litmus turns red in the presence of an acid, while phenolphthalein turns colorless.
4. Acids react with active metals to yield hydrogen gas. Recall that an activity series is a list of metals in descending order of reactivity. Metals that are above hydrogen in the activity series will replace the hydrogen from an acid in a single-replacement reaction, as shown below:

$$\mathrm{Zn}\,(s) + \mathrm{H_2SO_4}\,(aq) \rightarrow \mathrm{ZnSO_4}\,(aq) + \mathrm{H_2}\,(g) \tag{7.3.1}$$

5. Acids react with bases to produce a salt compound and water. When equal moles of an acid and a base are combined, the acid is neutralized by the base. The products of this reaction are an ionic compound, which is labeled as a salt, and water.

Common examples: Lemons, oranges, vinegar, urine, sulfuric acid, hydrochloric acid

Bases

Bases have properties that mostly contrast with those of acids.

1. Aqueous solutions of bases are also electrolytes. Bases can be either strong or weak, just as acids can.
2. Bases often have a bitter taste and are found in foods less frequently than acids. Many bases, like soaps, are slippery to the touch.
3. Bases also change the color of indicators. Litmus turns blue in the presence of a base while phenolphthalein turns pink.
4. Bases do not react with metals in the way that acids do.
5. Bases react with acids to produce a salt and water.

Common Examples: Soap, toothpaste, bleach, cleaning agents, limewater, ammonia water, sodium hydroxide.



Figure 7.3.2 Phenolphthalein indicator in presence of base.

Warning

Tasting chemicals and touching them are NOT good lab practices and should be avoided - in other words, don't do this at home.

Summary

- A brief summary of properties of acids and bases are given.
- Properties of acids and bases mostly contrast each other.

Contributors and Attributions

-
- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
- Wikipedia
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Formerly there were rather campy science-fiction television shows in which the hero was always being threatened with death by being plunged into a vat of boiling acid: "Mwa ha ha, Buck Rogers [or whatever the hero's name was], prepare to meet your doom by being dropped into a vat of boiling acid!" (The hero always escapes, of course.) This may have been interesting drama, but not very good chemistry. If the villain knew his/her/its science, the hero would have been dropped into a vat of boiling base.

Recall that the active component of a classic acid is the H^+ ion, while the active part of a classic base is the OH^- ion. Both ions are related to water in that all H^+ ion needs to become a water molecule is an OH^- ion, while all an OH^- ion needs to become water is an H^+ ion. Consider the relative masses involved: an ion of mass 1 needs an ion of mass 17 to make water, while an ion of mass 17 needs an ion of mass 1 to make water. Which process do you think will be easier?

In fact, bases are more potentially dangerous than acids because it is much easier for an OH^- ion to rip off an H^+ ion from surrounding matter than it is for an H^+ ion to rip off an OH^- ion. Certain household chemicals, such as some brands of cleaner, can be very concentrated bases, which makes them among the most potentially hazardous substances found around the home. If spilled on the skin, these strong caustic compounds can immediately remove H^+ ions from the flesh, resulting in chemical burns. Compare this to the fact that we occasionally purposefully ingest substances such as citrus fruits, vinegar, and wine—all of which contain acids. (Of course, some parts of the body, such as the eyes, are extremely sensitive to acids as well as bases.) It seems that our bodies are more capable of dealing with acids than with bases.



Figure 7.3.1 Household chemicals © Thinkstock. On the left is a common acid, and on the right is a common base. Which one is more potentially hazardous?

A note to all of the villains out there...get your chemistry right if you want to be successful!

Learning Objectives

- Identify an Arrhenius acid and an Arrhenius base.
- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Write chemical reactions between an Arrhenius acid and an Arrhenius base and between Brønsted-Lowry acid and a Brønsted-Lowry base.

There are three major classifications of substances known as acids or bases. The Arrhenius definition states that an acid produces H^+ in solution and a base produces OH^-. This theory was developed by Svante Arrhenius in 1883. Later, two more sophisticated and general theories were proposed. These are the Brønsted-Lowry and the Lewis definitions of acids and bases. The Lewis theory is discussed elsewhere.

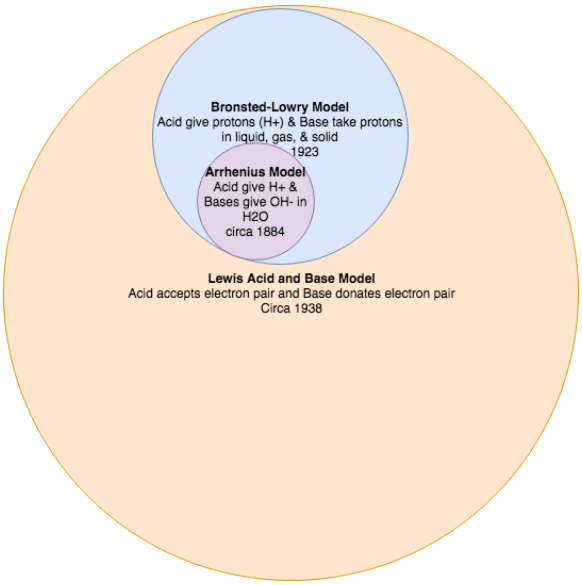


Figure 7.3.1 Acid base reaction theories as superset and subset models.

The Arrhenius Theory of Acids and Bases

In 1884, the Swedish chemist Svante Arrhenius proposed two specific classifications of compounds, termed acids and bases. When dissolved in an aqueous solution, certain ions were released into the solution. An **Arrhenius acid** is a compound that **increases the concentration of** $\mathrm{H^+}$ **ions** that are present when added to water. These H^+ ions form the hydronium ion (H_3O^+) when they combine with water molecules. This process is represented in a chemical equation by adding H_2O to the reactants side.

$$\mathrm{HCl\,(aq)\ \rightarrow H^+\,(aq) + Cl^-\,(aq)}$$

In this reaction, hydrochloric acid (HCl) dissociates completely into hydrogen (H^+) and chlorine (Cl^-) ions when dissolved in water, thereby releasing H^+ ions into solution. Formation of the hydronium ion equation:

$$\mathrm{HCl\,(aq) + H_2O\,(l)\ \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)}$$

An **Arrhenius base** is a compound that **increases the concentration of** $\mathrm{OH^-}$ **ions** that are present when added to water. The dissociation is represented by the following equation:

$$\mathrm{NaOH\ \ (aq)\ \rightarrow Na^+\ \ (aq) + OH^-\ \ (aq)}$$

In this reaction, sodium hydroxide (NaOH) disassociates into sodium (Na^+) and hydroxide (OH^-) ions when dissolved in water, thereby releasing OH^- ions into solution.

> Arrhenius acids are substances which produce hydrogen ions in solution and Arrhenius bases are substances which produce hydroxide ions in solution.

Limitations to the Arrhenius Theory

The Arrhenius theory has many more limitations than the other two theories. The theory does not explain the weak base ammonia (NH_3), which in the presence of water, releases hydroxide ions into solution, but does not contain OH- itself. Also, the Arrhenius definition of acid and base is limited to aqueous (i.e., water) solutions.

The Brønsted-Lowry Theory of Acids and Bases

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, **a Brønsted-Lowry acid** is a **proton donor (PD)**, while **a Brønsted-Lowry base** is a **proton acceptor (PA)**.

> A Brønsted-Lowry acid is a proton donor, while a Brønsted-Lowry base is a proton acceptor.

Let us use the reaction of ammonia in water to demonstrate the Brønsted-Lowry definitions of an acid and a base. Ammonia and water molecules are reactants, while the ammonium ion and the hydroxide ion are products:

$$\mathrm{NH_3\,(aq) + H_2O\ (\ell) \rightleftharpoons NH_4^+\,(aq) + OH^-\,(aq)} \tag{7.3.2}$$

What has happened in this reaction is that the original water molecule has donated a hydrogen ion to the original ammonia molecule, which in turn has accepted the hydrogen ion. We can illustrate this as follows (Figure 7.3.2):

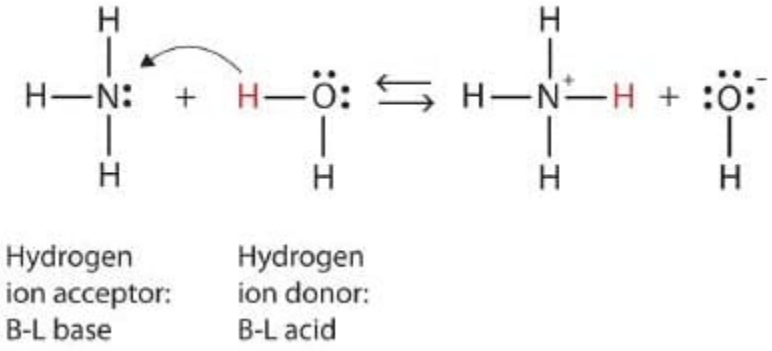


Figure 7.3.2 reaction of ammonia and water.

Because the water molecule donates a hydrogen ion to the ammonia, it is the Brønsted-Lowry acid, while the ammonia molecule—which accepts the hydrogen ion—is the Brønsted-Lowry base. Thus, ammonia acts as a base in both the Arrhenius sense and the Brønsted-Lowry sense.

Is an Arrhenius acid like hydrochloric acid still an acid in the Brønsted-Lowry sense? Yes, but it requires us to understand what really happens when HCl is dissolved in water. Recall that the hydrogen *atom* is a single proton surrounded by a single electron. To make the hydrogen *ion*, we remove the electron, leaving a bare proton. Do we *really* have bare protons floating around in aqueous solution? No, we do not. What really happens is that the H^+ ion attaches itself to H_2O to make H_3O^+, which is called the *hydronium ion*. For most purposes, H^+ and H_3O^+ represent the same species, but writing H_3O^+ instead of H^+ shows that we understand that there are no bare protons floating around in solution. Rather, these protons are actually attached to solvent molecules.

The Hydronium Ion

A proton in aqueous solution may be surrounded by more than one water molecule, leading to formulas like $\mathrm{H_5O_2^+}$ or $\mathrm{H_9O_4^+}$ rather than $\mathrm{H_3O^+}$. It is simpler, however, to use $\mathrm{H_3O^+}$ to represent the hydronium ion.

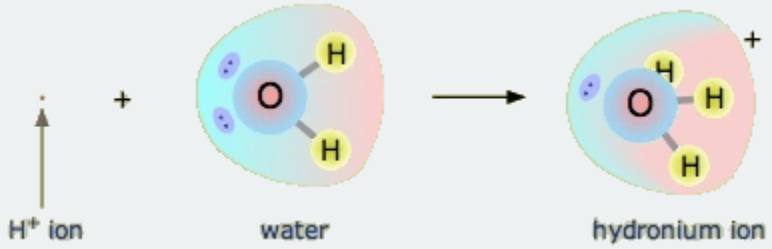


Figure 7.3.3 reaction of hydrogen and water to form hydronium. (CC BY-NC 3.0

With this in mind, how do we define HCl as an acid in the Brønsted-Lowry sense? Consider what happens when HCl is dissolved in H_2O:

$$\mathrm{HCl\,(g) + H_2O\ (\ell)\ \rightarrow H_3O^+\,(aq) + Cl^-\,(aq)} \tag{7.3.3}$$

We can depict this process using Lewis electron dot diagrams (Figure 7.3.4):

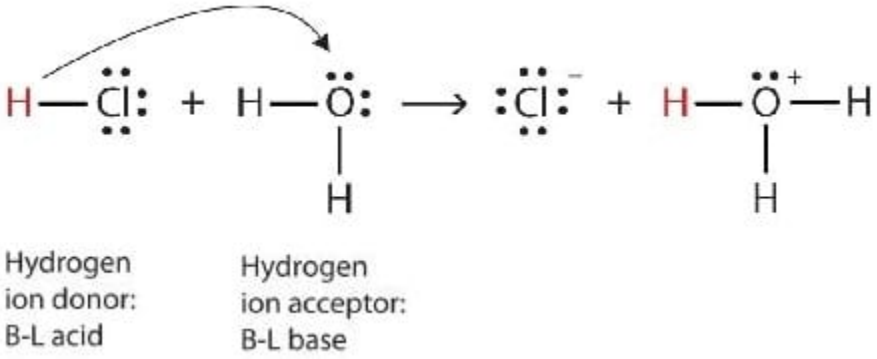


Figure 7.3.4 Lewis electron dot diagram of HCl dissolved in H_2O

Now we see that a hydrogen ion is transferred from the HCl molecule to the H_2O molecule to make chloride ions and hydronium ions. As the hydrogen ion donor, HCl acts as a Brønsted-Lowry acid; as a hydrogen ion acceptor, H_2O is a Brønsted-Lowry base. So HCl is an acid not just in the Arrhenius sense but also in the Brønsted-Lowry sense. Moreover, by the Brønsted-Lowry definitions, H_2O is a base in the formation of aqueous HCl. So the Brønsted-Lowry definitions of an acid and a base classify the dissolving of HCl in water as a reaction between an acid and a base—although the Arrhenius definition would not have labeled H_2O a base in this circumstance.

A summary of the acid-base definitions based on the Arrhenius and Bronstes-Lowry theories is given in Table 7.3.1.

Table 7.3.1 Acid-Base Definitions

Type	Acid	Base
Arrhenius	H^+ ions in solution	OH^- ions in solution
Brønsted-Lowry	H^+ donor	H^+ acceptor

All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

Example 7.3.1

Aniline ($C_6H_5NH_2$) is slightly soluble in water. It has a nitrogen atom that can accept a hydrogen ion from a water molecule just like the nitrogen atom in ammonia does. Write the chemical equation for this reaction and identify the Brønsted-Lowry acid and base.

Solution

$C_6H_5NH_2$ and H_2O are the reactants. When $C_6H_5NH_2$ accepts a proton from H_2O, it gains an extra H and a positive charge and leaves an OH^- ion behind. The reaction is as follows:

$$\mathrm{C_6H_5NH_2\,(aq) + H_2O(\ell) \rightleftharpoons C_6H_5NH_3{}^+\,(aq) + OH^-\,(aq)}$$

Because $C_6H_5NH_2$ accepts a proton, it is the Brønsted-Lowry base. The H_2O molecule, because it donates a proton, is the Brønsted-Lowry acid.

Exercise 7.3.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$H_2PO_4^- + H_2O \rightleftharpoons HPO_4^{2-} + H_3O^+$$

Answer available at the end of this section.

Exercise 7.3.2

Which of the following compounds is a Bronsted-Lowry base?

a. HCl
b. HPO_4^{2-}
c. H_3PO_4
d. NH_4^+
e. $CH_3NH_3^+$

Answer: available at the end of this section.

Summary

- An Arrhenius acid is a compound that increases the H^+ ion concentration and an Arrhenius base is a compound that increases the OH^- ion concentration in aqueous solution.
- A Brønsted-Lowry acid is a proton donor; a Brønsted-Lowry base is a proton acceptor.
- All Arrhenius acids and bases are Brønsted-Lowry acids and bases as well. However, not all Brønsted-Lowry acids and bases are Arrhenius acids and bases.

Contributors and Attributions

- Libretext: The Basics of GOB Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Libretext: Chemistry for Allied Health (Soult)

Learning Objective

- Identify an Arrhenius acid and an Arrhenius base.
- Write the chemical reaction between an Arrhenius acid and an Arrhenius base.

Historically, the first chemical definition of an acid and a base was put forward by Svante Arrhenius, a Swedish chemist, in 1884. An **Arrhenius acid** is a compound that increases the H^+ ion concentration in an aqueous solution. The H^+ ion is just a bare proton, and it is rather clear that bare protons are not floating around in an aqueous solution. Instead, chemistry has defined the **hydronium ion** ($H3O^+$) as the actual chemical species that represents an H^+ ion. H^+ ions and $H3O^+$ ions are often considered interchangeable when writing chemical equations (although a properly balanced chemical equation should also include the additional $H2O$). Classic Arrhenius acids can be considered ionic compounds in which H^+ is the cation. Table 7.3.1 lists some Arrhenius acids and their names.

Table 7.3.1 Some Arrhenius Acids

Formula	Name
$HC_2H_3O_2$ (also written CH_3COOH)	acetic acid
$HClO_3$	chloric acid
HCl	hydrochloric acid
HBr	hydrobromic acid
HI	hydriodic acid
HF	hydrofluoric acid
HNO_3	nitric acid
$H_2C_2O_4$	oxalic acid
$HClO_4$	perchloric acid
H_3PO_4	phosphoric acid
H_2SO_4	sulfuric acid
H_2SO_3	sulfurous acid

An **Arrhenius base** is a compound that increases the OH^- ion concentration in aqueous solution. Ionic compounds of the OH^- ion are classic Arrhenius bases.

Example 7.3.1

Identify each compound as an Arrhenius acid, an Arrhenius base, or neither.

a. HNO_3
b. CH_3OH
c. $Mg(OH)_2$

Solution

a. This compound is an ionic compound between H^+ ions and NO_3^- ions, so it is an Arrhenius acid.
b. Although this formula has an OH in it, we do not recognize the remaining part of the molecule as a cation. It is neither an acid nor a base. (In fact, it is the formula for methanol, an organic compound.)
c. This formula also has an OH in it, but this time we recognize that the magnesium is present as Mg^{2+} cations. As such, this is an ionic compound of the OH^- ion and is an Arrhenius base.

Exercise 7.3.1

Identify each compound as an Arrhenius acid, an Arrhenius base, or neither.

1. KOH
2. H_2SO_4
3. C_2H_6

Answer

1. Arrhenius base

2. Arrhenius acid
3. neither

Acids have some properties in common. They turn litmus, a plant extract, red. They react with some metals to give off H_2 gas. They react with carbonate and hydrogen carbonate salts to give off CO_2 gas. Acids that are ingested typically have a sour, sharp taste. (The name *acid* comes from the Latin word *acidus*, meaning "sour.") Bases also have some properties in common. They are slippery to the touch, turn litmus blue, and have a bitter flavor if ingested.

Acids and bases have another property: they react with each other to make water and an ionic compound called a salt. A **salt**, in chemistry, is any ionic compound made by combining an acid with a base. A reaction between an acid and a base is called a **neutralization reaction** and can be represented as:

$$\text{acid} + \text{base} \rightarrow H_2O + \text{salt}$$

The stoichiometry of the balanced chemical equation depends on the number of H^+ ions in the acid and the number of OH^- ions in the base.

Example 7.3.2

Write the balanced chemical equation for the neutralization reaction between H_2SO_4 and KOH. What is the name of the salt that is formed?

Solution

The general reaction is as follows:

$$H_2SO_4 + KOH \rightarrow H_2O + \text{salt}$$

Because the acid has two H^+ ions in its formula, we need two OH^- ions to react with it, making two H_2O molecules as product. The remaining ions, K^+ and SO_4^{2-}, make the salt potassium sulfate (K_2SO_4). The balanced chemical reaction is as follows:

$$H_2SO_4 + 2KOH \rightarrow 2H_2O + K_2SO_4$$

Exercise 7.3.2

Write the balanced chemical equation for the neutralization reaction between HCl and $Mg(OH)_2$. What is the name of the salt that is formed?

Answer available at the end of this section.

Key Takeaways

- An Arrhenius acid is a compound that increases the H^+ ion concentration in aqueous solution.
- An Arrhenius base is a compound that increases the OH^- ion concentration in aqueous solution.
- The reaction between an Arrhenius acid and an Arrhenius base is called neutralization and results in the formation of water and a salt.

Learning Objectives

- Identify a Brønsted-Lowry acid and a Brønsted-Lowry base.
- Identify conjugate acid-base pairs in an acid-base reaction.

The Arrhenius definition of acid and base is limited to aqueous (that is, water) solutions. Although this is useful because water is a common solvent, it is limited to the relationship between the H^+ ion and the OH^- ion. What would be useful is a general definition more applicable to other chemical reactions and, importantly, independent of H_2O.

In 1923, Danish chemist Johannes Brønsted and English chemist Thomas Lowry independently proposed new definitions for acids and bases, ones that focus on proton transfer. A **Brønsted-Lowry acid** is any species that can donate a proton (H^+) to another molecule. A **Brønsted-Lowry base** is any species that can accept a proton from another molecule. In short, a Brønsted-Lowry acid is a proton donor (PD), while a Brønsted-Lowry base is a proton acceptor (PA).

It is easy to see that the Brønsted-Lowry definition covers the Arrhenius definition of acids and bases. Consider the prototypical Arrhenius acid-base reaction:

$$\underset{acid}{H^+(aq)} + \underset{base}{OH^-(aq)} \rightarrow H_2O\,(l)$$

Acid species and base species are marked. The proton, however, is (by definition) a proton donor (labeled PD), while the OH^- ion is acting as the proton acceptor (labeled PA):

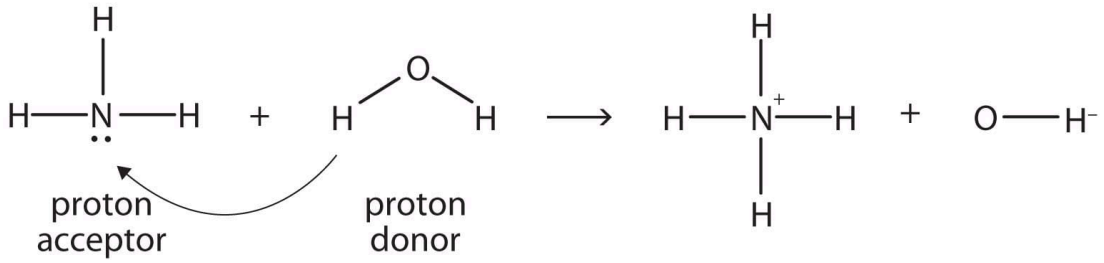
$$\underset{PD}{H^+(aq)} + \underset{PA}{OH^-(aq)} \rightarrow H_2O\,(l)$$

The proton donor is a Brønsted-Lowry acid, and the proton acceptor is the Brønsted-Lowry base:

$$\underset{BL\ acid}{H^+(aq)} + \underset{BL\ base}{OH^-(aq)} \rightarrow H_2O\,(l)$$

Thus H^+ is an acid by both definitions, and OH^- is a base by both definitions.

Ammonia (NH_3) is a base even though it does not contain OH^- ions in its formula. Instead, it generates OH^- ions as the product of a proton-transfer reaction with H_2O molecules; NH_3 acts like a Brønsted-Lowry base, and H_2O acts like a Brønsted-Lowry acid:



A reaction with water is called **hydrolysis**; we say that NH_3 hydrolyzes to make NH_4^+ ions and OH^- ions.

Even the dissolving of an Arrhenius acid in water can be considered a Brønsted-Lowry acid-base reaction. Consider the process of dissolving HCl(g) in water to make an aqueous solution of hydrochloric acid. The process can be written as follows:

$$\mathrm{HCl\,(g) + H_2O(\ell) \longrightarrow H_3O^+\,(aq) + Cl^-\,(aq)}$$

HCl(g) is the proton donor and therefore a Brønsted-Lowry acid, while H_2O is the proton acceptor and a Brønsted-Lowry base. These two examples show that H_2O can act as both a proton donor and a proton acceptor, depending on what other substance is in the chemical reaction. A substance that can act as a proton donor or a proton acceptor is called **amphiprotic**. Water is

probably the most common amphiprotic substance we will encounter, but other substances are also amphiprotic.

Example 7.3.1

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

$$C_6H_5OH + NH_2^- \longrightarrow C_6H_5O^- + NH_3$$

Solution

The C_6H_5OH molecule is losing an H^+; it is the proton donor and the Brønsted-Lowry acid. The NH_2^- ion (called the amide ion) is accepting the H^+ ion to become NH_3, so it is the Brønsted-Lowry base.

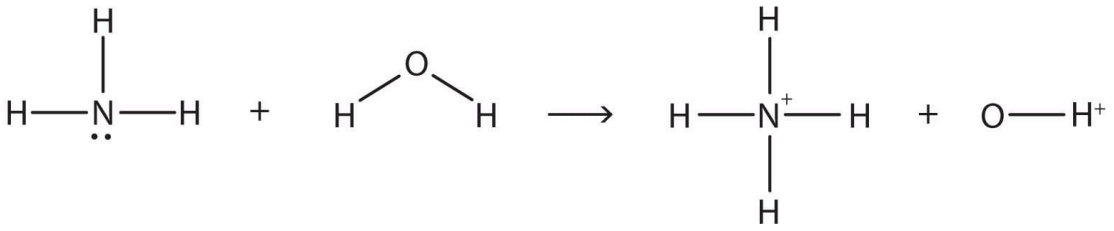
Exercise 7.3.1: Aluminum Ions in Solution

Identify the Brønsted-Lowry acid and the Brønsted-Lowry base in this chemical equation.

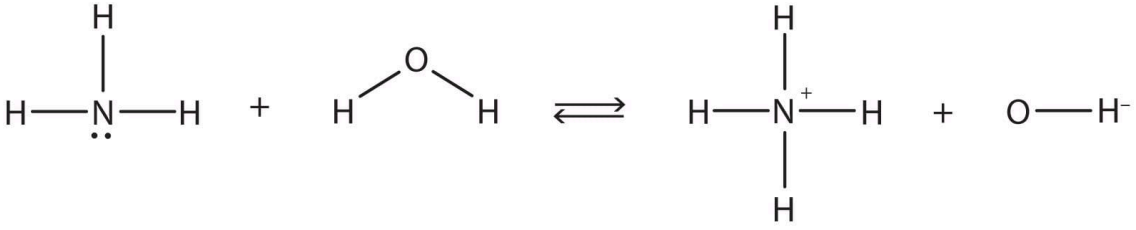
$$Al(H_2O)_6^{3+} + H_2O \longrightarrow Al(H_2O)_5(OH)^{2+} + H_3O^+$$

Answer available at the end of this section.

In the reaction between NH_3 and H_2O,



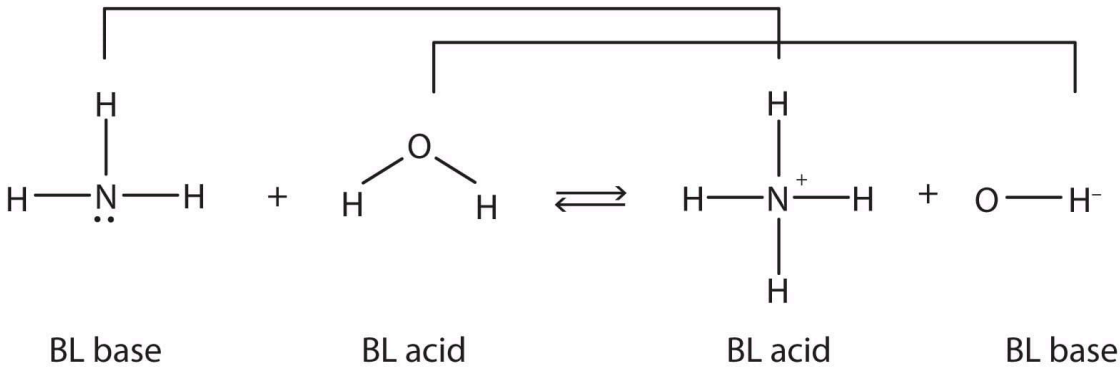
the chemical reaction does not go to completion; rather, the reverse process occurs as well, and eventually the two processes cancel out any additional change. At this point, we say the chemical reaction is at *equilibrium*. Both processes still occur, but any net change by one process is countered by the same net change of the other process; it is a *dynamic*, rather than a *static*, equilibrium. Because both reactions are occurring, it makes sense to use a double arrow instead of a single arrow:



What do you notice about the reverse reaction? The NH_4^+ ion is donating a proton to the OH^- ion, which is accepting it. This means that the NH_4^+ ion is acting as the proton donor, or Brønsted-Lowry acid, while the OH^- ion, the proton acceptor, is acting as a Brønsted-Lowry base. The reverse reaction is also a Brønsted-Lowry acid base reaction:

BL bases. NH4+ and H2O are the BL acids. lt-chem-64081" style="width: 750px; height: 173px;" width="750px" height="173px" src="/@api/deki/files/92035/3765cabac9591a0fb3dd5878f56075e2.jpg" data-quail-id="174">

This means that both reactions are acid-base reactions by the Brønsted-Lowry definition. If you consider the species in this chemical reaction, two sets of similar species exist on both sides. Within each set, the two species differ by a proton in their formulas, and one member of the set is a Brønsted-Lowry acid, while the other member is a Brønsted-Lowry base. These sets are marked here:



The two sets—NH_3/NH_4^+ and H_2O/OH^-—are called **conjugate acid-base pairs**. We say that NH_4^+ is the conjugate acid of NH_3, OH^- is the conjugate base of H_2O, and so forth. Every Brønsted-Lowry acid-base reaction can be labeled with two conjugate acid-base pairs.

Example 7.3.2

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H_2O and OH^-, where H_2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base. The other pair consists of $(CH_3)_3N$ and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and $(CH_3)_3N$ is the conjugate base.

Exercise 7.3.2

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

Answer available at the end of this section.

Chemistry is Everywhere: Household Acids and Bases

Many household products are acids or bases. For example, the owner of a swimming pool may use muriatic acid to clean the pool. Muriatic acid is another name for HCl(aq). In Section 4.6, vinegar was mentioned as a dilute solution of acetic acid [$HC_2H_3O_2$(aq)]. In a medicine chest, one may find a bottle of vitamin C tablets; the chemical name of vitamin C is ascorbic acid ($HC_6H_7O_6$).

One of the more familiar household bases is NH_3, which is found in numerous cleaning products. NH_3 is a base because it increases the OH^- ion concentration by reacting with H_2O:

$$NH_3(aq) + H_2O(\ell) \rightarrow NH_4^+(aq) + OH^-(aq)$$

Many soaps are also slightly basic because they contain compounds that act as Brønsted-Lowry bases, accepting protons from H_2O and forming excess OH^- ions. This is one explanation for why soap solutions are slippery.

Perhaps the most dangerous household chemical is the lye-based drain cleaner. Lye is a common name for NaOH, although it is also used as a synonym for KOH. Lye is an extremely caustic chemical that can react with grease, hair, food particles, and other substances that may build up and clog a water pipe. Unfortunately, lye can also attack body tissues and other substances in our bodies. Thus when we use lye-based drain cleaners, we must be very careful not to touch any of the solid drain cleaner or spill the water it was poured into. Safer, non-lye drain cleaners (like the one in the accompanying figure) use peroxide compounds to react on the materials in the clog and clear the drain.



Figure 7.3.1 Drain Cleaners. Drain cleaners can be made from a reactive material that is less caustic than a base. Source: Photo used by permission of Citrasolv, LLC.

Key Takeaways

- A Brønsted-Lowry acid is a proton donor; a Brønsted-Lowry base is a proton acceptor.
- Acid-base reactions include two sets of conjugate acid-base pairs.

Learning Objective

- Know the many uses of common acids and bases.

Antacids: A Basic Remedy

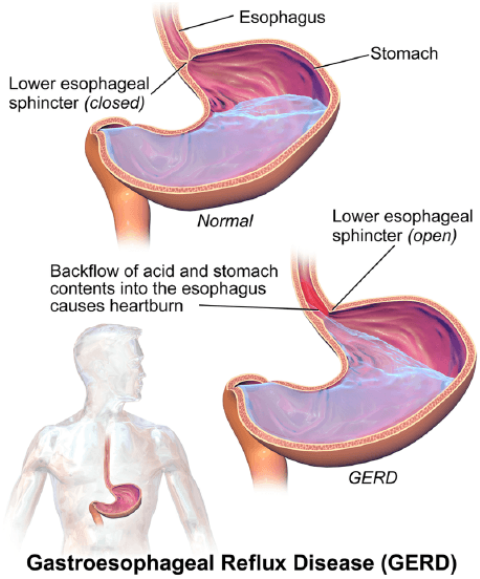


Figure 7.3.1. Antacids are available over the counter and are taken by mouth to quickly relieve occasional heartburn, the major symptom of gastroesophageal reflux disease and also indigestion. Treatment with antacids alone is symptomatic and only justified for minor symptoms.

Antacids contain alkaline ions that chemically neutralize stomach gastric acid, reducing damage and relieving pain. For example, Tums is a very commonly used antacid that can be bought over the counter. Other antacids, such as Alka Seltzer, work in similar ways. Below, show the mechanism and products for reaction in which calcium carbonate (the active ingredient in Tums) neutralizes HCl in a simple proton transfer reaction.

$$CaCO_3\,(s) + HCl\,(aq) \rightarrow H_2CO_3\,(aq) + CaCl_2\,(aq)$$

Following this acid-base reaction, carbonic acid quickly degrades into CO_2 and H_2O.

$$H_2CO_3\,(aq) \rightarrow CO_2(g) + H_2O(l)$$

Figure 7.3.1 *A comparison of a healthy condition to GERD. I*

mage used with permission (Cc BY-SA 4.0; BruceBlaus .

Under the generic name algeldrate, aluminium hydroxide is used as an antacid. Aluminium hydroxide is preferred over other alternatives such as sodium bicarbonate because $Al(OH)_3$, being insoluble, does not increase the pH of stomach above 7 and hence, does not trigger secretion of excess stomach acid. Brand names include Alu-Cap, Aludrox, Gaviscon, and Pepsamar. In 2016 Gaviscon was one of the biggest selling branded over-the-counter medications sold in Great Britain, with sales of £62 million. It reacts with excess acid in the stomach, reducing the acidity of the stomach content, which may relieve the symptoms of ulcers, heartburn or dyspepsia. Such products can cause constipation, because the aluminum ions inhibit the contractions of smooth muscle cells in the gastrointestinal tract, slowing peristalsis and lengthening the time needed for stool to pass through the colon. Some such products are formulated to minimize such effects through the inclusion of equal concentrations of magnesium hydroxide or magnesium carbonate, which have counterbalancing laxative effects.

Antacids are distinct from acid-reducing drugs like H_2-receptor antagonists or proton pump inhibitors and they do not kill the bacteria *Helicobacter pylori*, which causes most ulcers. Ranitidine (Zantac), famotidine (Pepcid AC), and cimetidine (Tagamet) are in a class of medications called H_2 blockers. It decreases the amount of acid made in the stomach and are used to treat ulcers; gastroesophageal reflux disease (GERD), a condition in which backward flow of acid from the stomach causes heartburn and injury of the food pipe (esophagus); and conditions where the stomach produces too much acid, such as Zollinger-Ellison syndrome. These over-the-counter medications are used to prevent and treat symptoms of heartburn associated with acid indigestion and sour stomach. Omeprazole (Prilosec) also works by decreasing the amount of acid made in the stomach and is in a class of medications called proton-pump inhibitors.

Acids and Bases in Industry and at Home

Common Acids

It should not be hard for you to name several common acids, but you might find that listing bases is just a little more difficult. Here's a partial list of some common acids, along with some chemical formulas:

Table 7.3.1 *Common Acids and Their Uses*

Chemist Name	Common Name	Uses
hydrochloric acid, HCl	muriatic acid (used in pools) and stomach acid	Used in cleaning (refining) metals, in maintenance of swimming pools, and for household cleaning.
sulfuric acid, H_2SO_4		Used in car batteries, and in the manufacture of fertilizers.
nitric acid, HNO_3		Used in the manufacture of fertilizers, explosives and in extraction of gold.
acetic acid, $HC_2H_3O_2$	vinegar	Main ingredient in vinegar.
carbonic acid, H_2CO_3	responsible for the "fiz" in carbonated drinks	As an ingredient in carbonated drinks.
citric acid, $C_6H_8O_7$		Used in food and dietary supplements. Also added as an acidulant in creams, gels, liquids, and lotions.
acetylsalicylic acid, $C_6H_4(OCOCH_3)CO_2H$	aspirin	The active ingredient in aspirin.

Hydrochloric Acid

Hydrochloric acid is a corrosive, strong mineral acid with many industrial uses. One of the most important applications of hydrochloric acid is in the pickling of steel, to remove rust or iron oxide scale from iron or steel before subsequent processing. Another major use of hydrochloric acid is in the production of organic compounds, such as vinyl chloride and dichloroethane for PVC. Other organic compounds produced with hydrochloric acid include bisphenol A for polycarbonate, activated carbon, and ascorbic acid, as well as numerous pharmaceutical products. Other inorganic compounds produced with hydrochloric acid include road application salt calcium chloride, nickel(II) chloride for electroplating, and zinc chloride for the galvanizing industry and battery production.

Sulfuric Acid

Sulfuric acid is a very important commodity chemical, and indeed, a nation's sulfuric acid production is a good indicator of its industrial strength. World production in the year 2004 was about 180 million tonnes, with the following geographic distribution: Asia 35%, North America (including Mexico) 24%, Africa 11%, Western Europe 10%, Eastern Europe and Russia 10%, Australia and Oceania 7%, South America 7%. Most of this amount (≈60%) is consumed for fertilizers, particularly superphosphates, ammonium phosphate and ammonium sulfates. About 20% is used in chemical industry for production of detergents, synthetic resins, dyestuffs, pharmaceuticals, petroleum catalysts, insecticides and antifreeze, as well as in various processes such as oil well acidicizing, aluminium reduction, paper sizing, water treatment. About 6% of uses are related to pigments and include paints, enamels, printing inks, coated fabrics and paper, and the rest is dispersed into a multitude of applications such as production of explosives, cellophane, acetate and viscose textiles, lubricants, non-ferrous metals, and batteries.



Figure 7.3.2: Drops of concentrated sulfuric acid rapidly decompose a piece of cotton towel by dehydration. (CC BY-SA 3.0; Toxic Walker).

Because the hydration of sulfuric acid is thermodynamically favorable (and is highly exothermic) and the affinity of it for water is sufficiently strong, sulfuric acid is an excellent dehydrating agent. Concentrated sulfuric acid has a very powerful dehydrating property, removing water (H_2O) from other compounds including sugar and other carbohydrates and producing carbon, heat, steam. Sulfuric acid behaves as a typical acid in its reaction with most metals by generating hydrogen gas (Equation 7.3.2).

Nitric Acid

Nitric acid (HNO_3) is a highly corrosive mineral acid and is also commonly used as a strong oxidizing agent. Nitric acid is normally considered to be a strong acid at ambient temperatures. The main industrial use of nitric acid is for the production of fertilizers. Nitric acid is neutralized with ammonia to give ammonium nitrate. This application consumes 75–80% of the 26 million tonnes produced annually (1987). The other main applications are for the production of explosives, nylon precursors, and specialty organic compounds.



Figure 7.3.3: Second degree burn caused by nitric acid. (CC BY-SA 3.0; Alcamán).

Table 7.3.2 Common Bases and Its Uses

Some Common Bases	Uses
sodium hydroxide, NaOH (lye or caustic soda)	Used in the manufacture of soaps and detergents and as the main ingredient in oven and drain cleaners.
potassium hydroxide, KOH (lye or caustic potash)	Used in the production of liquid soaps and soft soaps. Used in alkaline batteries.
magnesium hydroxide, $Mg(OH)_2$ (milk of magnesia)	Used as an ingredient in laxatives, antacids, and deodorants. Also used in the neutralization of acidic wastewater.
calcium hydroxide, $Ca(OH)_2$ (slaked lime)	Used in the manufacture of cement and lime water. Also, added to neutralize acidic soil.
aluminum hydroxide	Used in water purification and as an ingredient in antacids.
ammonia, NH_3	Used as a building block for the synthesis of many pharmaceutical products and in many commercial cleaning products. Used in the manufacture of fertilizers.

Sodium Hydroxide

Sodium hydroxide, also known as lye and caustic soda, is an inorganic compound with formula NaOH.

It is a popular strong base used in industry. Around 56% of sodium hydroxide produced is used by industry, 25% of which is used in the paper industry. Sodium hydroxide is also used in the manufacture of sodium salts and detergents, pH regulation, and organic synthesis. It is used in the Bayer process of aluminium production. In bulk, it is most often handled as an aqueous solution, since solutions are cheaper and easier to handle.

Sodium hydroxide is used in many scenarios where it is desirable to increase the alkalinity of a mixture, or to neutralize acids.



Figure 7.3.4 Sample of sodium hydroxide as pellets in a watchglass. (Public Domain; Walkerma).

Calcium Hydroxide

Calcium hydroxide (traditionally called slaked lime) is an inorganic compound with the chemical formula $\mathrm{Ca(OH)_2}$. It is a colorless crystal or white powder. It has many names including hydrated lime, caustic lime, builders' lime, slaked lime, cal, or pickling lime.

One significant application of calcium hydroxide is as a flocculant, in water and sewage treatment. It is also used in fresh water treatment for raising the pH of the water so that pipes will not corrode where the base water is acidic, because it is self-regulating and does not raise the pH too much.

It is also used in the preparation of ammonia gas (NH_3), using the following reaction:

$$Ca(OH)_2 + 2NH_4Cl \rightarrow 2NH_3 + CaCl_2 + 2H_2O$$

Another large application is in the paper industry, where it is an intermediate in the reaction in the production of sodium hydroxide. Because of its low toxicity and the mildness of its basic properties, slaked lime is widely used in the food industry:

Ammonia

Ammonia is a compound of nitrogen and hydrogen with the formula $\mathrm{NH_3}$ and is a colorless gas with a characteristic pungent smell. It is the active product of “smelling salts,” and can quickly revive the faint of heart and light of head. Although common in nature and in wide use, ammonia is both caustic and hazardous in its concentrated form.

Globally, approximately 88% (as of 2014) of ammonia is used as fertilizers either as its salts, solutions or anhydrously. When applied to soil, it helps provide increased yields of crops such as maize and wheat. 30% of agricultural nitrogen applied in the US is in the form of anhydrous ammonia and worldwide 110 million tonnes are applied each year. Ammonia is also a building block for the synthesis of many pharmaceutical products and is used in many commercial cleaning products.

Summary

Common acids and bases have many and various uses in industry and in daily life.

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
- Wikipedia
- TextMap: Chemistry-A Molecular Approach (Tro)
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)
- MedlinePlus

Learning Objectives

- Describe the autoionization of water.
- Calculate the concentrations of H^+ and OH^- in solutions, knowing the other concentration.

We have already seen that H_2O can act as an acid or a base:

$$NH_3 + H_2O \rightarrow NH_4^+ + OH^- \text{ (H_2O acts as an acid)}$$

$$HCl + H_2O \rightarrow H_3O^+ + Cl^- \text{ (H_2O acts as a base)}$$

It may not be surprising to learn, then, that within any given sample of water, some H_2O molecules are acting as acids, and other H_2O molecules are acting as bases. The chemical equation is as follows:

$$H_2O + H_2O \rightarrow H_3O^+ + OH^-$$

This occurs only to a very small degree: only about 6 in 10^8 H_2O molecules are participating in this process, which is called the **autoionization of water**. At this level, the concentration of both $H^+(aq)$ and $OH^-(aq)$ in a sample of pure H_2O is about 1.0×10^{-7} M. If we use square brackets—[ ]—around a dissolved species to imply the molar concentration of that species, we have

$$[H^+] = [OH^-] = 1.0 \times 10^{-7} \text{ M}$$

for *any* sample of pure water because H_2O can act as both an acid and a base. The product of these two concentrations is 1.0×10^{-14}:

$$[H^+] \times [OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

In acids, the concentration of $H^+(aq)$—$[H^+]$—is greater than 1.0×10^{-7} M, while for bases the concentration of $OH^-(aq)$—$[OH^-]$—is greater than 1.0×10^{-7} M. However, the *product* of the two concentrations—$[H^+][OH^-]$—is *always* equal to 1.0×10^{-14}, no matter whether the aqueous solution is an acid, a base, or neutral:

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$

This value of the product of concentrations is so important for aqueous solutions that it is called the **autoionization constant of water** and is denoted K_W:

$$K_W = [H^+][OH^-] = 1.0 \times 10^{-14}$$

This means that if you know $[H^+]$ for a solution, you can calculate what $[OH^-]$ has to be for the product to equal 1.0×10^{-14},

or if you know $[OH^-]$, you can calculate $[H^+]$. This also implies that as one concentration goes up, the other must go down to compensate so that their product always equals the value of K_W.

Example 7.3.1

What is $[OH^-]$ of an aqueous solution if $[H^+]$ is 1.0×10^{-4} M?

Solution

Using the expression and known value for K_W,

$$K_W = [H^+][OH^-] = 1.0 \times 10^{-14} = (1.0 \times 10^{-4})[OH^-]$$

We solve by dividing both sides of the equation by 1.0×10^{-4}:

$$[OH^-] = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} M$$

It is assumed that the concentration unit is molarity, so $[OH^-]$ is 1.0×10^{-10} M.

Exercise 7.3.1

What is $[H^+]$ of an aqueous solution if $[OH^-]$ is 1.0×10^{-9} M?

Answer available at the end of this section.

When you have a solution of a particular acid or base, you need to look at the formula of the acid or base to determine the number of H^+ or OH^- ions in the formula unit because $[H^+]$ or $[OH^-]$ may not be the same as the concentration of the acid or base itself.

Example 7.3.2

What is $[H^+]$ in a 0.0044 M solution of $Ca(OH)_2$?

Solution

We begin by determining $[OH^-]$. The concentration of the solute is 0.0044 M, but because $Ca(OH)_2$ is a strong base, there are two OH^- ions in solution for every formula unit dissolved, so the actual $[OH^-]$ is two times this, or 2×0.0044 M = 0.0088 M. Now we can use the K_W expression:

$$[H^+][OH^-] = 1.0 \times 10^{-14} = [H^+](0.0088 \text{ M})$$

Divide both sides by 0.0088:

$$[H^+] = \frac{1.0 \times 10^{-14}}{(0.0088)} = 1.1 \times 10^{-12} M$$

$[H^+]$ has decreased significantly in this basic solution.

Exercise 7.3.2

What is [OH^-] in a 0.00032 M solution of H_2SO_4? (Hint: assume both H^+ ions ionize.)

Answer

1.6×10^{-11} M

For strong acids and bases, [H^+] and [OH^-] can be determined directly from the concentration of the acid or base itself because these ions are 100% ionized by definition. However, for weak acids and bases, this is not so. The degree, or percentage, of ionization would need to be known before we can determine [H^+] and [OH^-].

Example 7.3.3

A 0.0788 M solution of $HC_2H_3O_2$ is 3.0% ionized into H^+ ions and $C_2H_3O_2^-$ ions. What are [H^+] and [OH^-] for this solution?

Solution

Because the acid is only 3.0% ionized, we can determine [H^+] from the concentration of the acid. Recall that 3.0% is 0.030 in decimal form:

[H^+] = 0.030 × 0.0788 = 0.00236 M

With this [H^+], then [OH^-] can be calculated as follows:

$$[OH^-] = \frac{1.0 \times 10^{-14}}{0.00236} = 4.2 \times 10^{-12} M$$

This is about 30 times higher than would be expected for a strong acid of the same concentration.

Exercise 7.3.3

A 0.0222 M solution of pyridine (C_5H_5N) is 0.44% ionized into pyridinium ions ($C_5H_5NH^+$) and OH^- ions. What are [OH^-] and [H^+] for this solution?

Answer available at the end of this section.

Summary

In any aqueous solution, the product of [H+] and [OH–] equals 1.0×10^{-14} (at room temperature).

Learning Objectives

- Define pH.
- Determine the pH of acidic and basic solutions.
- Determine the hydrogen (hydronium) ion concentration from pH and vice versa.

Molar concentration values of hydrogen $[H^+]$ can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions. This is known as the pH scale. The range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to make a quick determination whether a given aqueous solution is acidic, basic, or neutral.

pH is a *logarithmic* scale. A solution that has a pH of 1.0 has 10 times the [H^+] as a solution with a pH of 2.0, which in turn has 10 times the [H^+] as a solution with a pH of 3.0 and so forth.

Warning: The pH scale has no limits

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on $[H+]$, we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

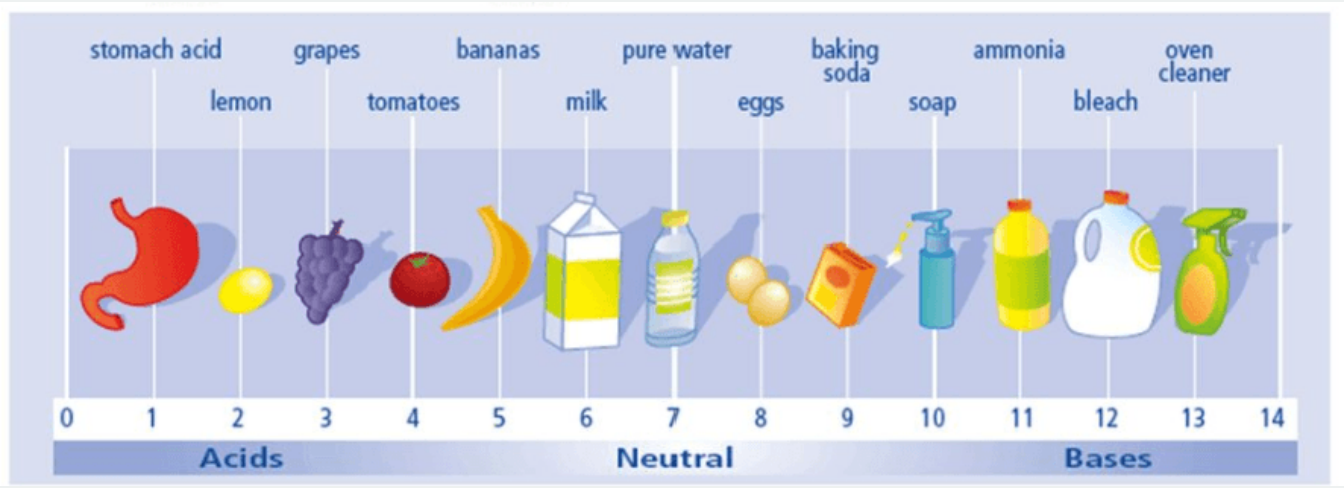


Figure 7.3.1 The pH scale.

Figure 7.3.2 illustrates the relationship between pH and the hydrogen ion concentration, along with some examples of various solutions. Because hydrogen ion concentrations are generally less than one (for example $1.3 \times 10^{-3}\ M$), the log of the number will be a negative number. To make pH even easier to work with, pH is defined as the ***negative* log of** $[H^+]$, which will give a positive value for pH.

The general formula for determining [H^+] from pH is as follows:

$$[H^+] = 10^{-pH}$$

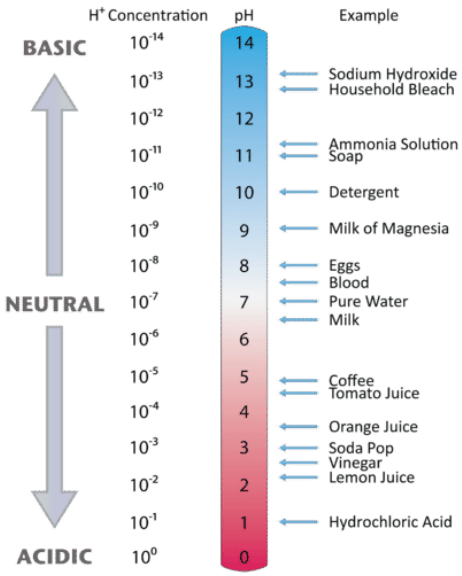


Figure 7.3.2 The pH values for several common materials with corresponding hydrogen ion concentration, [H+] .

Example 7.3.1

Label each solution as acidic, basic, or neutral based only on the stated pH.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Answer

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

Exercise 7.3.1

Identify each substance as acidic, basic, or neutral based only on the stated pH.

a. human blood with pH = 7.4
b. household ammonia with pH = 11.0
c. cherries with pH = 3.6

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Table 7.3.1: gives the typical pH values of some common substances. Note that several food items are on the list, and most of them are acidic.

*Table 7.3.1 Typical pH Values of Various Substances**

Substance	pH
stomach acid	1.7
lemon juice	2.2
vinegar	2.9
soda	3.0
wine	3.5
coffee, black	5.0
milk	6.9
pure water	7.0
blood	7.4
seawater	8.5
milk of magnesia	10.5
ammonia solution	12.5
1.0 M NaOH	14.0
***Actual values may vary depending on conditions**	

Example 7.3.2:

What is the $[H^+]$ for an aqueous solution whose pH is 6?

Solution

The pH value of 6 denotes that the exponent of 10 is -6. Therefore the answer is

$$[H^+] = 1.0 \times 10^{-6} M$$

Exercise 7.3.2

What is the $[H^+]$ for an aqueous solution whose pH is 11?

Answer

$[H^+] = 1.0 \times 10^{-11}$ M

Example 7.3.3:

What is the pH of an aqueous solution whose hydrogen ion concentration is 1.0×10^{-7}M? Is the solution acidic, basic, or neutral.

Solution

The hydrogen ion concentration is 1.0 x 10^{-7}M. The exponent of 10 is -7, which denoted that pH= -(-7) = **7**. The solution is **neutral**.

Exercise 7.3.3

What is the pH of an aqueous solution whose hydrogen ion concentration is 1.0 x 10^{-3}M? Is the solution acidic, basic, or neutral.

Solution

The hydrogen ion concentration is 1.0 x 10^{-3}M. The exponent of 10 is -3, which denoted that pH= -(-3) = **3**. The solution is **acidic**.

Summary

To make pH even easier to work with, pH is defined as the ***negative*** **log of** $[H^+]$, which will give a positive value for pH.

pH is usually (but not always) between 0 and 14.

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

Contributors and Attributions

- Peggy Lawson (Oxbow Prairie Heights School). Funded by Saskatchewan Educational Technology Consortium.
-
- Template:OpenStax
- Marisa Alviar-Agnew (Sacramento City College)
- Henry Agnew (UC Davis)

Learning Objectives

- Define *pH*.
- Determine the pH of acidic and basic solutions.

As we have seen, [H^+] and [OH^-] values can be markedly different from one aqueous solution to another. So chemists defined a new scale that succinctly indicates the concentrations of either of these two ions.

pH is commonly defined as a logarithmic function of [H^+]:

$$\mathrm{pH} = -\log[\mathrm{H}^+] \tag{7.3.4}$$

pH is usually (but not always) between 0 and 14. Knowing the dependence of pH on [H^+], we can summarize as follows:

- If pH < 7, then the solution is acidic.
- If pH = 7, then the solution is neutral.
- If pH > 7, then the solution is basic.

This is known as the pH scale and is the range of values from 0 to 14 that describes the acidity or basicity of a solution. You can use pH to quickly determine whether a given aqueous solution is acidic, basic, or neutral.

Example 7.3.1

Label each solution as acidic, basic, or neutral based only on the stated pH.

a. milk of magnesia, pH = 10.5
b. pure water, pH = 7
c. wine, pH = 3.0

Solution

a. With a pH greater than 7, milk of magnesia is basic. (Milk of magnesia is largely $Mg(OH)_2$.)
b. Pure water, with a pH of 7, is neutral.
c. With a pH of less than 7, wine is acidic.

Exercise 7.3.1

Identify each substance as acidic, basic, or neutral based only on the stated pH.

1. human blood, pH = 7.4
2. household ammonia, pH = 11.0
3. cherries, pH = 3.6

Answers

1. basic
2. basic
3. acidic

Table 7.3.1 gives the typical pH values of some common substances. Note that several food items are on the list, and most of them are acidic.

*Table 7.3.1 Typical pH Values of Various Substances**

Substance	pH
stomach acid	1.7
lemon juice	2.2
vinegar	2.9
soda	3.0
wine	3.5
coffee, black	5.0
milk	6.9
pure water	7.0
blood	7.4
seawater	8.5
milk of magnesia	10.5
***Actual values may vary depending on conditions**	

Substance	pH
ammonia solution	12.5
1.0 M NaOH	14.0
***Actual values may vary depending on conditions**	

pH is a *logarithmic* scale. A solution that has a pH of 1.0 has 10 times the $[H^+]$ as a solution with a pH of 2.0, which in turn has 10 times the $[H^+]$ as a solution with a pH of 3.0 and so forth.

Using the definition of pH (Equation 7.3.4), it is also possible to calculate $[H^+]$ (and $[OH^-]$) from pH and vice versa. The general formula for determining $[H^+]$ from pH is as follows:

$$[H^+] = 10^{-pH}$$

Key Takeaways

- pH is a logarithmic function of $[H^+]$.
- $[H^+]$ can be calculated directly from pH.
- pOH is related to pH and can be easily calculated from pH.

Learning Objectives

- Define a strong and a weak acid and base.
- Recognize an acid or a base as strong or weak.

Strong and Weak Acids

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider HCl(aq). When HCl is dissolved in H_2O, it completely dissociates into H^+(aq) and Cl^-(aq) ions; all the HCl molecules become ions:

$$HCl \overset{100\%}{\rightarrow} H^+(aq) + Cl^-(aq)$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. $HC_2H_3O_2$ is an example of a weak acid:

$$HC_2H_3O_2 \overset{\sim 5\%}{\rightarrow} H^+(aq) + C_2H_3O_2^-(aq)$$

Because this reaction does not go 100% to completion, it is more appropriate to write it as **a reversible reaction**:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

As it turns out, there are very few strong acids, which are given in Table 7.3.1. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

> Any acid that dissociates 100% into ions is called a strong acid. If it does not dissociate 100%, it is a weak acid.

Table 7.3.1 Strong Acids and Bases

Acids	Bases
HCl	LiOH

Acids	Bases
HBr	NaOH
HI	KOH
HNO_3	RbOH
H_2SO_4	CsOH
$HClO_3$	$Mg(OH)_2$
$HClO_4$	$Ca(OH)_2$
	$Sr(OH)_2$
	$Ba(OH)_2$

Strong and Weak Bases

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it is a **weak base**. There are very few strong bases (Table 7.3.1); any base not listed is a weak base. All strong bases are OH^- compounds. So a base based on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base.

Example 7.3.1: Identifying Strong and Weak Acids and Bases

Identify each acid or base as strong or weak.

a. HCl
b. $Mg(OH)_2$
c. C_5H_5N

Solution

a. Because HCl is listed in Table 7.3.1, it is a strong acid.
b. Because $Mg(OH)_2$ is listed in Table 7.3.1, it is a strong base.
c. The nitrogen in C_5H_5N would act as a proton acceptor and therefore can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

Exercise 7.3.1

Identify each acid or base as strong or weak.

a. RbOH
b. $\mathrm{HNO_2}$

Answer a available at the end of this section.

Example 7.3.2: Characterizing Base Ionization

Write the balanced chemical equation for the dissociation of $Ca(OH)_2$ and indicate whether it proceeds 100% to products or not.

Solution

This is an ionic compound of Ca^{2+} ions and OH^- ions. When an ionic compound dissolves, it separates into its constituent ions:

$$Ca(OH)_2 \longrightarrow Ca^{2+}(aq) + 2\,OH^-(aq)$$

Because Ca(OH)2 is listed in Table 7.3.1, this reaction proceeds 100% to products.

Exercise 7.3.2

Write the balanced chemical equation for the dissociation of hydrazoic acid (HN3) and indicate whether it proceeds 100% to products or not.

Answer a available at the end of this section.

Summary

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.

Contributors and Attributions

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Learning Objectives

- Define a strong and a weak acid and base.
- Recognize an acid or a base as strong or weak.
- Determine if a salt produces an acidic or a basic solution.

Except for their names and formulas, so far we have treated all acids as equals, especially in a chemical reaction. However, acids can be very different in a very important way. Consider $HCl(aq)$. When HCl is dissolved in H_2O, it completely dissociates into $H^+(aq)$ and $Cl^-(aq)$ ions; all the HCl molecules become ions:

$$HCl \xrightarrow{100\,\%} H^+(aq) + Cl^-(aq)$$

Any acid that dissociates 100% into ions is called a **strong acid**. If it does not dissociate 100%, it is a **weak acid**. Acetic acid ($HC_2H_3O_2$) is an example of a weak acid:

$$HC_2H_3O_2 \xrightarrow{\sim 5\,\%} H^+(aq) + C_2H_3O_2{}^-(aq)$$

Because this reaction does not go 100% to completion, it is more appropriate to write it as a reversible reaction:

$$HC_2H_3O_2 \rightleftharpoons H^+(aq) + C_2H_3O_2{}^-(aq)$$

As it turns out, there are very few strong acids, which are given in Table 7.3.1. If an acid is not listed here, it is a weak acid. It may be 1% ionized or 99% ionized, but it is still classified as a weak acid.

Table 7.3.1: Strong Acids and Bases

Acids	Bases
HCl	LiOH
HBr	NaOH
HI	KOH
HNO_3	RbOH
H_2SO_4	CsOH
$HClO_3$	$Mg(OH)_2$
$HClO_4$	$Ca(OH)_2$
	$Sr(OH)_2$
	$Ba(OH)_2$

The issue is similar with bases: a **strong base** is a base that is 100% ionized in solution. If it is less than 100% ionized in solution, it is a **weak base**. There are very few strong bases (Table 7.3.1); any base not listed is a weak base. All strong bases are OH^- compounds. So a base dependent on some other mechanism, such as NH_3 (which does not contain OH^- ions as part of its formula), will be a weak base.

Example 7.3.1

Identify each acid or base as strong or weak.

a. HCl
b. $Mg(OH)_2$
c. C_5H_5N

Solution

a. Because HCl is listed in Table 7.3.1, it is a strong acid.
b. Because $Mg(OH)_2$ is listed in Table 7.3.1, it is a strong base.
c. The nitrogen in C_5H_5N would act as a proton acceptor and therefore can be considered a base, but because it does not contain an OH compound, it cannot be considered a strong base; it is a weak base.

Exercise 7.3.1

Identify each acid or base as strong or weak.

a. RbOH
b. HNO_2

Answer a available at the end of this section.

Example 7.3.2

Write the balanced chemical equation for the dissociation of $Ca(OH)_2$ and indicate whether it proceeds 100% to products or not.

Solution

This is an ionic compound of Ca^{2+} ions and OH^- ions. When an ionic compound dissolves, it separates into its constituent ions:

$$Ca(OH)_2 \longrightarrow Ca^{2+}(aq) + 2\,OH^-(aq)$$

Because Ca(OH)2 is listed in Table 7.3.1, this reaction proceeds 100% to products.

Exercise 7.3.2

Write the balanced chemical equation for the dissociation of hydrazoic acid (HN3) and indicate whether it proceeds 100% to products or not.

Answer available at the end of this section.

Certain salts will also affect the acidity or basicity of aqueous solutions because some of the ions will undergo hydrolysis, just like NH3 does, to make a basic solution. The general rule is that salts with ions that are part of strong acids or bases will not hydrolyze, while salts with ions that are part of weak acids or bases will hydrolyze.

Consider NaCl. When it dissolves in an aqueous solution, it separates into Na^+ ions and Cl^- ions:

$$NaCl \longrightarrow Na^+(aq) + Cl^-(aq)$$

Will the Na^+(aq) ion hydrolyze? If it does, it will interact with the OH^- ion to make NaOH:

$$Na^+(aq) + H_2O \longrightarrow NaOH + H^+(aq)$$

However, NaOH is a strong base, which means that it is 100% ionized in solution:

$$NaOH \longrightarrow Na^+(aq) + OH^-(aq)$$

The free OH^-(aq) ion reacts with the H^+(aq) ion to remake a water molecule:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O$$

The net result? There is no change, so there is no effect on the acidity or basicity of the solution from the Na^+(aq) ion. What about the Cl^- ion? Will it hydrolyze? If it does, it will take an H^+ ion from a water molecule:

$$Cl^-(aq) + H_2O \longrightarrow HCl + OH^-$$

However, HCl is a strong acid, which means that it is 100% ionized in solution:

$$HCl \longrightarrow H^+(aq) + Cl^-(aq)$$

The free H^+(aq) ion reacts with the OH^-(aq) ion to remake a water molecule:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O$$

The net result? There is no change, so there is no effect on the acidity or basicity of the solution from the Cl^-(aq) ion. Because neither ion in NaCl affects the acidity or basicity of the solution, NaCl is an example of a neutral salt. A neutral salt is any ionic compound that does not affect the acidity or basicity of its aqueous solution.

Things change, however, when we consider a salt like NaC2H3O2. We already know that the Na^+ ion won't affect the acidity of the solution. What about the acetate ion? If it hydrolyzes, it will take an H^+ from a water molecule:

$$C_2H_3O_2^-(aq) + H_2O \rightleftharpoons HC_2H_3O_2 + OH^-(aq)$$

Does this happen? Yes, it does. Why? *Because HC2H3O2 is a weak acid.* Any chance a weak acid has to form, it will (the same with a weak base). As some $C_2H_3O_2^-$ ions hydrolyze with H_2O to make the molecular weak acid, OH^- ions are produced. OH^- ions make solutions basic. Thus $NaC_2H_3O_2$ solutions are slightly basic, so such a salt is called a **basic salt**.

There are also salts whose aqueous solutions are slightly acidic. NH_4Cl is an example. When NH_4Cl is dissolved in H_2O, it separates into NH_4^+ ions and Cl^- ions. We have already seen that the Cl^- ion does not hydrolyze. However, the NH_4^+ ion will:

$$NH_4^+(aq) + H_2O \rightleftharpoons NH_3(aq) + H_3O^+(aq)$$

Recall from Section 12.2, that H_3O^+ ion is the hydronium ion, the more chemically proper way to represent the H^+ ion. This is the classic acid species in solution, so a solution of NH_4^+(aq) ions is slightly acidic. NH_4Cl is an example of an **acid salt**. The molecule NH_3 is a weak base, and it will form when it can, just like a weak acid will form when it can.

So there are two general rules:

1. If an ion derives from a strong acid or base, it will not affect the acidity of the solution.
2. If an ion derives from a weak acid, it will make the solution basic; if an ion derives from a weak base, it will make the solution acidic.

Example 7.3.3

Identify each salt as acidic, basic, or neutral.

a. KCl
b. KNO_2
c. NH_4Br

Solution

a. The ions from KCl derive from a strong acid (HCl) and a strong base (KOH). Therefore, neither ion will affect the acidity of the solution, so KCl is a neutral salt.
b. Although the K^+ ion derives from a strong base (KOH), the NO_2^- ion derives from a weak acid (HNO_2). Therefore the solution will be basic, and KNO_2 is a basic salt.
c. Although the Br^- ions derive from a strong acid (HBr), the NH_4^+ ion derives from a weak base (NH_3), so the solution will be acidic, and NH_4Br is an acidic salt.

Exercise 7.3.3

Identify each salt as acidic, basic, or neutral.

A. $(C_5H_5NH)Cl$
B. Na_2SO_3

Answer a available at the end of this section.

Some salts are composed of ions that come from both weak acids and weak bases. The overall effect on an aqueous solution depends on which ion exerts more influence on the overall acidity. We will not consider such salts here.

Summary

- Strong acids and bases are 100% ionized in aqueous solution.
- Weak acids and bases are less than 100% ionized in aqueous solution.
- Salts of weak acids or bases can affect the acidity or basicity of their aqueous solutions.

Answers

Exercise 7.3.1

Brønsted-Lowry acid: $H_2PO_4^-$; Brønsted-Lowry base: H_2O

Exercise 7.3.2

A Brønsted-Lowry Base is a proton acceptor, which means it will take in an H^+. This eliminates HCl, H_3PO_4, NH_4^+ and $CH_3NH_3^+$ because they are Bronsted-Lowry acids. They all give away protons. In the case of HPO_4^{2-}, consider the following equation:

$$HPO_4^{2-}(aq) + H_2O(l) \rightarrow PO_4^{3-}(aq) + H_3O^+(aq)$$

Here, it is clear that HPO_4^{2-} is the acid since it donates a proton to water to make H_3O^+ and PO_4^{3-}. Now consider the following equation:

$$HPO_4^{2-}(aq) + H_2O(l) \rightarrow H_2PO_4^- + OH^-(aq)$$

In this case, HPO_4^{2-} is the base since it accepts a proton from water to form $H_2PO_4^-$ and OH^-. Thus, HPO_4^{2-} is an acid and base together, making it amphoteric.

Since HPO_4^{2-} is the only compound from the options that can act as a base, the answer is **(b) HPO_4^{2-}.**

Exercise 7.3.2

$2HCl + Mg(OH)_2 \rightarrow 2H_2O + MgCl_2$; magnesium chloride

Exercise 7.3.1**: Aluminum Ions in Solution**

Brønsted-Lowry acid: $Al(H_2O)_6^{3+}$; Brønsted-Lowry base: H_2O

Exercise 7.3.2

H_2O (acid) and OH^- (base); NH_2^- (base) and NH_3 (acid)

Exercise 7.3.1

1.0×10^{-5} M

Exercise 7.3.3

$[OH^-] = 9.77 \times 10^{-5}$ M; $[H^+] = 1.02 \times 10^{-10}$ M

Exercise 7.3.1

basic

Exercise 7.3.1

basic

Exercise 7.3.1

acidic

Exercise 7.3.1

strong base

weak acid

Exercise 7.3.2

The reaction is as follows:

$$HN_3 \longrightarrow H^+(aq) + N_3^-(aq)$$

It does not proceed 100% to products because hydrazoic acid is not a strong acid.

Exercise 7.3.1

strong base

weak acid

Exercise 7.3.2

The reaction is as follows:

$$HN_3 \longrightarrow H^+(aq) + N_3^-(aq)$$

It does not proceed 100% to products because hydrazoic acid is not a strong acid.

Exercise 7.3.3

acidic

basic

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7.4: Buffers

Learning Objectives

- Identify conjugate acid base pair.
- Define buffers and know the composition of different buffer systems.
- Describe how buffers work.

Conjugate Acid-Base Pair

In reality, all acid-base reactions involve the transfer of protons between acids and bases. For example, consider the acid-base reaction that takes place when ammonia is dissolved in water. A water molecule (functioning as an acid) transfers a proton to an ammonia molecule (functioning as a base), yielding the conjugate base of water, $\mathrm{OH^-}$, and the conjugate acid of ammonia, $\mathrm{NH_4^+}$:

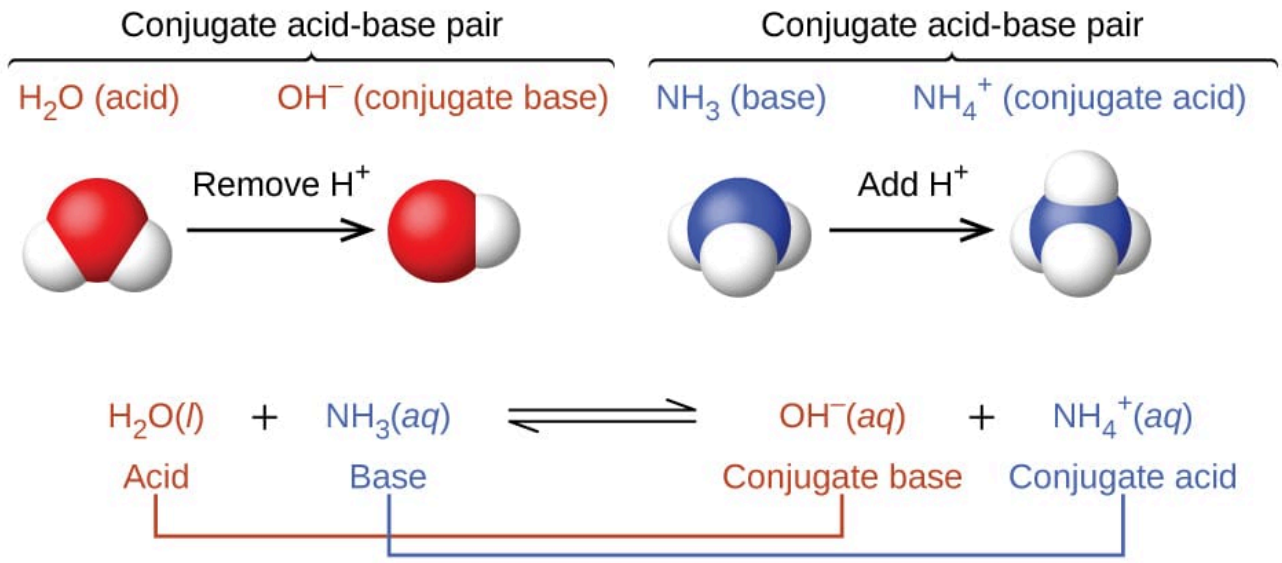


Figure 7.4.2.This figure has three parts in two rows. In the first row, two diagrams of acid-base pairs are shown. On the left, a space filling model of H subscript 2 O is shown with a red O atom at the center and two smaller white H atoms attached in a bent shape. Above this model is the label “H subscript 2 O (acid)” in purple. An arrow points right, which is labeled “Remove H superscript plus.” To the right is another space filling model with a single red O atom to which a single smaller white H atom is attached. The label in purple above this model reads, “O H superscript negative (conjugate base).” Above both of these red and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” To the right is a space filling model with a central blue N atom to which three smaller white H atoms are attached in a triangular pyramid arrangement. A label in green above reads “N H subscript 3 (base).” An arrow labeled “Add H superscript plus” points right. To the right of the arrow is another space filling model with a blue central N atom and four smaller white H atoms in a tetrahedral arrangement. The green label above reads “N H subscript 3 superscript plus (conjugate acid).” Above both of these blue and white models is an upward pointing bracket that is labeled “Conjugate acid-base pair.” The second row of the figure shows the chemical reaction, H subscript 2 O (l) is shown in purple, and is labeled below in purple as “acid,” plus N H subscript 3 (a q) in green, labeled below in green as “base,” followed by a double sided arrow arrow and O H superscript negative (a q) in purple, labeled in purple as “conjugate base,” plus N H subscript 4 superscript plus (a q)” in green, which is labeled in green as “conjugate acid.” The acid on the left side of the equation is connected to the conjugate base on the right with a purple line. Similarly, the base on the left is connected to the conjugate acid on the right side.

In the reaction of ammonia with water to give ammonium ions and hydroxide ions, ammonia acts as a base by accepting a proton from a water molecule, which in this case means that water is acting as an acid. In the reverse reaction, an ammonium ion acts as an acid by donating a proton to a hydroxide ion, and the hydroxide ion acts as a base. The conjugate acid–base pairs for this reaction are NH_4^+/NH_3 and H_2O/OH^-.

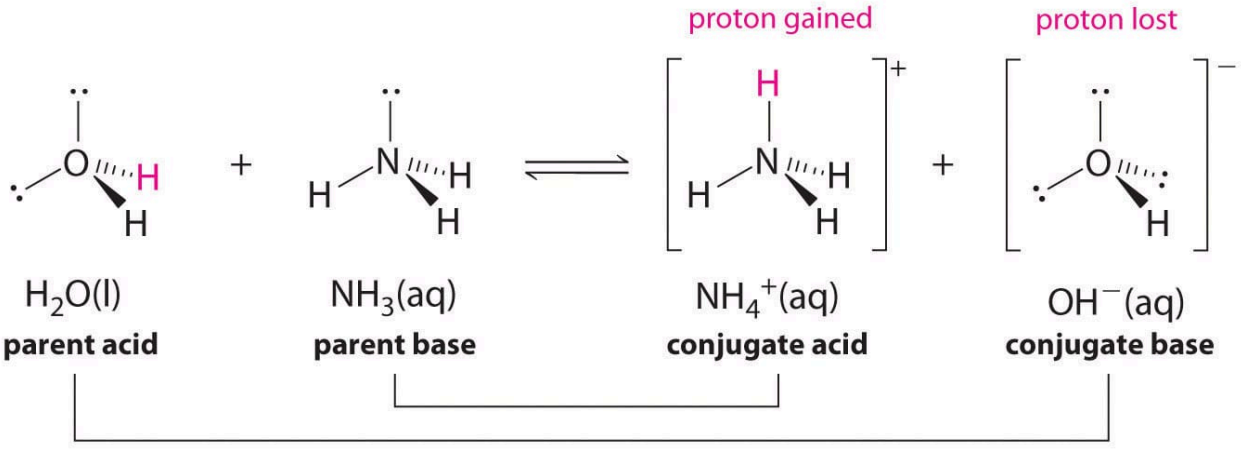


Figure 7.4.2. The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid

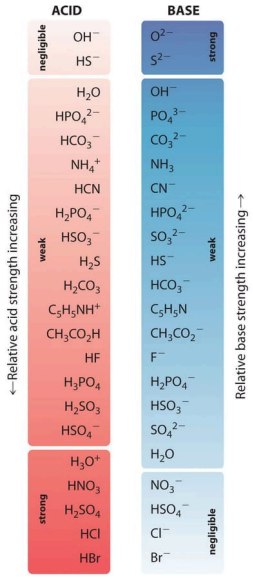


Figure 7.4.3 The Relative Strengths of Some Common Conjugate Acid–Base Pairs

The strongest acids are at the bottom left, and the strongest bases are at the top right. The conjugate base of a strong acid is a very weak base, and, conversely, the conjugate acid of a strong base is a very weak acid.

Example 7.4.1

Identify the conjugate acid-base pairs in this equilibrium.

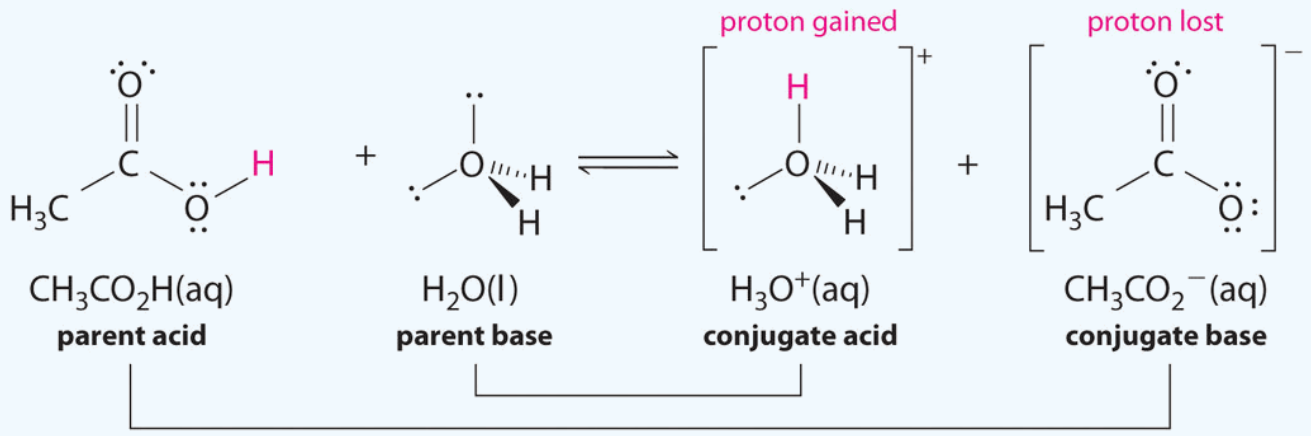
$$\mathrm{CH_3CO_2H + H_2O \rightleftharpoons H_3O^+ + CH_3CO_2^-}$$

Solution

Similarly, in the reaction of acetic acid with water, acetic acid **donates** a proton to water, which acts as the base. In the reverse reaction, H_3O^+ is the acid that donates a proton to the acetate ion, which acts as the base.

Once again, we have two conjugate acid–base pairs:

- the parent acid and its conjugate base ($CH_3CO_2H/CH_3CO_2^-$) and
- the parent base and its conjugate acid (H_3O^+/H_2O).



Example 7.4.1 the reaction of acetic acid with water

Example 7.4.2

Identify the conjugate acid-base pairs in this equilibrium.

$$(CH_3)_3N + H_2O \rightleftharpoons (CH_3)_3NH^+ + OH^-$$

Solution

One pair is H2O and OH^-, where H2O has one more H^+ and is the conjugate acid, while OH^- has one less H^+ and is the conjugate base.

The other pair consists of (CH3)3N and $(CH_3)_3NH^+$, where $(CH_3)_3NH^+$ is the conjugate acid (it has an additional proton) and (CH3)3N is the conjugate base.

Exercise 7.4.1

Identify the conjugate acid-base pairs in this equilibrium.

$$NH_2^- + H_2O \rightleftharpoons NH_3 + OH^-$$

Answer: available at the end of this section.

Buffer Solutions

Weak acids are relatively common, even in the foods we eat. But we occasionally encounter a strong acid or base, such as stomach acid, which has a strongly acidic pH of 1.7. By definition, strong acids and bases can produce a relatively large amount of H^+ or OH^- ions and consequently have marked chemical activities. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [approximated as 0.1 M HCl(aq)] were added to the bloodstream and no correcting mechanism were present, the pH of the blood would decrease from about 7.4 to about 4.7—a pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

This mechanism involves a **buffer**, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid, or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved HC2H3O2 (a weak acid) and NaC2H3O2 (the salt derived from that weak acid). Another example of a buffer is a solution containing NH3 (a weak base) and NH4Cl (a salt derived from that weak base).

Let us use an $HC_2H_3O_2/NaC_2H_3O_2$ buffer to demonstrate how buffers work. If a strong base—a source of $OH^-(aq)$ ions—is added to the buffer solution, those OH^- ions will react with the $HC_2H_3O_2$ in an acid-base reaction:

$$HC_2H_3O_2\,(aq) + OH^-\,(aq) \rightarrow H_2O(\ell) + C_2H_3O_2^-\,(aq) \tag{7.4.1}$$

Rather than changing the pH dramatically by making the solution basic, the added OH^- ions react to make H_2O, so the pH does not change much.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of H^+ ions and $C_2H_3O_2^-$ ions are present in the same solution, they will come together to make $HC_2H_3O_2$:

$$H^+\,(aq) + C_2H_3O_2^-\,(aq) \rightarrow HC_2H_3O_2\,(aq) \tag{7.4.2}$$

Rather than changing the pH dramatically and making the solution acidic, the added H^+ ions react to make molecules of a weak acid. Figure 7.4.2 illustrates both actions of a buffer.

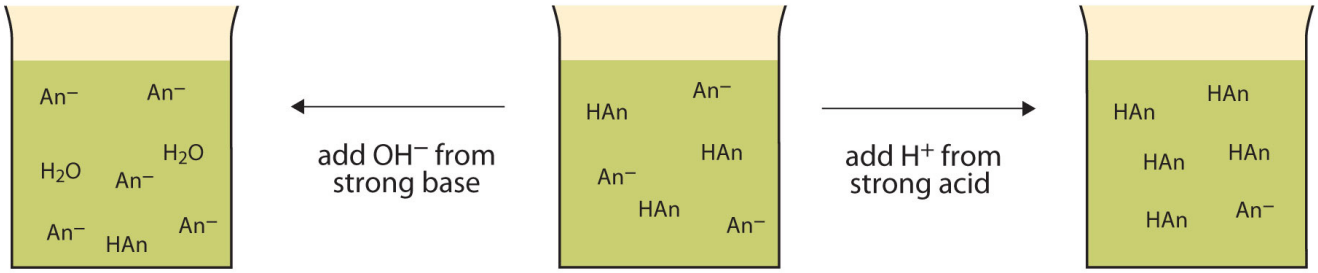


Figure 7.4.4 The actions of buffers. Buffers can react with both strong acids (top) and strong bases (side) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl, NH_3 molecules can react with any excess H^+ ions introduced by strong acids:

$$NH_3(aq) + H^+(aq) \rightarrow NH_4^+(aq)$$

while the $NH_4^+(aq)$ ion can react with any OH^- ions introduced by strong bases:

$$NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(\ell)$$

Some common buffer systems are listed in the table below.

Table 7.4.1 Some Common Buffers

Table 7.4.1 Some Common Buffers		
Buffer System	**Buffer Components**	**pH of buffer (equal molarities of both components)**
Acetic acid/acetate ion	CH_3COOH/CH_3COO^-	4.74
Carbonic acid/hydrogen carbonate ion	H_2CO_3/HCO_3^-	6.38
Dihydrogen phosphate ion/hydrogen phosphate ion	$H_2PO_4^-/HPO_4^{2-}$	7.21
Ammonia/ammonium ion	NH_3/NH_4^+	9.25

Buffers work well only for limited amounts of added strong acid or base. Once either solute is completely reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain **capacity**. Buffers that have more solute dissolved in them to start with have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [the second compound is another way to write $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Example 7.4.3:

Which combinations of compounds can make a buffer solution?

1. $HCHO_2$ and $NaCHO_2$
2. HCl and NaCl
3. CH_3NH_2 and CH_3NH_3Cl
4. NH_3 and NaOH

Solution

1. $HCHO_2$ is formic acid, a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid (the formate ion [CHO_2^-]). The combination of these two solutes would make a buffer solution.
2. HCl is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
3. CH_3NH_2 is methylamine, which is like NH_3 with one of its H atoms substituted with a CH_3 group. Because it is not listed in Table 7.4.1, we can assume that it is a weak base. The compound CH_3NH_3Cl is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
4. NH_3 is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

Exercise 7.4.2

Which combinations of compounds can make a buffer solution?

1. $NaHCO_3$ and NaCl
2. H_3PO_4 and NaH_2PO_4
3. NH_3 and $(NH_4)_3PO_4$
4. NaOH and NaCl

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Food and Drink Application: the Acid that Eases Pain

Although medicines are not exactly "food and drink," we do ingest them, so let's take a look at an acid that is probably the most common medicine: acetylsalicylic acid, also known as aspirin. Aspirin is well known as a pain reliever and antipyretic (fever reducer).

The structure of aspirin is shown in the accompanying figure. The acid part is circled; it is the H atom in that part that can be donated as aspirin acts as a Brønsted-Lowry acid. Because it is not given in Table 7.4.1, acetylsalicylic acid is a weak acid. However, it is still an acid, and given that some people consume relatively large amounts of aspirin daily, its acidic nature can cause problems in the stomach lining, despite the stomach's defenses against its own stomach acid.

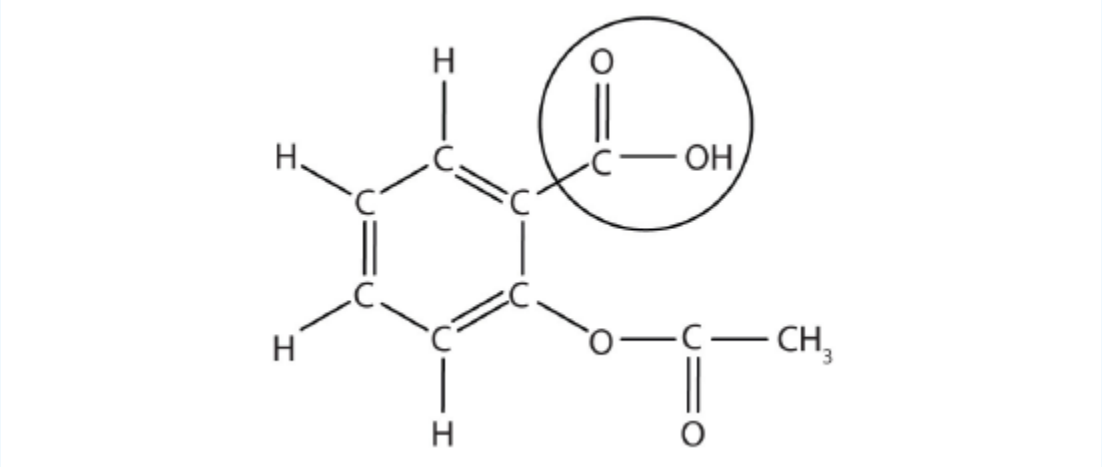


Figure 7.4.5: The Molecular Structure of Aspirin. The circled atoms are the acid part of the molecule.

Because the acid properties of aspirin may be problematic, many aspirin brands offer a "buffered aspirin" form of the medicine. In these cases, the aspirin also contains a buffering agent—usually MgO—that regulates the acidity of the aspirin to minimize its acidic side effects.

As useful and common as aspirin is, it was formally marketed as a drug starting in 1899. The US Food and Drug Administration (FDA), the governmental agency charged with overseeing and approving drugs in the United States, wasn't formed until 1906. Some have argued that if the FDA had been formed before aspirin was introduced, aspirin may never have gotten approval due to its potential for side effects—gastrointestinal bleeding, ringing in the ears, Reye's syndrome (a liver problem), and some allergic reactions. However, recently aspirin has been touted for its effects in lessening heart attacks and strokes, so it is likely that aspirin will remain on the market.

Summary

- A buffer is a solution that resists sudden changes in pH.
- Reactions showing how buffers regulate pH are described.

Contributors and Attributions

-

- TextMap: Beginning Chemistry (Ball et al.)

- Marisa Alviar-Agnew (Sacramento City College)

Learning Objectives

- Define *buffer*.
- Correctly identify the two components of a buffer.

As indicated in Section 12.5, weak acids are relatively common, even in the foods we eat. But we occasionally encounter a strong acid or base, such as stomach acid, which has a strongly acidic pH of 1.7. By definition, strong acids and bases can produce a relatively large amount of H^+ or OH^- ions and consequently have marked chemical activities. In addition, very small amounts of strong acids and bases can change the pH of a solution very quickly. If 1 mL of stomach acid [approximated as 0.1 M HCl(aq)]

were added to the bloodstream and no correcting mechanism were present, the pH of the blood would decrease from about 7.4 to about 4.7—a pH that is not conducive to continued living. Fortunately, the body has a mechanism for minimizing such dramatic pH changes.

This mechanism involves a **buffer**, a solution that resists dramatic changes in pH. Buffers do so by being composed of certain pairs of solutes: either a weak acid plus a salt derived from that weak acid, or a weak base plus a salt of that weak base. For example, a buffer can be composed of dissolved $HC_2H_3O_2$ (a weak acid) and $NaC_2H_3O_2$ (the salt derived from that weak acid). Another example of a buffer is a solution containing NH_3 (a weak base) and NH_4Cl (a salt derived from that weak base).

Let us use an $HC_2H_3O_2/NaC_2H_3O_2$ buffer to demonstrate how buffers work. If a strong base—a source of $OH^-(aq)$ ions—is added to the buffer solution, those OH^- ions will react with the $HC_2H_3O_2$ in an acid-base reaction:

Label 'Eq1' multiply defined

Rather than changing the pH dramatically by making the solution basic, the added OH^- ions react to make H_2O, so the pH does not change much.

If a strong acid—a source of H^+ ions—is added to the buffer solution, the H^+ ions will react with the anion from the salt. Because $HC_2H_3O_2$ is a weak acid, it is not ionized much. This means that if lots of H^+ ions and $C_2H_3O_2^-$ ions are present in the same solution, they will come together to make $HC_2H_3O_2$:

Label 'Eq2' multiply defined

Rather than changing the pH dramatically and making the solution acidic, the added H^+ ions react to make molecules of a weak acid. Figure 7.4.1 illustrates both actions of a buffer.

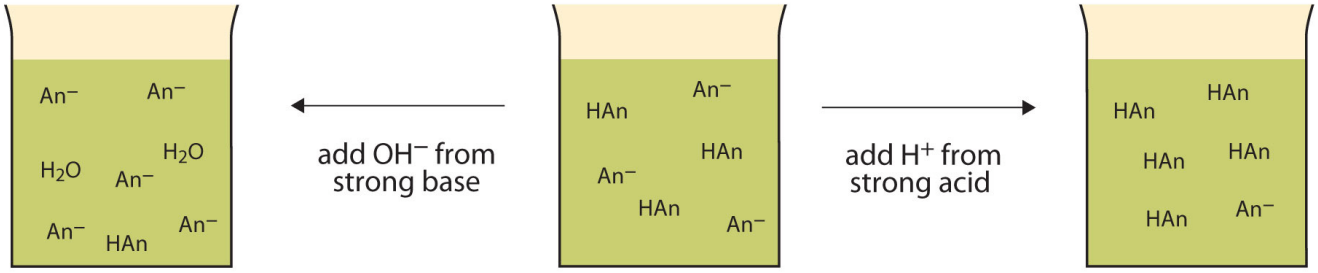


Figure 7.4.1 The Actions of Buffers. Buffers can react with both strong acids (top) and strong bases (side) to minimize large changes in pH.

Buffers made from weak bases and salts of weak bases act similarly. For example, in a buffer containing NH_3 and NH_4Cl, NH_3 molecules can react with any excess H^+ ions introduced by strong acids:

$$NH_3\,(aq) + H^+\,(aq) \rightarrow NH_4^+\,(aq) \tag{7.4.3}$$

while the **ammonium ion** ($NH_4^+\,(aq)$) can react with any hydroxide ions introduced by strong bases:

$$NH_4^+\,(aq) + OH^-\,(aq) \rightarrow NH_3\,(aq) + H_2O(\ell) \tag{7.4.4}$$

Example 7.4.1

Which combinations of compounds can make a buffer solution?

1. $HCHO_2$ and $NaCHO_2$
2. HCl and NaCl

3. CH_3NH_2 and CH_3NH_3Cl
4. NH_3 and NaOH

Solution

1. $HCHO_2$ is formic acid, a weak acid, while $NaCHO_2$ is the salt made from the anion of the weak acid (the formate ion [CHO_2^-]). The combination of these two solutes would make a buffer solution.
2. HCl is a strong acid, not a weak acid, so the combination of these two solutes would not make a buffer solution.
3. CH_3NH_2 is methylamine, which is like NH_3 with one of its H atoms substituted with a CH_3 group. Because it is not listed in Table 7.4.1, we can assume that it is a weak base. The compound CH_3NH_3Cl is a salt made from that weak base, so the combination of these two solutes would make a buffer solution.
4. NH_3 is a weak base, but NaOH is a strong base. The combination of these two solutes would not make a buffer solution.

Exercise 7.4.1

Which combinations of compounds can make a buffer solution?

1. $NaHCO_3$ and NaCl
2. H_3PO_4 and NaH_2PO_4
3. NH_3 and $(NH_4)_3PO_4$
4. NaOH and NaCl

Answer a available at the end of this section.

Answer b available at the end of this section.

Answer c available at the end of this section.

Answer d available at the end of this section.

Buffers work well only for limited amounts of added strong acid or base. Once either solute is completely reacted, the solution is no longer a buffer, and rapid changes in pH may occur. We say that a buffer has a certain **capacity**. Buffers that have more initial solute dissolved within them have larger capacities, as might be expected.

Human blood has a buffering system to minimize extreme changes in pH. One buffer in blood is based on the presence of HCO_3^- and H_2CO_3 [the second compound is another way to write $CO_2(aq)$]. With this buffer present, even if some stomach acid were to find its way directly into the bloodstream, the change in the pH of blood would be minimal. Inside many of the body's cells, there is a buffering system based on phosphate ions.

Food and Drink Application: the Acid that Eases Pain

Although medicines are not exactly "food and drink," we do ingest them, so let's take a look at an acid that is probably the most common medicine: acetylsalicylic acid, also known as aspirin. Aspirin is well known as a pain reliever and antipyretic (fever reducer).

The structure of aspirin is shown in the accompanying figure. The acid part is circled; it is the H atom in that part that can be donated as aspirin acts as a Brønsted-Lowry acid. Because it is not given in Table 7.4.1, acetylsalicylic acid is a weak acid. However, it is still an acid, and given that some people consume relatively large amounts of aspirin daily, its acidic nature can cause problems in the stomach lining, despite the stomach's defenses against its own stomach acid.

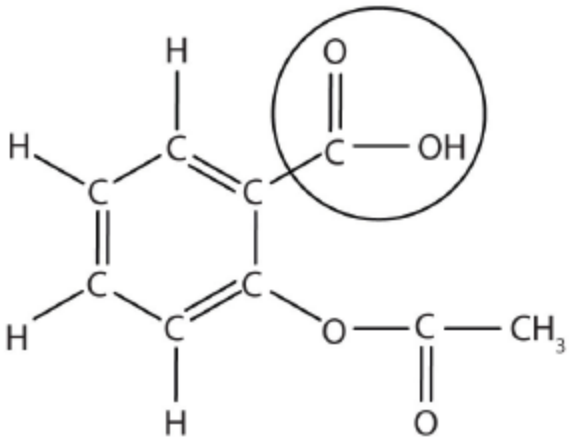


Figure 7.4.2: The Molecular Structure of Aspirin. The circled atoms are the acid part of the molecule.
The molecule of aspartame is shown. The CO2H is circled to indicate that it is the acidic part of the molecule.

Because the acid properties of aspirin may be problematic, many aspirin brands offer a "buffered aspirin" form of the medicine. In these cases, the aspirin also contains a buffering agent—usually MgO—that regulates the acidity of the aspirin to minimize its acidic side effects.

As useful and common as aspirin is, it was formally marketed as a drug starting in 1899. The US Food and Drug Administration (FDA), the governmental agency charged with overseeing and approving drugs in the United States, wasn't formed until 1906. Some have argued that if the FDA had been formed before aspirin was introduced, aspirin may never have gotten approval due to its potential for side effects—gastrointestinal bleeding, ringing in the ears, Reye's syndrome (a liver problem), and some allergic reactions. However, recently aspirin has been touted for its effects in lessening heart attacks and strokes, so it is likely that aspirin will remain on the market.

Summary

A buffer is a solution that resists sudden changes in pH.

Answers

Exercise 7.4.1

H_2O (acid) and OH^- (base); NH_2^- (base) and NH_3 (acid)

Exercise 7.4.2

Yes.

Exercise 7.4.2

No. Need a weak acid or base and a salt of its conjugate base or acid.

Exercise 7.4.2

Yes.

Exercise 7.4.2

No. Need a weak base or acid.

Exercise 7.4.1

Yes.

Exercise 7.4.1

No. Need a weak acid or base and a salt of its conjugate base or acid.

Exercise 7.4.1

Yes.

Exercise 7.4.1

No. Need a weak base or acid.

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7.5: Redox Chemistry

Learning Objectives

- Define *oxidation* and *reduction*.
- Assign oxidation numbers to atoms in simple compounds.
- Recognize a reaction as an oxidation-reduction reaction.

Oxygen is an element that has been known for centuries. In its pure elemental form, oxygen is highly reactive, and it readily makes compounds with most other elements. It is also the most abundant element by mass in the Earth's crust. The class of reactions called oxidation and reduction were originally defined with respect to the element oxygen.

Many elements simply combine with oxygen to form the oxide of that element. Heating magnesium in air allows it to combine with oxygen to form magnesium oxide.

$$2\mathrm{Mg}\,(s) + \mathrm{O_2}\,(g) \rightarrow 2\mathrm{MgO}\,(s)$$

Many compounds react with oxygen as well, often in very exothermic processes that are generally referred to as combustion reactions. For example, when methane burns, carbon dioxide and water are produced.

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

Carbon dioxide is an oxide of carbon, while water is an oxide of hydrogen. Early scientists viewed oxidation as a process in which a substance was reacted with oxygen to produce one or more oxides. In the previous examples, magnesium and methane are being oxidized.

The definitions of oxidation and reduction were eventually broadened to include similar types of reactions that do not necessarily involve oxygen. Oxygen is more electronegative than any element except for fluorine. Therefore, when oxygen is bonded to any element other than fluorine, electrons from the other atom are shifted away from that atom and toward the oxygen atom. An **oxidation-reduction reaction** (sometimes abbreviated as a **redox reaction**) *is a reaction that involves the full or partial transfer of electrons from one reactant to another*. **Oxidation** *involves a full or partial loss of electrons,* while **reduction** *involves a full or partial gain of electrons*.

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use **oxidation numbers** to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on four rules. Oxidation numbers are not necessarily equal to the charge on the atom (although sometimes they can be); we must keep the concepts of charge and oxidation numbers separate.

Assigning Oxidation Numbers

The **oxidation number** is a positive or negative number that is assigned to an atom to indicate its degree of oxidation or reduction. In oxidation-reduction processes, the driving force for chemical change is in the exchange of electrons between chemical species. A series of rules have been developed to help us.

Assigning Oxidation Numbers

The rules for assigning oxidation numbers to atoms are as follows:

1. Atoms in their elemental state are assigned an oxidation number of 0.
   Examples: $\mathrm{H_2}$, $\mathrm{Br_2}$, Na, Be, K, $\mathrm{O_2}$, $\mathrm{P_4}$, all have oxidation number of 0.
2. Atoms in monatomic (i.e., one-atom) ions are assigned an oxidation number equal to their charge.
   Example: In MgCl2, magnesium has an oxidation number of +2, while each chlorine atom has an oxidation number of –1

3. In compounds, fluorine is assigned a –1 oxidation number; oxygen is usually assigned a –2 oxidation number [except in peroxide compounds (where it is –1) and in binary compounds with fluorine (where it is positive)]; and hydrogen is usually assigned a +1 oxidation number [except when it exists as the hydride ion (H^-), in which case rule 2 prevails].
   Example: In H_2O, the H atoms each have an oxidation number of +1, while the O atom has an oxidation number of –2, even though hydrogen and oxygen do not exist as ions in this compound.
4. In compounds, all other atoms are assigned an oxidation number so that the sum of the oxidation numbers on all the atoms in the species equals the charge on the species (which is zero if the species is neutral).
   Example: In SO_2. Each O atom has an oxidation number of –2; for the sum of the oxidation numbers to equal the charge on the species (which is zero), the S atom is assigned an oxidation number of +4. Does this mean that the sulfur atom has a 4+ charge on it? No, it means only that the S atom is assigned a +4 oxidation number by our rules of apportioning electrons among the atoms in a compound.

Note

Oxidation numbers are usually written with the sign first, then the magnitude, to differentiate them from charges.

Example 7.5.1

Assign oxidation numbers to the atoms in each substance.

1. Cl_2
2. GeO_2
3. $CaCl_2$

Solution

1. Cl_2 is the elemental form of chlorine. Rule 1 states each atom has an oxidation number of 0.
2. By rule 3, oxygen is normally assigned an oxidation number of –2. For the sum of the oxidation numbers to equal the charge on the species (zero), the Ge atom is assigned an oxidation number of +4.
3. The Ca^{2+} ion has an oxidation number of +2 by rule 2. Also according to rule 2, each chlorine atom is assigned an oxidation number of –1.

Exercise 7.5.1: Phosphoric Acid

Assign oxidation numbers to the atoms in H_3PO_4.

Answer available at the end of this section.

Oxidation and Reduction Reactions

The processes of oxidation and reduction are commonly viewed in terms of the gain or loss of electrons or the increase or decrease in oxidation numbers. However, many reactions in organic chemistry that involve the gain or loss of oxygen or hydrogen are also considered as redox reactions.

The table below summarizes the processes of oxidation and reduction.

Table 7.5.1 The Processes of Oxidation and Reduction.

Oxidation	Reduction
Complete loss of electrons (ionic reaction)	Complete gain of electrons (ionic reaction)

Oxidation	Reduction
Increase in oxidation number	Decrease in oxidation number
Gain of oxygen	Loss of oxygen
Loss of hydrogen in a molecular compound	Gain of hydrogen in a molecular compound

1. Oxidation is complete loss of electrons. Reduction is complete gain of electrons

To understand electron-transfer reactions like the one between zinc metal and hydrogen ions, chemists separate them into two parts: one part focuses on the loss of electrons, and one part focuses on the gain of electrons. The loss of electrons is called oxidation. The gain of electrons is called reduction.

Electrons that are lost are written as products; electrons that are gained are written as reactants. For example, the reaction below represents when acid is added to zinc metal.

$$\mathrm{Zn\,(s) + 2\,H^+\,(aq) \longrightarrow Zn^{2+}\,(aq) + H_2\,(g)}$$

zinc atoms are oxidized to Zn^{2+}. The half reaction for the oxidation reaction, omitting phase labels, is as follows:

$$\mathrm{Zn \longrightarrow Zn^{2+} + 2\,e^-}$$

This half reaction is balanced in terms of the number of zinc atoms, and it also shows the two electrons that are needed as products to account for the zinc atom losing two negative charges to become a 2+ ion. With half reactions, there is one more item to balance: the overall charge on each side of the reaction. If you check each side of this reaction, you will note that both sides have a zero net charge.

Hydrogen is reduced in the reaction. The balanced reduction half reaction is as follows:

$$\mathrm{2\,H^+ + 2\,e^- \longrightarrow H_2}$$

There are two hydrogen atoms on each side, and the two electrons written as reactants serve to neutralize the 2+ charge on the reactant hydrogen ions. Again, the overall charge on both sides is zero.

Note: "LEO says GER"

LEO stands for Losing Electrons is Oxidation, while GER stands for Gaining Electrons is Reduction.



Figure 7.5.1

> *This guy knows his redox reactions.*

2. Oxidation is an increase in oxidation number. Reduction is a decrease in oxidation number.

Consider the reaction below between elemental iron and copper sulfate:

$$\mathrm{Fe + CuSO_4 \rightarrow FeSO_4 + Cu}$$

In the course of the reaction, the oxidation number of Fe increases from zero to $+2$. The oxidation number of copper decreases from $+2$ to 0. This result is in accordance with the activity series. Iron is above copper in the series, so will be more likely to form $\mathrm{Fe^{2+}}$ while converting the $\mathrm{Cu^{2+}}$ to metallic copper $(\mathrm{Cu^0})$.

3. Oxidation is gain of oxygen. Reduction is loss of oxygen.

Oxidation and reduction can also be defined in terms of changes in composition. The original meaning of oxidation was “adding oxygen,” so when oxygen is added to a molecule, the molecule is being oxidized. The reverse is true for reduction: if a molecule loses oxygen atoms, the molecule is being reduced. For example, the acetaldehyde ($CH3CHO$) molecule takes on an oxygen atom to become acetic acid ($\mathrm{CH_3COOH}$).

$$\mathrm{2\,CH_3CHO + O_2 \longrightarrow 2\,CH_3COOH}$$

Thus, acetaldehyde is being oxidized.

4. Oxidation is loss of hydrogen in a molecular compound. Reduction is gain of hydrogen in a molecular compound.

Oxidation and reduction can be defined in terms of the gain or loss of hydrogen atoms. If a molecule adds hydrogen atoms, it is being reduced. If a molecule loses hydrogen atoms, the molecule is being oxidized. For example, in the conversion of acetaldehyde into ethanol CH_3CHO, hydrogen atoms are added to acetaldehyde, so the acetaldehyde is being reduced:

$$\mathrm{CH_3CHO + H_2 \longrightarrow CH_3CH_2OH}$$

Example 7.5.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $N_2 \rightarrow NH_3$
b. $CH_3CH_2OHCH_3 \rightarrow CH_3COCH_3$
c. $HCHO \rightarrow HCOOH$

Solution

a. Hydrogen is being added to the original reactant molecule, so reduction is occurring.
b. Hydrogen is being removed from the original reactant molecule, so oxidation is occurring.
c. Oxygen is being added to the original reactant molecule, so oxidation is occurring.

Exercise 7.5.2

In each conversion, indicate whether oxidation or reduction is occurring.

a. $CH_4 \rightarrow CO_2 + H_2O$
b. $NO_2 \rightarrow N_2$
c. $CH_2{=}CH_2 \rightarrow CH_3CH_3$

Answers

a. oxidation

b. reduction
c. reduction

Example 7.5.3:

Identify what is being oxidized and reduced in this redox reaction.

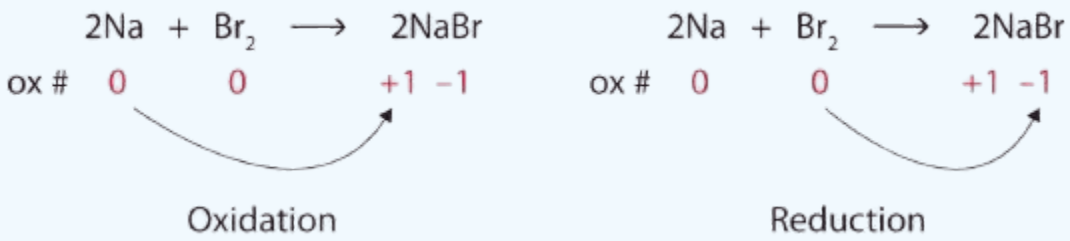
$$2\,\mathrm{Na} + \mathrm{Br_2} \longrightarrow 2\,\mathrm{NaBr}$$

Solution

Both reactants are the elemental forms of their atoms, so the Na and Br atoms have oxidation numbers of 0. In the ionic product, the Na^+ ions have an oxidation number of +1, while the Br^- ions have an oxidation number of –1.

$$\overset{0}{2Na} + \overset{0}{Br_2} \rightarrow \overset{+1-1}{2NaBr}$$

Sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; bromine is decreasing its oxidation number from 0 to –1, so it is being reduced:



Example 7.5.3

$$2\,\mathrm{Na} + \mathrm{Br_2} \longrightarrow 2\,\mathrm{NaBr}$$

Because oxidation numbers are changing, this is a redox reaction. The total number of electrons being lost by sodium (two, one lost from each Na atom) is gained by bromine (two, one gained for each Br atom).

Exercise 7.5.3

Identify what is being oxidized and reduced in this redox reaction.

$$\mathrm{C} + \mathrm{O_2} \longrightarrow \mathrm{CO_2}$$

Answer available at the end of this section.

Example 7.5.4: Gain or Loss of Electrons

For each half reaction, identify if there was a loss or gain of electrons.

a. $2N^{3-} \rightarrow N_2$
b. $Cr^{6+} \rightarrow Cr^{3+}$

Solutions

a. When we compare the overall charges on each side of the equation, we find a charge of 2(-3)=-6 on the left but a charge of 0 on the right. This equation is not properly balanced. To balance it, six electrons must be added on the right side as follows:

\[\ce{2N^{3-} → N2 + 6e^{-}}\nonumber \]

So there was a *loss of six electrons*. Now the equation is balanced, not only in terms of elements but also in terms of charge.

b. When we compare the overall charges on each side of the equation, we find a charge of +6 on the left and a charge of +3 on the right. This equation is not properly balanced. To end up with an oxidation number of +3 on the right side 3 electrons must be added on the left side as follows:

\[\ce{Cr^{6+} + 3e^{-} → Cr^{3+}\nonumber \]

So there was a *gain of three electrons*. Now the equation is balanced in terms of charge.

Exercise 7.5.4 Gain or Loss of Electrons

For each half reaction, identify if there was a loss or gain of electrons.

a. $O_2 \rightarrow 2O^{2-}$
b. $Cu^{+} \rightarrow Cu^{2+}$

Answers

a. gain of four electrons

b. loss of one electron

Example 7.5.5: Reducing Silver Ions

Write and balance the redox reaction that has silver ions and aluminum metal as reactants and silver metal and aluminum ions as products.

Solution

We start by using symbols of the elements and ions to represent the reaction:

$$\mathrm{Ag^+ + Al \longrightarrow Ag + Al^{3+}}$$

The equation looks balanced as it is written. However, when we compare the overall charges on each side of the equation, we find a charge of +1 on the left but a charge of +3 on the right. This equation is not properly balanced. To balance it, let us write the two half reactions. Silver ions are reduced, and it takes one electron to change Ag^+ to Ag:

$$\mathrm{Ag^+ + e^- \longrightarrow Ag}$$

Aluminum is oxidized, losing three electrons to change from Al to Al^{3+}:

$$\mathrm{Al \longrightarrow Al^{3+} + 3\,e^-}$$

To combine these two half reactions and cancel out all the electrons, we need to multiply the silver reduction reaction by 3:

$$\begin{array}{r} 3\left(\mathrm{Ag^+ + e^- \rightarrow Ag}\right) \\ \mathrm{Al \rightarrow Al^{3+} + 3e^-} \\ \hline \mathrm{3Ag^+ + Al \rightarrow 3Ag + Al^{3+}} \end{array}$$

Example 7.5.5 Silver reduction reaction

Now the equation is balanced, not only in terms of elements but also in terms of charge.

Exercise 7.5.5

Write and balance the redox reaction that has calcium ions and potassium metal as reactants and calcium metal and potassium ions as products.

Summary

- Rules for determining oxidation numbers are listed.
- Examples of oxidation number determinations are provided.
- Chemical reactions in which electrons are transferred are called oxidation-reduction, or redox, reactions.
- Oxidation is defined in terms of the loss of electrons, gain of oxygen, loss of hydrogen, or an increase in oxidation number.
- Reduction is defined in terms of gain of electrons, loss of oxygen, gain of hydrogen, or a decrease in oxidation number.
- Oxidation and reduction always occur together, even though they can be written as separate chemical equations.

Contributors and Attributions

-

- Libretext: Beginning Chemistry (Ball et al.)
- Libretext: The Basics of GOB Chemistry (Ball et al.)
- Libretext: Chemistry for Allied Health (Soult)

- Marisa Alviar-Agnew (Sacramento City College)

- Ed Vitz (Kutztown University), John W. Moore (UW-Madison), Justin Shorb (Hope College), Xavier Prat-Resina (University of Minnesota Rochester), Tim Wendorff, and Adam Hahn.

Learning Objective

- Identify oxidizing and reducing agents.

Life on planet Earth is a complicated and well-organized set of processes. Animals are designed to breathe oxygen and plants are designed to produce oxygen. Photosynthesis is the means by which we get the oxygen we need for life. Light striking a plant pigment known as chlorophyll initiates a complex series of reactions, many of which involve redox processes complete with movement of electrons. In this series of reactions, water is converted to oxygen gas, and we have something to sustain our lives.

The reaction below is a redox reaction that produces zinc sulfide:

$$\mathrm{Zn} + \mathrm{S} \rightarrow \mathrm{ZnS}$$

The half-reactions can be written:

$$\mathrm{Oxidation:\ Zn} \rightarrow \blacksquare + 2\mathrm{e}^-$$
$$\mathrm{Reduction:\ S} + 2\mathrm{e}^- \rightarrow \blacksquare$$

In the reaction above, zinc is being oxidized by losing electrons. However, there must be another substance present that gains those electrons and in this case that is the sulfur. In other words, the sulfur is causing the zinc to be oxidized. Sulfur is called the oxidizing agent. The zinc causes the sulfur to gain electrons and become reduced and so the zinc is called the reducing agent. The **oxidizing agent** is a substance that causes oxidation by accepting electrons; therefore, its oxidation state decreases. The **reducing agent** is a substance that causes reduction by losing electrons; therefore its oxidation state increases. The simplest way to think of this is that the oxidizing agent is the substance that is reduced, while the reducing agent is the substance that is oxidized as shown in Figure 7.5.1 and summarized in Table 7.5.1.

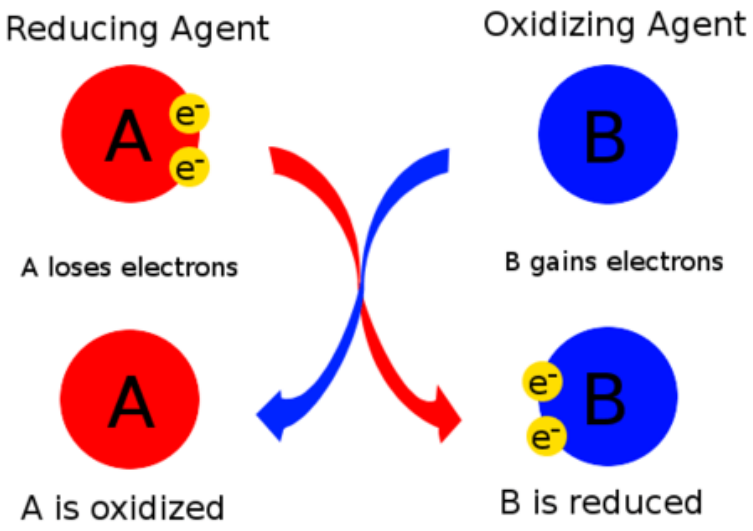


Figure 7.5.1 A reducing agent is oxidized and an oxidizing agent is reduced.

Note

Both the oxidizing and reducing agents are the *reactants* and therefore appear on the *left*-hand side of an equation.

Table 7.5.1: A Comparison of Oxidizing and Reducing Agents.

	Oxidizing Agents	Reducing Agents
Oxidation State	Decreases	Increases
# of Electrons	Gained	Lost
Substance is...	Reduced	Oxidized

The examples below show how to analyze a redox reaction and identify oxidizing and reducing agents.

Example 7.5.1 Half-equations

When chlorine gas is bubbled into a solution of sodium bromide, a reaction occurs which produces aqueous sodium chloride and bromine. Determine what is being oxidized and what is being reduced. Identify the oxidizing and reducing agents.

$$\mathrm{Cl_2\,(g) + 2NaBr\,(aq) \rightarrow 2NaCl\,(aq) + Br_2\,(l)}$$

Solution:

Step 1: Plan the problem.

Break the reaction down into a net ionic equation and then into half-reactions. The substance that loses electrons is being oxidized and is the reducing agent. The substance that gains electrons is being reduced and is the oxidizing agent.

Step 2: Solve.

$$\mathrm{Cl_2\,(g) + \cancel{2Na^+\,(aq)} + 2Br^-\,(aq) \rightarrow \cancel{2Na^+\,(aq)} + 2Cl^-\,(aq) + Br_2\,(l)}$$

$$\mathrm{Cl_2\,(g) + 2Br^-\,(aq) \rightarrow 2Cl^-\,(aq) + Br_2\,(l) \quad (net\ ionic\ equation)}$$

$$\mathrm{Reduction:\ Cl_2\,(g) + 2e^- \rightarrow 2Cl^-\,(aq)}$$
$$\mathrm{Oxidation:\ 2Br^-\,(aq) \rightarrow Br_2\,(l) + 2e^-}$$

The $\mathrm{Cl_2}$ is being reduced and is the oxidizing agent. The $\mathrm{Br^-}$ is being oxidized and is the reducing agent.

Exercise 7.5.1 : Half-equations

Write the following reaction in the form of half-equations. Identify each half-equation as an oxidation or a reduction. Also identify the oxidizing agent and the reducing agent in the overall reaction

$$\mathrm{Zn + 2\,Fe^{3+} \longrightarrow Zn^{2+} + 2\,Fe^{2+}}$$

Answer

The half-equations are

$\mathrm{Zn \longrightarrow Zn^{2+} + 2\,e^-}$ oxidation—loss of electrons

$\mathrm{2\,e^- + 2\,Fe^{3+} \longrightarrow 2\,Fe^{2+}}$ reduction—gain of electrons

Zinc has been oxidized, the oxidizing *agent* must have been the other reactant, namely, iron(III).

Iron(III) ion has been reduced, the zinc must be the reducing agent.

Example 7.5.2: Identify reducing and oxidizing agents

Identify the reducing and oxidizing agents in the balanced redox reaction:

$$Cl_2(aq) + 2Br^-(aq) \rightarrow 2Cl^-(aq) + Br_2(aq)$$

Oxidation half reaction

$$2Br^-(aq) \rightarrow Br_2(aq)$$

Oxidation States: -1 to 0

Reduction Half Reaction

$$Cl_2(aq) \rightarrow 2Cl^-(aq)$$

Oxidation States: 0 to -1

Overview

- Br^- *loses* an electron; it is *oxidized* from Br^- to Br_2; thus, Br^- is the *reducing agent*.
- Cl_2 *gains* one electron; it is *reduced* from Cl_2 to 2 Cl^-; thus, Cl_2 is the *oxidizing agent*.

Exercise 7.5.2: Identify reducing and oxidizing agents

Identify the oxidizing agent and the reducing agent in the following redox reaction:

$$MnO_4^- + SO_3^{2-} \rightarrow Mn^{2+} + SO_4^{2-}$$

Solution

S is the reducing agent and Mn is the oxidizing agent.

Summary

- An oxidizing agent is a substance that causes oxidation by accepting electrons; therefore, it gets reduced.
- A reducing agent is a substance that causes reduction by losing electrons; therefore it gets oxidized.
- Examples of how to identify oxidizing and reducing agents are shown.

Contributors and Attributions

-

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Learning Objectives

- Define *oxidation* and *reduction*.
- Assign oxidation numbers to atoms in simple compounds.
- Recognize a reaction as an oxidation-reduction reaction.

Consider this chemical reaction:

$$\mathrm{Mg\,(s) + Cl_2\,(g) \longrightarrow MgCl_2}$$

The reactants are two electrically neutral elements; they have the same number of electrons as protons. The product, however, is ionic; it is composed of Mg^{2+} and Cl^- ions. Somehow, the individual Mg atoms lose two electrons to make the Mg^{2+} ion, while the Cl atoms gain an electron to become Cl^- ions. This reaction involves the *transfer of electrons* between atoms.

The process of losing and gaining electrons occurs simultaneously. However, mentally we can separate the two processes. **Oxidation** is defined as the loss of one or more electrons by an atom. **Reduction** is defined as the gain of one or more electrons by an atom. So oxidation and reduction always occur together; it is only mentally that we can separate them. Chemical reactions that involve the transfer of electrons are called **oxidation-reduction (or redox) reactions**.

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use **oxidation numbers** to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on four rules. Oxidation numbers are not necessarily equal to the charge on the atom (although sometimes they can be); we must keep the concepts of charge and oxidation numbers separate.

Assigning Oxidation Numbers

The rules for assigning oxidation numbers to atoms are as follows:

1. Atoms in their elemental state are assigned an oxidation number of 0.
2. Atoms in monatomic (i.e., one-atom) ions are assigned an oxidation number equal to their charge. Oxidation numbers are usually written with the sign first, then the magnitude, to differentiate them from charges.
3. In compounds, fluorine is assigned a –1 oxidation number; oxygen is usually assigned a –2 oxidation number (except in peroxide compounds [where it is –1] and in binary compounds with fluorine [where it is positive]); and hydrogen is usually assigned a +1 oxidation number (except when it exists as the hydride ion [H^-], in which case rule 2 prevails).
4. In compounds, all other atoms are assigned an oxidation number so that the sum of the oxidation numbers on all the atoms in the species equals the charge on the species (which is zero if the species is neutral).

Here are some examples for practice. In H_2, both H atoms have an oxidation number of 0 by rule 1. In $MgCl_2$, magnesium has an oxidation number of +2, while chlorine has an oxidation number of –1 by rule 2. In H_2O, the H atoms each have an oxidation number of +1, while the O atom has an oxidation number of –2, even though hydrogen and oxygen do not exist as ions in this compound (rule 3). By contrast, by rule 3, each H atom in hydrogen peroxide (H_2O_2) has an oxidation number of +1, while each O atom has an oxidation number of –1. We can use rule 4 to determine oxidation numbers for the atoms in SO_2. Each O atom has an oxidation number of –2; for the sum of the oxidation numbers to equal the charge on the species (which is zero), the S atom is assigned an oxidation number of +4. Does this mean that the sulfur atom has a 4+ charge on it? No, it means only that the S atom is assigned a +4 oxidation number by our rules of apportioning electrons among the atoms in a compound.

Example 7.5.1

Assign oxidation numbers to the atoms in each substance.

a. $\mathrm{Cl_2}$
b. $\mathrm{GeO_2}$
c. $\mathrm{Ca(NO_3)_2}$

Solution

1. $\mathrm{Cl_2}$ is the elemental form of chlorine. Rule 1 states each atom has an oxidation number of 0.
2. By rule 3, oxygen is normally assigned an oxidation number of –2. For the sum of the oxidation numbers to equal the charge on the species (zero), the Ge atom is assigned an oxidation number of +4.
3. $\mathrm{Ca(NO_3)_2}$ can be separated into two parts: the $\mathrm{Ca^{2+}}$ ion and the $\mathrm{NO_3^-}$ ion. Considering these separately, the $\mathrm{Ca^{2+}}$ ion has an oxidation number of +2 by rule 2. Now consider the $\mathrm{NO_3^-}$ ion. Oxygen is assigned an oxidation number of –2, and there are three of them. According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation x + 3(–2) = –1 where x is the oxidation number of the N atom and the –1 represents the charge on the species. Evaluating for x,

$$\begin{aligned} x + (-6) &= -1 \\ x &= +5 \end{aligned}$$

Thus the oxidation number on the N atom in the $\mathrm{NO_3^-}$ ion is +5.

Exercise 7.5.1: Phosphoric Acid

Assign oxidation numbers to the atoms in H_3PO_4.

Answer available at the end of this section.

All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When an oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidized*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction can also be defined in terms of increasing or decreasing oxidation numbers, respectively.

Example 7.5.2

Identify what is being oxidized and reduced in this redox reaction.

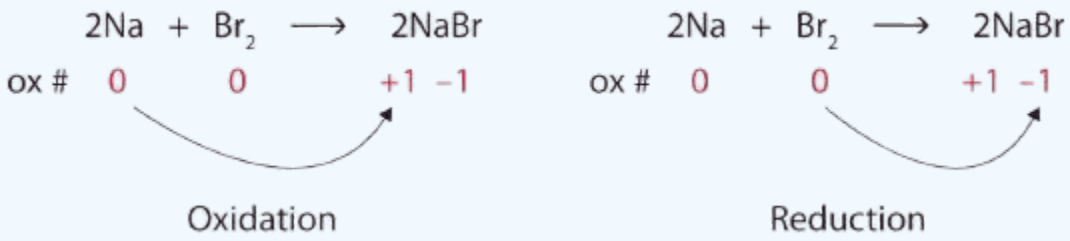
$$\mathrm{2\,Na + Br_2 \longrightarrow 2\,NaBr}$$

Solution

Both reactants are the elemental forms of their atoms, so the Na and Br atoms have oxidation numbers of 0. In the ionic product, the Na^+ ions have an oxidation number of +1, while the Br^- ions have an oxidation number of –1.

$$\overset{0}{2\,Na} + \overset{0}{Br_2} \rightarrow \overset{+1-1}{2\,NaBr}$$

Sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; bromine is decreasing its oxidation number from 0 to –1, so it is being reduced:



Because oxidation numbers are changing, this is a redox reaction. The total number of electrons being lost by sodium (two, one lost from each Na atom) is gained by bromine (two, one gained for each Br atom).

Exercise 7.5.2

Identify what is being oxidized and reduced in this redox reaction.

$$C + O_2 \longrightarrow CO_2$$

Answer available at the end of this section.

Oxidation reactions can become quite complex, as attested by the following redox reaction:

$$6\,H^+(aq) + 2\,\overset{+7}{MnO_4}^-(aq) + 5\,\overset{-1}{H_2O_2}(l) \longrightarrow 2\,\overset{+2}{[redacted]}(aq) + 5\,\overset{0}{O_2}(g) + 8\,H_2O(l)$$

To demonstrate that this is a redox reaction, the oxidation numbers of the species being oxidized and reduced are listed; can you determine what is being oxidized and what is being reduced? This is also an example of a net ionic reaction; spectator ions that do not change oxidation numbers are not displayed in the equation. Eventually, we will need to learn techniques for writing correct (i.e., balanced) redox reactions.

Food and Drink Application: Fortifying Food with Iron

Iron is an essential mineral in our diet; iron-containing compounds like the heme protein in hemoglobin could not function without it. Most biological iron has the form of the Fe^{2+} ion; iron with other oxidation numbers is almost inconsequential in human biology (although the body does contain an enzyme to reduce Fe^{3+} to Fe^{2+}, so Fe^{3+} must have some biological significance, albeit minor). To ensure that we ingest enough iron, many foods are enriched with iron. Although Fe^{2+} compounds are the most logical substances to use, some foods use "reduced iron" as an ingredient (bread and breakfast cereals are the most well-known examples). Reduced iron is simply iron metal; iron is added as a fine metallic powder. The metallic iron is oxidized to Fe^{2+} in the digestive system and then absorbed by the body, but the question remains: Why are we ingesting metallic iron? Why not just use Fe^{2+} salts as an additive?

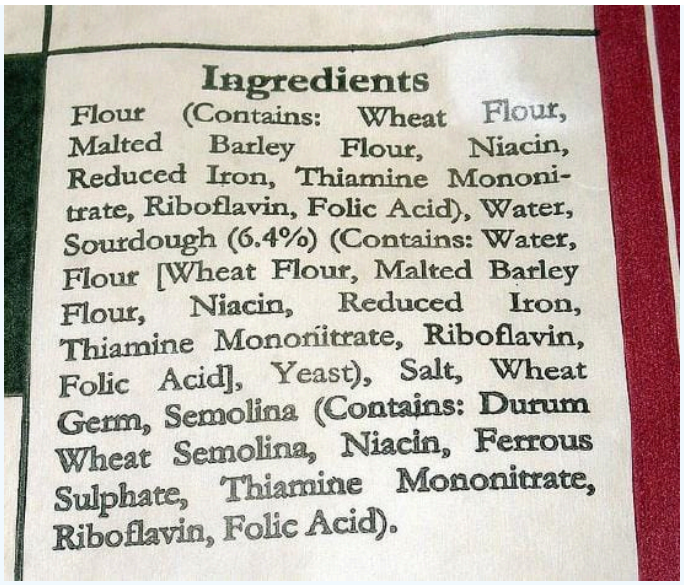


Figure 7.5.1 Ingredients. Many prepared foods list reduced iron in their ingredients list.

Although it is difficult to establish conclusive reasons, a search of scientific and medical literature suggests a few reasons. One reason is that fine iron filings do not affect the taste of the product. The size of the iron powder (several dozen micrometers) is not noticeable when chewing iron-supplemented foods, and the tongue does not detect any changes in flavor that can be detected when using Fe^{2+} salts. Fe^{2+} compounds can affect other properties of foodstuffs during preparation and cooking, like dough pliability, yeast growth, and color. Finally, of the common iron substances that might be used, metallic iron is the least expensive. These factors appear to be among the reasons why metallic iron is the supplement of choice in some foods.

Key Takeaways

- Oxidation-reduction (redox) reactions involve the transfer of electrons from one atom to another.
- Oxidation numbers are used to keep track of electrons in atoms.
- There are rules for assigning oxidation numbers to atoms.
- Oxidation is an increase in oxidation number (loss of electrons); reduction is a decrease in oxidation number (gain of electrons).

Learning Outcomes

- Assign oxidation numbers to free elements or elements in a compound or ion.

Oxidation Numbers

An **oxidation number** *is a positive or negative number that is assigned to an atom to indicate its degree of oxidation or reduction*. The term oxidation state is often used interchangeably with oxidation number. A partial electron transfer is a shift in the electron density near an atom as a result of a change in the other atoms to which it is covalently bonded. That charge shift is based on the relative electronegativities of the atoms involved in the bond.

Overall, the oxidation number of an atom in a molecule is the charge that the atom would have if all polar covalent and ionic bonds resulted in a complete transfer of electrons from the less electronegative atom to the more electronegative one. Oxidation numbers can be assigned using the set of rules outlined below.

1. The oxidation number of an atom in a neutral free element is zero. A free element is considered to be any element in an uncombined state, whether monatomic or polyatomic. For example, the oxidation number of each atom in Fe, Li, $\mathrm{N_2}$, Ar, and

P_4 is zero.
2. The oxidation number of a monatomic (composed of one atom) ion is the same as the charge of the ion. For example, the oxidation numbers of K^+, Se^{2-}, and Au^{3+} are $+1$, -2, and $+3$, respectively.
3. The oxidation number of oxygen in most compounds is -2.
4. The oxidation number of hydrogen in most compounds is $+1$.
5. The oxidation number of fluorine in all compounds is -1. Other halogens usually have an oxidation number of -1 in binary compounds, but can have variable oxidation numbers depending on the bonding environment.
6. In a neutral molecule, the sum of the oxidation numbers of all atoms is zero. For example, in H_2O, the oxidation numbers of H and O are $+1$ and -2, respectively. Because there are two hydrogen atoms in the formula, the sum of all the oxidation numbers in H_2O is $2\,(+1) + 1\,(-2) = 0$.
7. In a polyatomic ion, the sum of the oxidation numbers of all atoms is equal to the overall charge on the ion. For example, in $SO_4{}^{2-}$, the oxidation numbers of S and O are $+6$ and -2, respectively. The sum of all oxidation numbers in the sulfate ion would be $1\,(+6) + 4\,(-2) = -2$, which is the charge of the ion.

An examination of the rules for assigning oxidation numbers reveals that there are many elements for which there are no specific rules, such as nitrogen, sulfur, and chlorine. These elements, as well as some others, can have variable oxidation numbers depending on the other atoms to which they are covalently bonded in a molecular compound. It is useful to analyze a few molecules in order to see the strategy to follow in assigning oxidation numbers to other atoms.

Oxidation numbers for the atoms in a binary ionic compound are easy to assign because they are equal to the charge of the ion (rule 2). In $FeCl_3$, the oxidation number of iron is $+3$, while the oxidation number of chlorine is -1. In Ca_3P_2, the calcium is $+2$, while the phosphorus is -3. This is because an ionic compound is in the form of a crystal lattice that is actually composed of these ions.

Assigning oxidation numbers for molecular compounds is trickier. The key is to remember rule 6: that the sum of all the oxidation numbers for any neutral species must be zero. Make sure to account for any subscripts which appear in the formula. As an example, consider the compound nitric acid, HNO_3. According to rule 4, the oxidation number of hydrogen is $+1$. According to rule 3, the oxidation number of oxygen is -2. There is no rule regarding nitrogen, but its oxidation number can be calculated as follows.

$$1\,(+1) + x + 3\,(-2) = 0, \text{ where } x \text{ is the oxidation number of nitrogen} \tag{7.5.1}$$

$$\text{Solving: } x = 0 - 1 - (-6) = +5 \tag{7.5.2}$$

The oxidation number of the nitrogen atom in HNO_3 is $+5$. Often when assigning oxidation numbers, it is convenient to write it above the symbol within the formula.

$$\overset{+1+5\ -2}{HNO_3} \tag{7.5.3}$$

You may wonder if there are any limits on the value of oxidation numbers. The key point to consider is the octet rule. Since nitrogen has 5 valence electrons, the most that it can "lose" while forming bonds in a molecule is 5, so its highest possible oxidation number is $+5$. Alternatively, it could gain up to 3 electrons, and so its lowest (most negative) possible oxidation number is -3. Similarly, chlorine can have oxidation numbers ranging from -1 to $+7$.

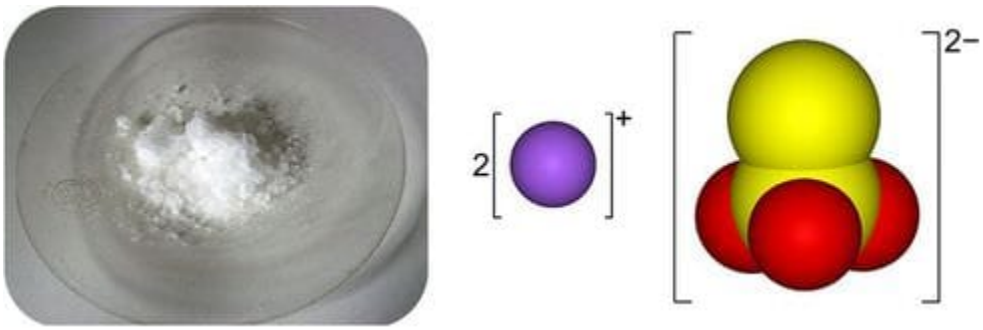


Figure 7.5.1: Sodium thiosulfate is a white crystalline compound (left) composed of two sodium ions (Na^+) for every one thiosulfate ion $(S_2O_3{}^{2-})$ (right).

Now consider the ionic compound sodium thiosulfate, $Na_2S_2O_3$ (Figure 7.5.1). It contains the thiosulfate polyatomic ion, $S_2O_3{}^{2-}$. The sodium is not part of the covalently bonded polyatomic ion, and so its oxidation number is the same as it would be in a binary

ionic compound, $+1$. The sulfur is the atom whose oxidation number is not covered by one of the rules. The oxidation number of sulfur is assigned the variable x in the following calculation. Remember the sum of the oxidation numbers of all the elements must equal zero because $\mathrm{Na_2S_2O_3}$ is a neutral compound.

$$2\,(+1) + 2\,(x) + 3\,(-2) = 0 \tag{7.5.4}$$
$$2 + 2x - 6 = 0 \tag{7.5.5}$$
$$-4 + 2x = 0 \tag{7.5.6}$$
$$2x = +4 \tag{7.5.7}$$
$$x = +2 \tag{7.5.8}$$

Sulfur has an oxidation number of $+2$ in $\mathrm{Na_2S_2O_3}$. Notice how the subscript of 2 for the S atom had to be accounted for by dividing the result of the subtraction by 2. When assigning oxidation numbers, you do so for each individual atom. In the above example, the oxidation number of sulfur could also have been determined by looking at just the thiosulfate ion, $\mathrm{{S_2O_3}^{2-}}$.

$$2\,(x) + 3\,(-2) = -2 \tag{7.5.9}$$
$$2x - 6 = -2 \tag{7.5.10}$$
$$2x = +4 \tag{7.5.11}$$
$$x = +2 \tag{7.5.12}$$

Contributors and Attributions

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- Allison Soult, Ph.D. (Department of Chemistry, University of Kentucky)

Learning Objectives

- Define *oxidation* and *reduction*.
- Assign oxidation numbers to atoms in simple compounds.
- Recognize a reaction as an oxidation-reduction reaction.

Consider this chemical reaction:

$$\mathrm{Mg\,(s) + Cl_2\,(g) \longrightarrow MgCl_2}$$

The reactants are two electrically neutral elements; they have the same number of electrons as protons. The product, however, is ionic; it is composed of $\mathrm{Mg^{2+}}$ and $\mathrm{Cl^-}$ ions. Somehow, the individual Mg atoms lose two electrons to make the $\mathrm{Mg^{2+}}$ ion, while the Cl atoms gain an electron to become $\mathrm{Cl^-}$ ions. This reaction involves the *transfer of electrons* between atoms.

The process of losing and gaining electrons occurs simultaneously. However, mentally we can separate the two processes. **Oxidation** is defined as the loss of one or more electrons by an atom. **Reduction** is defined as the gain of one or more electrons by an atom. So oxidation and reduction always occur together; it is only mentally that we can separate them. Chemical reactions that involve the transfer of electrons are called **oxidation-reduction (or redox) reactions**.

Redox reactions require that we keep track of the electrons assigned to each atom in a chemical reaction. How do we do that? We use **oxidation numbers** to keep track of electrons in atoms. Oxidation numbers are assigned to atoms based on four rules. Oxidation numbers are not necessarily equal to the charge on the atom (although sometimes they can be); we must keep the concepts of charge and oxidation numbers separate.

Assigning Oxidation Numbers

The rules for assigning oxidation numbers to atoms are as follows:

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2. Atoms in monatomic (i.e., one-atom) ions are assigned an oxidation number equal to their charge. Oxidation numbers are usually written with the sign first, then the magnitude, to differentiate them from charges.
3. In compounds, fluorine is assigned a –1 oxidation number; oxygen is usually assigned a –2 oxidation number (except in peroxide compounds [where it is –1] and in binary compounds with fluorine [where it is positive]); and hydrogen is usually assigned a +1 oxidation number (except when it exists as the hydride ion [H^-], in which case rule 2 prevails).
4. In compounds, all other atoms are assigned an oxidation number so that the sum of the oxidation numbers on all the atoms in the species equals the charge on the species (which is zero if the species is neutral).

Here are some examples for practice. In H_2, both H atoms have an oxidation number of 0 by rule 1. In $MgCl_2$, magnesium has an oxidation number of +2, while chlorine has an oxidation number of –1 by rule 2. In H_2O, the H atoms each have an oxidation number of +1, while the O atom has an oxidation number of –2, even though hydrogen and oxygen do not exist as ions in this compound (rule 3). By contrast, by rule 3, each H atom in hydrogen peroxide (H_2O_2) has an oxidation number of +1, while each O atom has an oxidation number of –1. We can use rule 4 to determine oxidation numbers for the atoms in SO_2. Each O atom has an oxidation number of –2; for the sum of the oxidation numbers to equal the charge on the species (which is zero), the S atom is assigned an oxidation number of +4. Does this mean that the sulfur atom has a 4+ charge on it? No, it means only that the S atom is assigned a +4 oxidation number by our rules of apportioning electrons among the atoms in a compound.

Example 7.5.1

Assign oxidation numbers to the atoms in each substance.

a. Cl_2
b. GeO_2
c. $Ca(NO_3)_2$

Solution

1. Cl_2 is the elemental form of chlorine. Rule 1 states each atom has an oxidation number of 0.
2. By rule 3, oxygen is normally assigned an oxidation number of –2. For the sum of the oxidation numbers to equal the charge on the species (zero), the Ge atom is assigned an oxidation number of +4.
3. $Ca(NO_3)_2$ can be separated into two parts: the Ca^{2+} ion and the NO_3^- ion. Considering these separately, the Ca^{2+} ion has an oxidation number of +2 by rule 2. Now consider the NO_3^- ion. Oxygen is assigned an oxidation number of –2, and there are three of them. According to rule 4, the sum of the oxidation numbers on all atoms must equal the charge on the species, so we have the simple algebraic equation x + 3(–2) = –1 where x is the oxidation number of the N atom and the –1 represents the charge on the species. Evaluating for x,

$$x + (-6) = -1$$
$$x = +5$$

Thus the oxidation number on the N atom in the NO_3^- ion is +5.

Exercise 7.5.1: Phosphoric Acid

Assign oxidation numbers to the atoms in H_3PO_4.

Answer available at the end of this section.

All redox reactions occur with a simultaneous change in the oxidation numbers of some atoms. At least two elements must change their oxidation numbers. When an oxidation number of an atom is increased in the course of a redox reaction, that atom is being *oxidized*. When an oxidation number of an atom is decreased in the course of a redox reaction, that atom is being *reduced*. Oxidation and reduction can also be defined in terms of increasing or decreasing oxidation numbers, respectively.

Example 7.5.2

Identify what is being oxidized and reduced in this redox reaction.

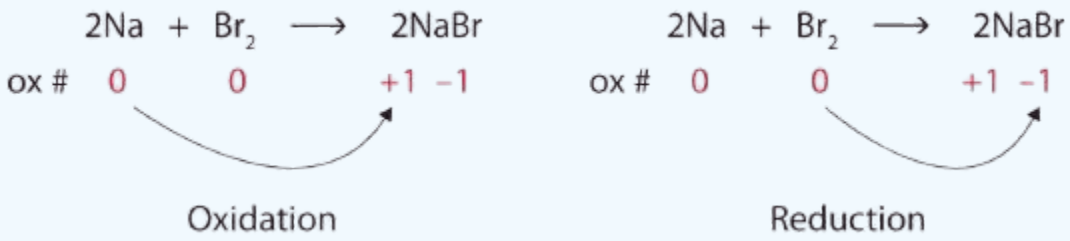
$$2\,\mathrm{Na} + \mathrm{Br_2} \longrightarrow 2\,\mathrm{NaBr}$$

Solution

Both reactants are the elemental forms of their atoms, so the Na and Br atoms have oxidation numbers of 0. In the ionic product, the Na^+ ions have an oxidation number of +1, while the Br^- ions have an oxidation number of –1.

$$2\,\overset{0}{\mathrm{Na}} + \overset{0}{\mathrm{Br_2}} \rightarrow 2\,\overset{+1-1}{\mathrm{NaBr}}$$

Sodium is increasing its oxidation number from 0 to +1, so it is being oxidized; bromine is decreasing its oxidation number from 0 to –1, so it is being reduced:



Because oxidation numbers are changing, this is a redox reaction. The total number of electrons being lost by sodium (two, one lost from each Na atom) is gained by bromine (two, one gained for each Br atom).

Exercise 7.5.2

Identify what is being oxidized and reduced in this redox reaction.

$$\mathrm{C} + \mathrm{O_2} \longrightarrow \mathrm{CO_2}$$

Answer available at the end of this section.

Oxidation reactions can become quite complex, as attested by the following redox reaction:

$$6\,\mathrm{H^+(aq)} + 2\,\overset{+7}{\mathrm{MnO_4^-}}\mathrm{(aq)} + 5\,\overset{-1}{\mathrm{H_2O_2}}\mathrm{(l)} \longrightarrow 2\,\overset{+2}{\blacksquare}\mathrm{(aq)} + 5\,\overset{0}{\mathrm{O_2}}\mathrm{(g)} + 8\,\mathrm{H_2O\,(l)}$$

To demonstrate that this is a redox reaction, the oxidation numbers of the species being oxidized and reduced are listed; can you determine what is being oxidized and what is being reduced? This is also an example of a net ionic reaction; spectator ions that do not change oxidation numbers are not displayed in the equation. Eventually, we will need to learn techniques for writing correct (i.e., balanced) redox reactions.

Food and Drink Application: Fortifying Food with Iron

Iron is an essential mineral in our diet; iron-containing compounds like the heme protein in hemoglobin could not function without it. Most biological iron has the form of the Fe^{2+} ion; iron with other oxidation numbers is almost inconsequential in human biology (although the body does contain an enzyme to reduce Fe^{3+} to Fe^{2+}, so Fe^{3+} must have some biological significance, albeit minor). To ensure that we ingest enough iron, many foods are enriched with iron. Although Fe^{2+} compounds are the most logical substances to use, some foods use "reduced iron" as an ingredient (bread and breakfast

cereals are the most well-known examples). Reduced iron is simply iron metal; iron is added as a fine metallic powder. The metallic iron is oxidized to Fe^{2+} in the digestive system and then absorbed by the body, but the question remains: Why are we ingesting metallic iron? Why not just use Fe^{2+} salts as an additive?

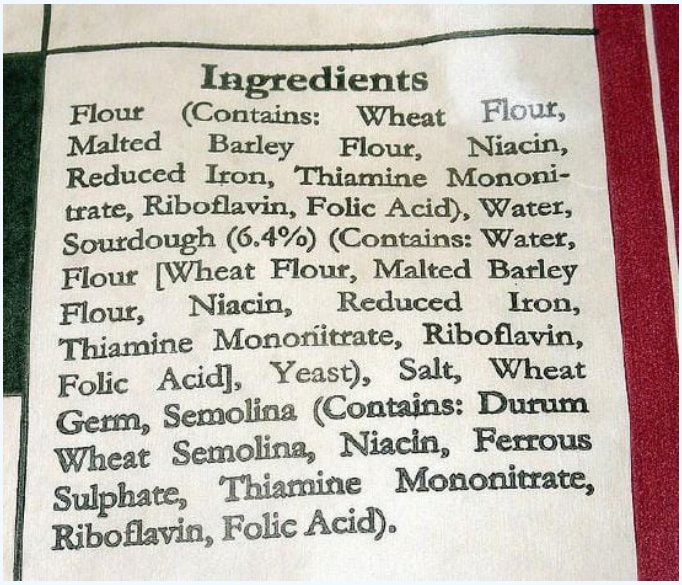


Figure 7.5.1 Ingredients. Many prepared foods list reduced iron in their ingredients list.

Although it is difficult to establish conclusive reasons, a search of scientific and medical literature suggests a few reasons. One reason is that fine iron filings do not affect the taste of the product. The size of the iron powder (several dozen micrometers) is not noticeable when chewing iron-supplemented foods, and the tongue does not detect any changes in flavor that can be detected when using Fe^{2+} salts. Fe^{2+} compounds can affect other properties of foodstuffs during preparation and cooking, like dough pliability, yeast growth, and color. Finally, of the common iron substances that might be used, metallic iron is the least expensive. These factors appear to be among the reasons why metallic iron is the supplement of choice in some foods.

Key Takeaways

- Oxidation-reduction (redox) reactions involve the transfer of electrons from one atom to another.
- Oxidation numbers are used to keep track of electrons in atoms.
- There are rules for assigning oxidation numbers to atoms.
- Oxidation is an increase in oxidation number (loss of electrons); reduction is a decrease in oxidation number (gain of electrons).

Answers

Exercise 7.5.1: Phosphoric Acid

H: +1; O: −2; P: +5

Exercise 7.5.3

C is being oxidized from 0 to +4; O is being reduced from 0 to −2.

Exercise 7.5.1: Phosphoric Acid

H: +1; O: −2; P: +5

Exercise 7.5.2

C is being oxidized from 0 to +4.

O is being reduced from 0 to −2.

Exercise 7.5.1: Phosphoric Acid

H: +1; O: −2; P: +5

Exercise 7.5.2

C is being oxidized from 0 to +4.
O is being reduced from 0 to −2.

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7.6: Applications of Redox Chemistry

Learning Objectives

- Know the occurrence and properties of oxygen.
- Know reactions involving oxygen as an oxidizing agent.

Elemental oxygen is a strong oxidizing agent. It reacts with most other elements and many compounds. Oxygen was isolated by Michael Sendivogius before 1604, but it is commonly believed that the element was discovered independently by Carl Wilhelm Scheele, in Uppsala, in 1773 or earlier, and Joseph Priestley in Wiltshire, in 1774. The name oxygen was coined in 1777 by Antoine Lavoisier, who first recognized oxygen as a chemical element and correctly characterized the role it plays in combustion.

Oxygen: Occurrence and Properties

By mass, oxygen is the third-most abundant element in the universe, after hydrogen and helium. Oxygen is the most abundant element on the earth's crust. About 50% of the mass of the earth's crust consists of oxygen (combined with other elements, principally silicon). Oxygen occurs as O_2 molecules and, to a limited extent, as O_3 (ozone) molecules in air. It forms about 20% of the mass of the air. About 89% of water by mass consists of combined oxygen. In combination with carbon, hydrogen, and nitrogen, oxygen is a large part of plants and animals.

Oxygen is a colorless, odorless, and tasteless gas at ordinary temperatures. It is slightly denser than air. Although it is only slightly soluble in water (49 mL of gas dissolves in 1 L at STP), oxygen's solubility is very important to aquatic life.

Oxygen is essential in combustion processes such as the burning of fuels. All combustion reactions are also examples of redox reactions. A combustion reaction occurs when a substance reacts with oxygen to create heat. One example is the combustion of octane, the principle component of gasoline:

$$2\,C_8H_{18}\,(l) + 25\,O_2\,(g) \longrightarrow 16\,CO_2\,(g) + 18\,H_2O\,(g)$$

Combustion reactions are a major source of energy for modern industry. Plants and animals use the oxygen from the air in respiration. The administration of oxygen-enriched air is an important medical practice when a patient is receiving an inadequate supply of oxygen because of shock, pneumonia, or some other illness.

The chemical industry employs oxygen for oxidizing many substances. A significant amount of oxygen produced commercially is important in the removal of carbon from iron during steel production. Large quantities of pure oxygen are also necessary in metal fabrication and in the cutting and welding of metals with oxyhydrogen and oxyacetylene torches. Liquid oxygen is important to the space industry. It is an oxidizing agent in rocket engines. It is also the source of gaseous oxygen for life support in space.

As we know, oxygen is very important to life. The energy required for the maintenance of normal body functions in human beings and in other organisms comes from the slow oxidation of chemical compounds. Oxygen is the final oxidizing agent in these reactions. In humans, oxygen passes from the lungs into the blood, where it combines with hemoglobin, producing oxyhemoglobin. In this form, blood transports the oxygen to tissues, where it is transferred to the tissues. The ultimate products are carbon dioxide and water. The blood carries the carbon dioxide through the veins to the lungs, where the blood releases the carbon dioxide and collects another supply of oxygen. Digestion and assimilation of food regenerate the materials consumed by oxidation in the body; the energy liberated is the same as if the food burned outside the body.

Oxygen: Reactions with Other Elements

Oxygen reacts directly at room temperature or at elevated temperatures with all other elements except the noble gases, the halogens, and few second- and third-row transition metals of low reactivity (those with higher reduction potentials than copper). Rust is an example of the reaction of oxygen with iron. The more active metals form peroxides or superoxides. Less active metals and the nonmetals give oxides. Two examples of these reactions are:

$$2\,\mathrm{Mg\,(s)} + \mathrm{O_2\,(g)} \longrightarrow 2\,\mathrm{MgO\,(s)}$$

$$\mathrm{P_4(s)}^{+}\ 5\,\mathrm{O_2\,(g)} \longrightarrow \mathrm{P_4O_{10}\,(s)}$$

The oxides of halogens, at least one of the noble gases, and metals with higher reduction potentials than copper do not form by the direct action of the elements with oxygen.

Oxygen: Reaction with Compounds

Elemental oxygen also reacts with some compounds. If it is possible to oxidize any of the elements in a given compound, further oxidation by oxygen can occur. For example, hydrogen sulfide, H_2S, contains sulfur with an oxidation state of 2–. Because the sulfur does not exhibit its maximum oxidation state, we would expect H_2S to react with oxygen. It does, yielding water and sulfur dioxide. The reaction is:

$$2\,\mathrm{H_2S}(g)+3\,\mathrm{O_2}(g) \longrightarrow 2\,\mathrm{H_2O}(l)+2\,\mathrm{SO_2}(g)$$

It is also possible to oxidize oxides such as CO and P_4O_6 that contain an element with a lower oxidation state. The ease with which elemental oxygen picks up electrons is mirrored by the difficulty of removing electrons from oxygen in most oxides. Of the elements, only the very reactive fluorine can oxidize oxides to form oxygen gas.

Most nonmetals react with oxygen to form nonmetal oxides. Depending on the available oxidation states for the element, a variety of oxides might form. Fluorine will combine with oxygen to form fluorides such as OF_2, where the oxygen has a 2+-oxidation state.

Sulfur Oxygen Compounds

The two common oxides of sulfur are sulfur dioxide, SO_2, and sulfur trioxide, SO_3. The odor of burning sulfur comes from sulfur dioxide. Sulfur dioxide, occurs in volcanic gases Figure 7.6.1 and in the atmosphere near industrial plants that burn fuel containing sulfur compounds.



Figure 7.6.1: Mount Mayon volcano (Philippines) erupting in the early morning of January 23, 2018. (CC BY-SA 4.0; Ezra Acayan via Wikipedia)

Ozone: Another Form of Oxygen

When dry oxygen is passed between two electrically charged plates, ozone (O_3, illustrated in Figure 7.6.2), an allotrope of oxygen possessing a distinctive odor, forms. The formation of ozone from oxygen is an endothermic reaction, in which the energy comes from an electrical discharge, heat, or ultraviolet light:

$$3\,O_2(g) \xrightarrow{\text{electric discharge}} 2\,O_3(g)$$

The sharp odor associated with sparking electrical equipment is due, in part, to ozone.

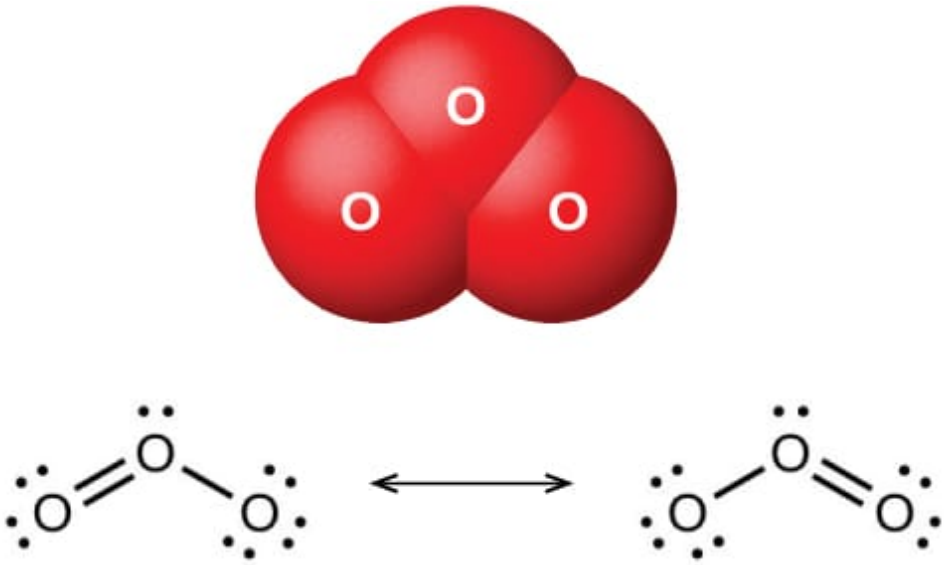


Figure 7.6.2 The image shows the bent ozone (O3) molecule and the resonance structures necessary to describe its bonding.

Ozone forms naturally in the upper atmosphere by the action of ultraviolet light from the sun on the oxygen there. Most atmospheric ozone occurs in the stratosphere, a layer of the atmosphere extending from about 10 to 50 kilometers above the earth's surface. This ozone acts as a barrier to harmful ultraviolet light from the sun by absorbing it via a chemical decomposition reaction:

$$O_3(g) \xrightarrow{\text{ultraviolet light}} O(g) + O_2(g)$$

The reactive oxygen atoms recombine with molecular oxygen to complete the ozone cycle. The presence of stratospheric ozone decreases the frequency of skin cancer and other damaging effects of ultraviolet radiation. It has been clearly demonstrated that chlorofluorocarbons, CFCs (known commercially as Freons), which were present as aerosol propellants in spray cans and as refrigerants, caused depletion of ozone in the stratosphere. This occurred because ultraviolet light also causes CFCs to decompose, producing atomic chlorine. The chlorine atoms react with ozone molecules, resulting in a net removal of O_3 molecules from stratosphere. This process is explored in detail in our coverage of chemical kinetics. There is a worldwide effort to reduce the amount of CFCs used commercially, and the ozone hole is already beginning to decrease in size as atmospheric concentrations of atomic chlorine decrease. While ozone in the stratosphere helps protect us, ozone in the troposphere is a problem. This ozone is a toxic component of photochemical smog.

The uses of ozone depend on its reactivity with other substances. It can be used as a bleaching agent for oils, waxes, fabrics, and starch: It oxidizes the colored compounds in these substances to colorless compounds. It is an alternative to chlorine as a disinfectant for water.

Low level ozone (or tropospheric ozone) is an atmospheric pollutant.[42] It is not emitted directly by car engines or by industrial operations, but formed by the reaction of sunlight on air containing hydrocarbons and nitrogen oxides that react to form ozone directly at the source of the pollution or many kilometers down wind.

Ozone reacts directly with some hydrocarbons such as aldehydes and thus begins their removal from the air, but the products are themselves key components of smog.

Ozone can be used to remove iron and manganese from water, forming a precipitate which can be filtered:

$$2Fe^{2+} + O_3 + 5H_2O \rightarrow 2Fe(OH)_3(s) + O_2 + 4H^+$$

$$2Mn^{2+} + 2O_3 + 4H_2O \rightarrow 2MnO(OH)_2(s) + 2O_2 + 4H^+$$

Ozone will also oxidize dissolved hydrogen sulfide in water to sulfurous acid:

$$3O_3 + H_2S \rightarrow H_2SO_3 + 3O_2$$

These three reactions are central in the use of ozone based well water treatment.

Ozone will also detoxify cyanides by converting them to cyanates.

$$CN^- + O_3 \rightarrow CNO^- + O_2$$

Ozone will also completely decompose urea:

$$(NH_2)_2CO + O_3 \rightarrow N_2 + CO_2 + 2H_2O$$

Other Common Oxidizing Agents

Chlorine is usually used (in the form of hypochlorous acid) to kill bacteria and other microbes in drinking water supplies and public swimming pools. In most private swimming pools, chlorine itself is not used, but rather sodium hypochlorite, formed from chlorine and sodium hydroxide, or solid tablets of chlorinated isocyanurates. The drawback of using chlorine in swimming pools is that the chlorine reacts with the proteins in human hair and skin. The distinctive 'chlorine aroma' associated with swimming pools is not the result of chlorine itself, but of chloramine, a chemical compound produced by the reaction of free dissolved chlorine with amines in organic substances. As a disinfectant in water, chlorine is more than three times as effective against Escherichia coli as bromine, and more than six times as effective as iodine. Increasingly, chloramine itself is being directly added to drinking water for purposes of disinfection, a process known as chloramination.

It is often impractical to store and use poisonous chlorine gas for water treatment, so alternative methods of adding chlorine are used. These include hypochlorite solutions, which gradually release chlorine into the water, and compounds like sodium dichloro-s-triazinetrione (dihydrate or anhydrous), sometimes referred to as "dichlor", and trichloro-s-triazinetrione, sometimes referred to as "trichlor". These compounds are stable while solid and may be used in powdered, granular, or tablet form. When added in small amounts to pool water or industrial water systems, the chlorine atoms hydrolyze from the rest of the molecule forming hypochlorous acid (HOCl), which acts as a general biocide, killing germs, micro-organisms, algae, and so on.

Breathalizers

The breathalyzer is a redox reaction. When the potassium dichromate reacts with ethanol it loses an oxygen atom (gets reduced), going from the orange dichromate to the green chromium sulfate. At the same time dichromate is being reduced, ethanol gains an oxygen atom (gets oxidized), forming acetic acid. The sulfuric acid helps to remove the ethanol from the exhaled air into the test solution and also provides the necessary acidic conditions.

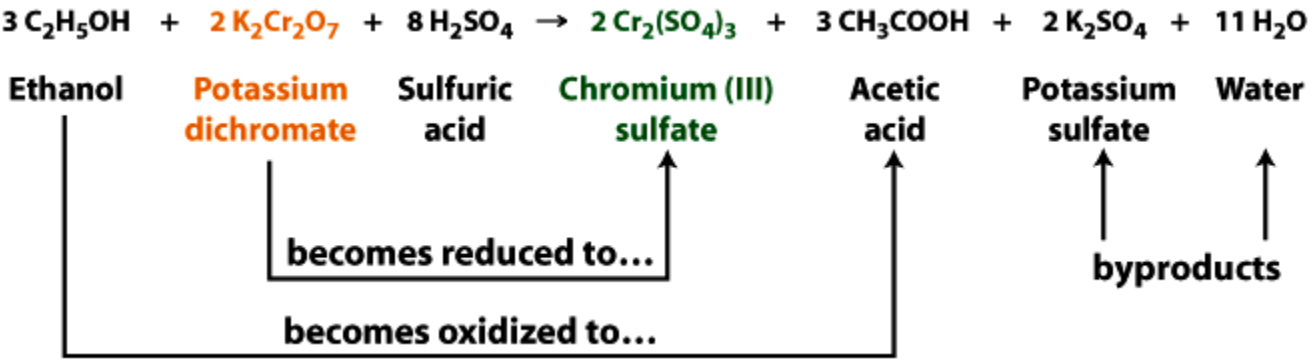


Figure 7.6.3 In this reaction, the chromium atom is reduced from Cr^{6+} to Cr^{3+}, and the ethanol is oxidized to acetic acid.

In this reaction, the chromium atom is reduced from Cr^{6+} to Cr^{3+}, and the ethanol is oxidized to acetic acid.

Hydrogen peroxide can be used for the sterilization of various surfaces, including surgical tools and may be deployed as a vapor (VHP) for room sterilization. H_2O_2 demonstrates broad-spectrum efficacy against viruses, bacteria, yeasts, and bacterial spores. In general, greater activity is seen against Gram-positive than Gram-negative bacteria; however, the presence of catalase or other

peroxidases in these organisms can increase tolerance in the presence of lower concentrations. Higher concentrations of $\mathrm{H_2O_2}$ (10 to 30%) and longer contact times are required for sporicidal activity.

Hydrogen peroxide is seen as an environmentally safe alternative to chlorine-based bleaches, as it degrades to form oxygen and water and it is generally recognized as safe as an antimicrobial agent by the U.S. Food and Drug Administration (FDA).

Historically hydrogen peroxide was used for disinfecting wounds, partly because of its low cost and prompt availability compared to other antiseptics. It is now thought to inhibit healing and to induce scarring because it destroys newly formed skin cells. Only a very low concentration of

$\mathrm{H_2O_2}$ can induce healing, and only if not repeatedly applied. Surgical use can lead to gas embolism formation. Despite this, it is still used for wound treatment in many countries but is also prevalent as a major first aid antiseptic in the United States

Diluted $\mathrm{H_2O_2}$ (between 1.9% and 12%) mixed with aqueous ammonia is used to bleach human hair. The chemical's bleaching property lends its name to the phrase "peroxide blonde". Hydrogen peroxide is also used for tooth whitening. It can be found in most whitening toothpastes.



Figure 7.6.3: Skin shortly after exposure to 35% $\mathrm{H_2O_2}$. (Public Domain; Olli Niemitalo via Wikipedia)

Almost all applications of potassium permanganate ($\mathrm{KMnO_4}$) exploit its oxidizing properties. As a strong oxidant that does not generate toxic byproducts, $\mathrm{KMnO_4}$ has many niche uses. Potassium permanganate is used for a number of skin conditions. This includes fungal infections of the foot, impetigo, pemphigus, superficial wounds, dermatitis, and tropical ulcers. It is on the WHO Model List of Essential Medicines, the most important medications needed in a basic health system. Potassium permanganate is used extensively in the water treatment industry.

Benzoyl peroxide (BPO) is a medication and industrial chemical. As a medication, it is used to treat mild to moderate acne. For more severe cases, it may be used with other treatments. Some versions are sold mixed with antibiotics such as clindamycin. Other uses include bleaching flour, hair bleaching, teeth whitening, and textile bleaching.[5] It is also used in the plastic industry

Sodium hypochlorite is most often encountered as a pale greenish-yellow dilute solution commonly known as liquid bleach or simply bleach, a household chemical widely used (since the 18th century) as a disinfectant or a bleaching agent. The compound in solution is unstable and easily decomposes, liberating chlorine, which is the active principle of such products. Indeed, sodium hypochlorite is the oldest and still most important chlorine-based bleach.

Warning

While sodium hypochlorite is non-toxic, its corrosive properties, common availability, and reaction products make it a significant safety risk. In particular, mixing liquid bleach with other cleaning products, such as acids or ammonia, may produce toxic fumes.

Summary

- Oxygen is one of the most reactive elements. This reactivity, coupled with its abundance, makes the chemistry of oxygen very rich and well understood.
- Many compounds of the representative metals and non metals react with oxygen to form oxides.
- Other than oxygen, ozone, chlorine, potassium dichromate, hydrogen peroxide, potassium peroxide, benzoyl peroxide, sodium hypochlorite are common oxidizing agents mentioned in this section, that have many beneficial uses.

Contributors and Attributions

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

- Wikipedia

- Marisa Alviar-Agnew (Sacramento City College)

Learning Objectives

- Know the occurrence and properties of hydrogen.
- Know reactions involving common reducing agents such as carbon, hydrogen, and antioxidants.

Metal Extraction

Smelting is a process of applying heat to ore in order to extract a base metal. It uses heat and a chemical reducing agent to decompose the ore, leaving the metal base behind. It is used to extract many metals from their ores, including silver, iron, copper, and other base metals. The reducing agent is commonly a source of carbon, such as coke—or, in earlier times, charcoal. **Coke** is a grey, hard, and porous fuel with a high carbon content and few impurities, made by heating coal or oil in the absence of air.

Coke is used in smelting iron ore for example in a blast furnace. The reaction can be represented as follows:

$$2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$$

Antioxidants

In food chemistry, the substances known as antioxidants are reducing agents. Ascorbic acid (vitamin C; $\mathrm{C_6H_8O_6}$) is thought to retard potentially damaging oxidation of living cells. In the process, it is oxidized to dehydroascorbic acid ($\mathrm{C_6H_6O_6}$). In the stomach, ascorbic acid reduces the nitrite ion ($\mathrm{NO_2}^-$) to nitric oxide (NO):

$$\mathrm{C_6H_8O_6 + 2\,H^+ + 2\,NO_2^-} \rightarrow \mathrm{C_6H_6O_6 + 2\,H_2O + 2\,NO} \tag{7.6.1}$$

If reaction in Equation 7.6.1 did not occur, nitrite ions from foods would oxidize the iron in hemoglobin, destroying its ability to carry oxygen.

Tocopherol (vitamin E) is also an antioxidant. In the body, vitamin E is thought to act by scavenging harmful by-products of

metabolism, such as the highly reactive molecular fragments called free radicals. In foods, vitamin E acts to prevent fats from being oxidized and thus becoming rancid. Vitamin C is also a good antioxidant (Figure 7.6.2).



Figure 7.6.2: Citrus Fruits. Citrus fruits, such as oranges, lemons, and limes, are good sources of vitamin C, which is an antioxidant. Wedges of pink grapefruit, lime, and lemon, and a half orange (clockwise from top). from Wikipedia.

Hydrogen as a Reducing Agent

Large quantities of H_2 are needed in the petroleum and chemical industries. The largest application of H_2 is for the processing ("upgrading") of fossil fuels, and in the production of ammonia. Mass production of Ammonia mostly uses the Haber–Bosch process, reacting hydrogen (H_2) and nitrogen (N_2) at a moderately-elevated temperature (450 °C) and high pressure (100 standard atmospheres (10,000 kPa)):

$$3H_2(g) + N_2(g) \rightarrow 2NH_3(g)$$

H_2 has several other important uses. H_2 is used as a hydrogenating agent (Chapter 17), particularly in increasing the level of saturation of unsaturated fats and oils (found in items such as margarine), and in the production of methanol. It is similarly the source of hydrogen in the manufacture of hydrochloric acid. H_2 is also used as a reducing agent of metallic ores.

Hydrogen gas is highly flammable producing a large amount of heat when it reacts with oxygen gas as shown below.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l) + 572kJ$$

Hydrogen gas forms explosive mixtures with air in concentrations from 4–74% and with chlorine at 5–95%. The explosive reactions may be triggered by spark, heat, or sunlight.

A Closer Look at Hydrogen

Hydrogen is among the ten most abundant elements on the planet, but very little is found in elemental form due to its low density and reactivity. Much of the terrestrial hydrogen is locked up in water molecules and organic compounds like hydrocarbons.

Hydrogen is the fuel for reactions of the Sun and other stars (fusion reactions). Hydrogen is the lightest and most abundant element in the universe. About 70%- 75% of the universe is composed of hydrogen by mass. All stars are essentially large masses of hydrogen gas that produce enormous amounts of energy through the fusion of hydrogen atoms at their dense cores. In smaller stars, hydrogen atoms collided and fused to form helium and other light elements like nitrogen and carbon(essential for life). In the larger stars, fusion produces the lighter and heavier elements like calcium, oxygen, and silicon.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

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On Earth, hydrogen is mostly found in association with oxygen; its most abundant form being water (H_2O). Hydrogen is only .9% by mass and 15% by volume abundant on the earth, despite water covering about 70% of the planet. Because hydrogen is so light, there is only 0.5 ppm (parts per million) in the atmosphere, which is a good thing considering it is EXTREMELY flammable.

Hydrogen gas can be prepared by reacting a dilute strong acid like hydrochloric acids with an active metal. The metal becomes oxides, while the H^+ (from the acid) is reduced to hydrogen gas. This method is only practical for producing small amounts of hydrogen in the lab, but is much too costly for industrial production:

$$Zn_{(s)} + 2H^+_{(aq)} \rightarrow Zn^{2+}_{(aq)} + H_{2(g)}$$

The purest form of $H_2(g)$ can come from electrolysis of $H_2O(l)$, the most common hydrogen compound on this plant. This method is also not commercially viable because it requires a significant amount of energy ($\Delta H = 572 \ \ kJ$):

$$2H_2O_{(l)} \rightarrow 2H_{2(g)} + O_{2(g)}$$

\(H_2O\) is the most abundant form of hydrogen on the planet, so it seems logical to try to extract hydrogen from water without electrolysis of water. To do so, we must reduce the hydrogen with +1 oxidation state to hydrogen with 0 oxidation state (in hydrogen gas).

Three commonly used reducing agents are carbon (in coke or coal), carbon monoxide, and methane. These react with water vapor form $H_2(g)$:

$$C_{(s)} + 2H_2O_{(g)} \rightarrow CO(g) + H_{2(g)}$$

$$CO_{(g)} + 2H_2O_{(g)} \rightarrow CO2 + H_{2(g)}$$

Reforming of Methane:

$$CH_{4(g)} + H_2O_{(g)} \rightarrow CO(g) + 3H_{2(g)}$$

These three methods are most industrially feasible (cost effective) methods of producing $H_2(g)$.

Summary

- Common reducing agents include carbon (in the form of coke or coal), hydrogen gas, as well as those substances referred to in the food chemistry as antioxidants (e.g. ascorbic acid and vitamin E).
- Various reactions involving reducing agents were given.

Contributors and Attributions

- Libretext:The Basics of GOB Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Wikipedia
- Ridhi Sachdev (UC Davis)
- Stephen R. Marsden

Learning Objective

- Know the overall reactions involved in respiration and photosynthesis.

Oxidation and reduction reactions are especially crucial in biological processes such as cellular respiration and photosynthesis.

Cellular respiration is a set of metabolic reactions and processes that take place in the cells of organisms to convert biochemical energy from nutrients into adenosine triphosphate (ATP), and then release waste products. Cellular respiration is considered an exothermic redox reaction which releases heat. The overall reaction occurs in a series of biochemical steps, most of which are redox reactions themselves. Although cellular respiration is technically a combustion reaction, it clearly does not resemble one when it occurs in a living cell because of the slow release of energy from the series of reactions.

Nutrients that are commonly used by animal and plant cells in respiration include sugar, amino acids and fatty acids, and the most common oxidizing agent (electron acceptor) is molecular oxygen (O_2).

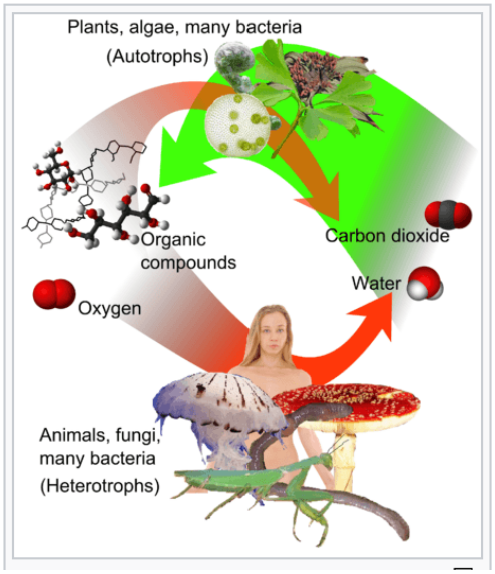


Figure 7.6.1 Aerobic respiration (red arrows) is the main means by which both fungi and animals utilize chemical energy in the form of organic compounds that were previously created through photosynthesis (green arrow). (CC-BY-SA 2.0, Mikael Häggström, Wikimedia)

The reaction for aerobic respiration is essentially the reverse of photosynthesis and is simplified as:

$$C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O + 2880kJ/mol$$

Green plants carry out the redox reaction that makes possible almost all life on Earth. They do this through a process called **photosynthesis**, in which carbon dioxide and water are converted to glucose ($C_6H_{12}O_6$). The synthesis of glucose requires a variety of proteins called enzymes and a green pigment called chlorophyll that converts sunlight into chemical energy

A simplified overall formula for photosynthesis is:

$$6CO_2 + 6H_2O + photons \rightarrow C_6H_{12}O_6 + 6O_2$$

or simply

carbon dioxide + water + sunlight → glucose + oxygen gas

In this reaction, carbon dioxide is reduced to glucose, and water is oxidized to oxygen gas. Other reactions convert the glucose to more complex carbohydrates, plant proteins, and oils.

Summary

- Redox reactions are common in organic and biological chemistry, including aerobic respiration, and photosynthesis.

Contributors and Attributions

- Libretext: Basics of GOB Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Wikipedia

Learning Objectives

- Understand the process of corrosion.
- Understand the process of explosion.

Corrosion is a galvanic process by which metals deteriorate through oxidation—usually but not always to their oxides. For example, when exposed to air, iron rusts, silver tarnishes, and copper and brass acquire a bluish-green surface called a patina. Of the various metals subject to corrosion, iron is by far the most important commercially. An estimated $100 billion per year is spent in the United States alone to replace iron-containing objects destroyed by corrosion.

Explosions can occur in nature due to a large influx of energy. Most natural explosions arise from volcanic processes of various sorts. Chemical explosions, usually involve a rapid and violent oxidation reaction that produces large amounts of hot gas. Gunpowder was the first explosive to be discovered and put to use. Other notable early developments in chemical explosive technology were Frederick Augustus Abel's development of nitrocellulose in 1865 and Alfred Nobel's invention of dynamite in 1866. Chemical explosions (both intentional and accidental) are often initiated by an electric spark or flame in the presence of oxygen. Accidental explosions may occur in fuel tanks, rocket engines, etc.

The Rusting of Iron

Perhaps the most familiar example of corrosion is the formation of rust on iron. Iron will rust when it is exposed to oxygen and water. The main steps in the rusting of iron appear to involve the following. Once exposed to the atmosphere, iron rapidly oxidizes.

$$\mathrm{Fe\,(s)} \longrightarrow \mathrm{Fe^{2+}\,(aq)} + 2\,\mathrm{e^-} \qquad \text{(cathode)}$$

The electrons reduce oxygen in the air in acidic solutions.

$$\mathrm{O_2\,(g)} + 4\,\mathrm{H^+\,(aq)} + 4\,\mathrm{e^-} \longrightarrow 2\,\mathrm{H_2O\,(l)} \qquad \text{(anode)}$$

$$2\,Fe\,(s) + O_2\,(g) + 4\,H^+\,(aq) \longrightarrow 2\,Fe^{2+}\,(aq) + 2\,H_2O\,(l) \qquad \text{(over all)}$$

What we call rust is hydrated iron(III) oxide, which forms when iron(II) ions react further with oxygen.

$$4\,Fe^{2+}\,(aq) + O_2\,(g) + (4 + 2\,x)H_2O\,(l) \longrightarrow 2\,Fe_2O_3\ \cdot x\,H_2O\,(s) + 8\,H^+\,(aq)$$

The number of water molecules is variable, so it is represented by *x*. Unlike the patina on copper, the formation of rust does not create a protective layer and so corrosion of the iron continues as the rust flakes off and exposes fresh iron to the atmosphere.

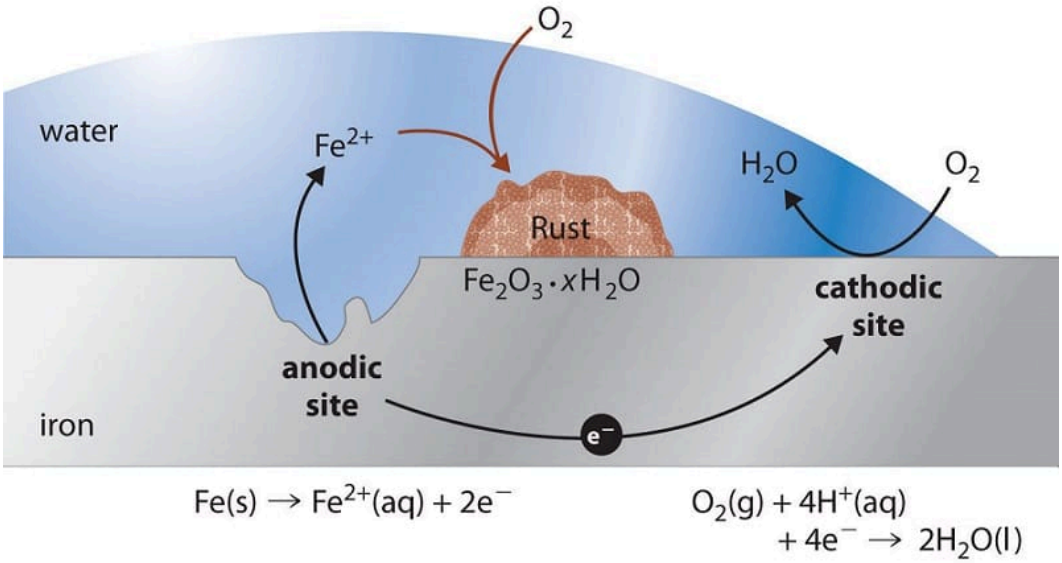


Figure 7.6.1 Rust, the Result of Corrosion of Metallic Iron. Iron is oxidized to Fe2+(aq) at an anodic site on the surface of the iron, which is often an impurity or a lattice defect. Oxygen is reduced to water at a different site on the surface of the iron, which acts as the cathode. Electrons are transferred from the anode to the cathode through the electrically conductive metal.

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$$2\,Cu\,(s) + \tfrac{1}{2}\,O_2\,(g) \rightarrow \underset{\text{red}}{Cu_2O\,(s)}$$

$$Cu_2O\,(s) + \tfrac{1}{2}\,O_2\,(g) \rightarrow 2\,\underset{\text{black}}{CuO\,(s)}$$

Coal, which was often high in sulfur, was burned extensively in the early part of the last century. As a result, sulfur trioxide, carbon dioxide, and water all reacted with the CuO.

$$2\,CuO\,(s) + CO_2\,(g) + H_2O\,(l) \rightarrow \underset{\text{green}}{Cu_2CO_3(OH)_2\,(s)}$$

$$3\,\mathrm{CuO\,(s)} + 2\,\mathrm{CO_2\,(g)} + \mathrm{H_2O\,(l)} \rightarrow \underset{\text{blue}}{\mathrm{Cu_2(CO_3)_2(OH)_2\,(s)}}$$

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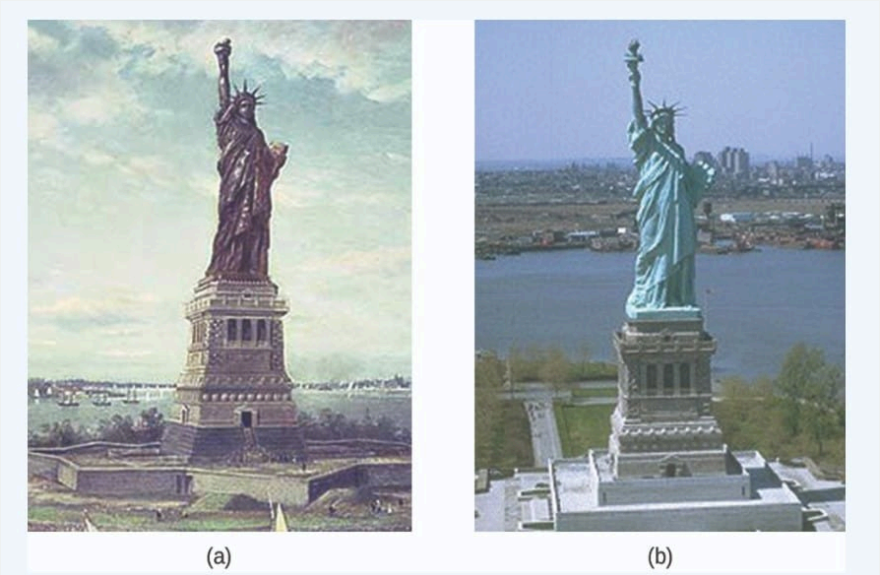


Figure 7.6.2: (a) The Statue of Liberty is covered with a copper skin, and was originally brown, as shown in this painting. (b) Exposure to the elements has resulted in the formation of the blue-green patina seen today.

Silver Tarnishes

Even though coinage metals are known to be resistant to oxygen in air, silver will tarnish due to $\mathrm{H_2S\,(g)}$

$$2\,\mathrm{Ag\,(s)} + \mathrm{H_2S\,(g)} \longrightarrow \mathrm{Ag_2S\,(s)} + \mathrm{H_2\,(g)}$$

Explosive Reactions

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Although blocks of TNT are available in various sizes (e.g. 250 g, 500 g, 1,000 g), it is more commonly encountered in synergistic explosive blends comprising a variable percentage of TNT plus other ingredients

Upon detonation, TNT decomposes as follows:

$$2\,\mathrm{C_7H_5N_3O_6} \longrightarrow 3\,\mathrm{N_2} + 5\,\mathrm{H_2O} + 7\,\mathrm{CO} + 7\,\mathrm{C}$$

$$2\,\mathrm{C_7H_5N_3O_6} \longrightarrow 3\,\mathrm{N_2} + 5\,\mathrm{H_2} + 12\,\mathrm{CO} + 2\,\mathrm{C}$$

The reaction is exothermic but has a high activation energy in the gas phase (~62 kcal/mol). The condensed phases (solid or liquid) show markedly lower activation energies of roughly 35 kcal/mol due to unique bimolecular decomposition routes at elevated densities. Because of the production of carbon, TNT explosions have a sooty appearance. Because TNT has an excess of carbon, explosive mixtures with oxygen-rich compounds can yield more energy per kilogram than TNT alone. During the 20th century, amatol, a mixture of TNT with ammonium nitrate was a widely used military explosive.

TNT can be detonated with a high velocity initiator or by efficient concussion. For many years, TNT used to be the reference point for the Figure of Insensitivity. TNT had a rating of exactly 100 on the "F of I" scale. The reference has since been changed to a more sensitive explosive called RDX, which has an F of I rating of 80.

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The chemistry of ANFO detonation is the reaction of ammonium nitrate with a long-chain alkane ($\mathrm{C_nH_{2n+2}}$) to form nitrogen, carbon dioxide, and water.

$$\mathrm{NH_4NO_3 + C_nH_{2n+2} \longrightarrow \mathit{x}\,N_2 + \mathit{y}\,H_2O + \mathit{z}\,CO_2}$$

In an ideal stoichiometrically balanced reaction, ANFO is composed of about 94.3% AN and 5.7% FO by \(\ce{weight. In practice, a slight excess of fuel oil is added, as underdosing results in reduced performance while overdosing merely results in more post-blast fumes. When detonation conditions are optimal, the aforementioned gases are the only products. In practical use, such conditions are impossible to attain, and blasts produce moderate amounts of toxic gases such as carbon monoxide and nitrogen oxides (NOx).

$$\mathrm{NH_4NO_3 + C_nH_{2n+2} \longrightarrow \mathit{v}\,CO + \mathit{w}\,NOx + \mathit{y}\,H_2O + \mathit{z}\,CO_2}$$

ANFO has found wide use in coal mining, quarrying, metal mining, and civil construction in applications where its low cost and ease of use may outweigh the benefits of other explosives, such as water resistance, oxygen balance, higher detonation velocity, or performance in small-diameter columns. ANFO is also widely used in avalanche hazard mitigation.

Summary

- Corrosion and explosion are two major redox processes.
- Examples of corrosion (metal deterioration) and explosion (using TNT and ANFO) are given.

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- Marisa Alviar-Agnew (Sacramento City College)

- TextMap: Chemistry-A Central Science (Brown et al.)
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Learning Objectives

- Understand the process of corrosion.
- Understand the process of explosion.

Corrosion is a galvanic process by which metals deteriorate through oxidation—usually but not always to their oxides. For example, when exposed to air, iron rusts, silver tarnishes, and copper and brass acquire a bluish-green surface called a patina.

Of the various metals subject to corrosion, iron is by far the most important commercially. An estimated $100 billion per year is spent in the United States alone to replace iron-containing objects destroyed by corrosion.

Explosions can occur in nature due to a large influx of energy. Most natural explosions arise from volcanic processes of various sorts. Chemical explosions, usually involve a rapid and violent oxidation reaction that produces large amounts of hot gas. Gunpowder was the first explosive to be discovered and put to use. Other notable early developments in chemical explosive technology were Frederick Augustus Abel's development of nitrocellulose in 1865 and Alfred Nobel's invention of dynamite in 1866. Chemical explosions (both intentional and accidental) are often initiated by an electric spark or flame in the presence of oxygen. Accidental explosions may occur in fuel tanks, rocket engines, etc.

The Rusting of Iron

Perhaps the most familiar example of corrosion is the formation of rust on iron. Iron will rust when it is exposed to oxygen and water. The main steps in the rusting of iron appear to involve the following. Once exposed to the atmosphere, iron rapidly oxidizes.

$$\mathrm{Fe\,(s)} \longrightarrow \mathrm{Fe^{2+}\,(aq)} + 2\,\mathrm{e^-} \qquad \text{(cathode)}$$

The electrons reduce oxygen in the air in acidic solutions.

$$\mathrm{O_2\,(g)} + 4\,\mathrm{H^+\,(aq)} + 4\,\mathrm{e^-} \longrightarrow 2\,\mathrm{H_2O\,(l)} \qquad \text{(anode)}$$

$$2\,\mathrm{Fe\,(s)} + \mathrm{O_2\,(g)} + 4\,\mathrm{H^+\,(aq)} \longrightarrow 2\,\mathrm{Fe^{2+}\,(aq)} + 2\,\mathrm{H_2O\,(l)} \qquad \text{(over all)}$$

What we call rust is hydrated iron(III) oxide, which forms when iron(II) ions react further with oxygen.

$$4\,\mathrm{Fe^{2+}\,(aq)} + \mathrm{O_2\,(g)} + (4+2\,\mathrm{x})\mathrm{H_2O\,(l)} \longrightarrow 2\,\mathrm{Fe_2O_3\ \cdot}\,x\,\mathrm{H_2O\,(s)} + 8\,\mathrm{H^+\,(aq)}$$

The number of water molecules is variable, so it is represented by *x*. Unlike the patina on copper, the formation of rust does not create a protective layer and so corrosion of the iron continues as the rust flakes off and exposes fresh iron to the atmosphere.

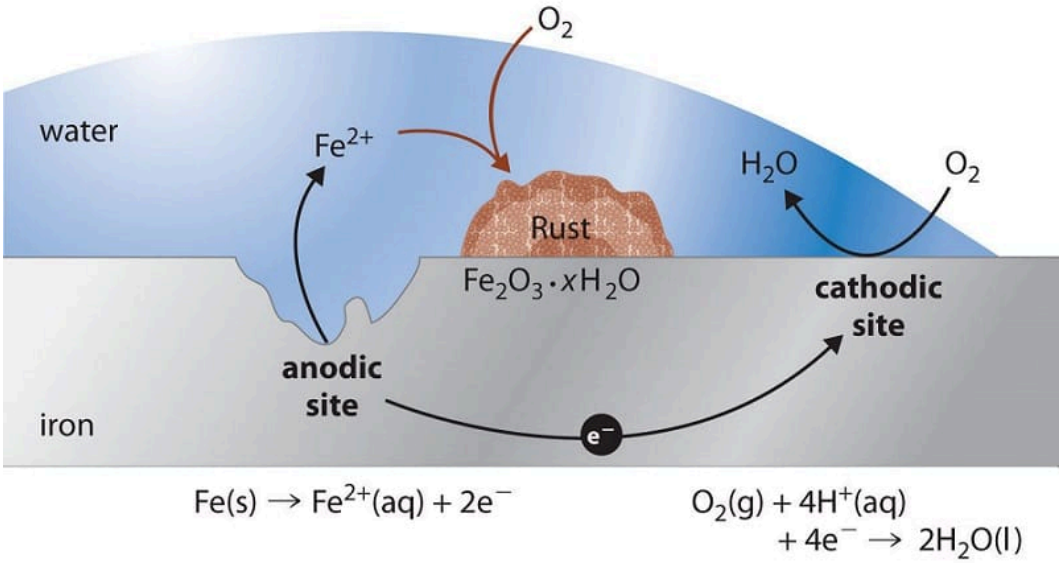


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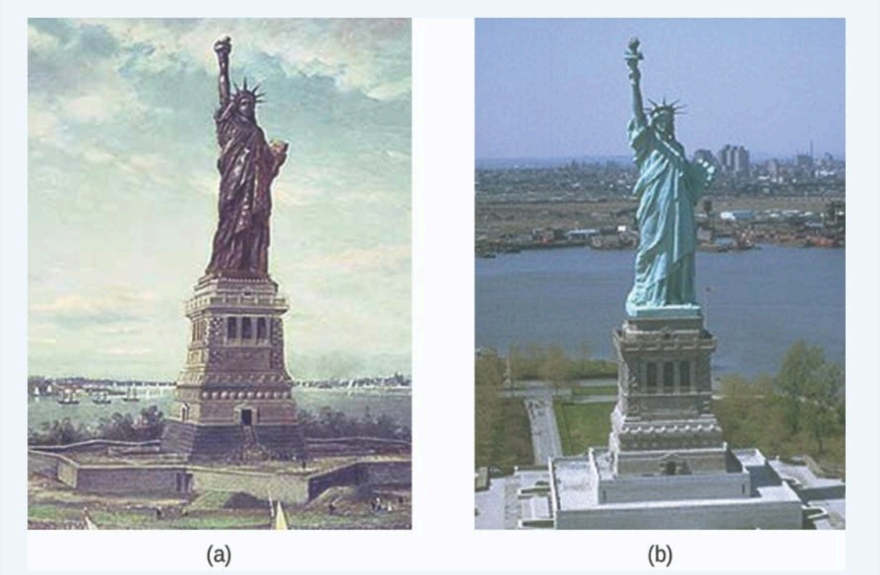


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Learning Objectives

- Define electrochemistry.
- Describe the basic components of electrochemical cells.
- List some of the characteristics, applications and limitations of cells and batteries.
- Know the difference between galvanic and electrolytic cells.
- Define electrolysis and list several of its applications.

Metal exposed to the outside elements will usually corrode if not protected. The corrosion process is a series of redox reactions involving the metal of the sculpture. In some situations, the metals are deliberately left unprotected so that the surface will undergo changes that may enhance the esthetic value of the work.

Electrochemical Reactions

Chemical reactions either absorb or release energy, which can be in the form of electricity. **Electrochemistry** is a branch of chemistry that deals with the interconversion of chemical energy and electrical energy. Electrochemistry has many common applications in everyday life. All sorts of batteries, from those used to power a flashlight to a calculator to an automobile, rely on chemical reactions to generate electricity. Electricity is used to plate objects with decorative metals like gold or chromium. Electrochemistry is important in the transmission of nerve impulses in biological systems. Redox chemistry, the transfer of electrons, is behind all electrochemical processes.

An **electrochemical cell** is any device that converts chemical energy into electrical energy or electrical energy into chemical energy. There are three components that make up an electrochemical reaction. There must be a solution where redox reactions can occur. These reactions generally take place in water to facilitate electron and ion movement. A conductor must exist for electrons to be transferred. This conductor is usually some kind of wire so that electrons can move from one site to another. Ions also must be able to move through some form of salt bridge that facilitates ion migration.

Luigi Galvani

Luigi Galvani (1737 - 1798) was an Italian physician and scientist who did research on nerve conduction in animals. His accidental observation of the twitching of frog legs when they were in contact with an iron scalpel while the legs hung on copper hooks led to studies on electrical conductivity in muscles and nerves. He believed that animal tissues contained an "animal electricity" similar to the natural electricity that caused lightning to form.



Figure 7.6.1: Alessandro Giuseppe Antonio Anastasio Volta (Public Domain; via Wikipedia)

Alessandro Volta developed the first "voltaic cell" in 1800. This battery consisted of alternating disks of zinc and silver with pieces of cardboard soaked in brine separating the disks. Since there were no voltmeters at the time (and no idea that the electric current was due to electron flow), Volta had to rely on another measure of battery strength: the amount of shock produced (it's never a good idea to test things on yourself). He found that the intensity of the shock increased with the number of metal plates in the system. Devices with twenty plates produced a shock that was quite painful. It's a good thing we have voltmeters today to measure electric current instead of the "stick your finger on this and tell me what you feel" method.

Galvanic (Voltaic) Cells

Galvanic cells, also known as voltaic cells, are electrochemical cells in which spontaneous oxidation-reduction reactions produce electrical energy. In writing the equations, it is often convenient to separate the oxidation-reduction reactions into half-reactions to facilitate balancing the overall equation and to emphasize the actual chemical transformations.

Consider what happens when a clean piece of copper metal is placed in a solution of silver nitrate (Figure 7.6.1). As soon as the copper metal is added, silver metal begins to form and copper ions pass into the solution. The blue color of the solution on the far right indicates the presence of copper ions. The reaction may be split into its two half-reactions. Half-reactions separate the oxidation from the reduction, so each can be considered individually.

$$\begin{aligned}
&\text{oxidation: } \mathrm{Cu}(s) \longrightarrow \mathrm{Cu}^{2+}(aq) + 2\,\mathrm{e}^- \\
&\text{reduction: } 2 \times (\mathrm{Ag}^+(aq) + \mathrm{e}^- \longrightarrow \mathrm{Ag}(s)) \qquad \text{or} \qquad 2\,\mathrm{Ag}^+(aq) + 2\,\mathrm{e}^- \longrightarrow 2\,\mathrm{Ag}(s) \\
&\text{overall: } 2\,\mathrm{Ag}^+(aq) + \mathrm{Cu}(s) \longrightarrow 2\,\mathrm{Ag}(s) + \mathrm{Cu}^{2+}(aq)
\end{aligned}$$

The equation for the reduction half-reaction had to be doubled so the number electrons “gained” in the reduction half-reaction equaled the number of electrons “lost” in the oxidation half-reaction.

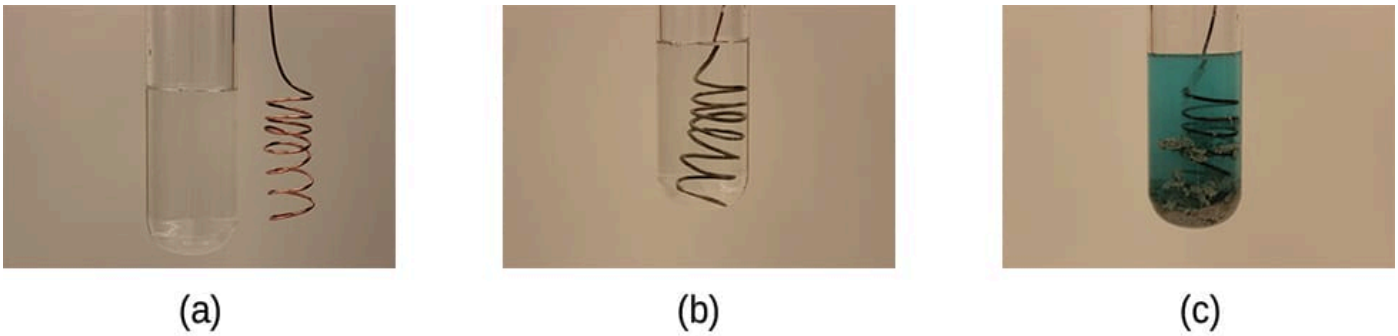


Figure 7.6.1 When a clean piece of copper metal is placed into a clear solution of silver nitrate (a), an oxidation-reduction reaction occurs that results in the exchange of Cu2+ for Ag+ ions in solution. As the reaction proceeds (b), the solution turns blue (c) because of the copper ions present, and silver metal is deposited on the copper strip as the silver ions are removed from solution. (CC BY 4.0; modification of work by Mark Ott via OpenStax)

Galvanic or voltaic cells involve spontaneous electrochemical reactions in which the half-reactions are separated (Figure 7.6.2) so that current can flow through an external wire. The beaker on the left side of the figure is called a half-cell, and contains a 1 *M* solution of copper(II) nitrate [$Cu(NO_3)_2$] with a piece of copper metal partially submerged in the solution. The copper metal is an electrode. The copper is undergoing oxidation; therefore, the copper electrode is the **anode**. The anode is connected to a voltmeter with a wire and the other terminal of the voltmeter is connected to a silver electrode by a wire. The silver is undergoing reduction; therefore, the silver electrode is the **cathode**. The half-cell on the right side of the figure consists of the silver electrode in a 1 *M* solution of silver nitrate ($AgNO_3$). At this point, no current flows—that is, no significant movement of electrons through the wire occurs because the circuit is open. The circuit is closed using a salt bridge, which transmits the current with moving ions.

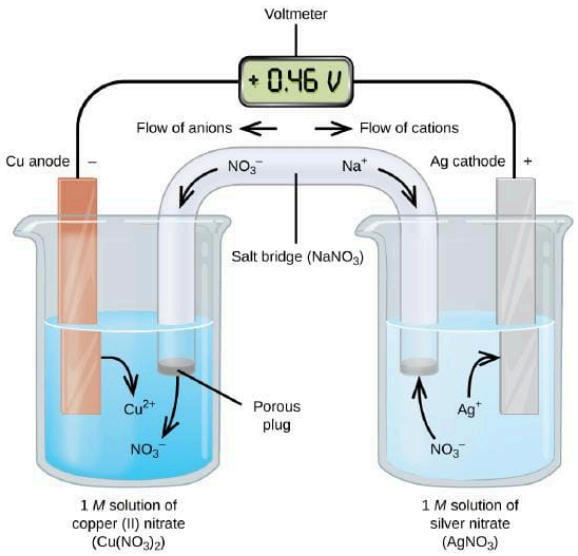


Figure 7.6.2 A standard Cu-Ag galvanic cell.

The salt bridge consists of a concentrated, nonreactive, electrolyte solution such as the sodium nitrate ($NaNO_3$) solution used in this example. As electrons flow from left to right through the electrode and wire, nitrate ions (anions) pass through the porous plug on the left into the copper(II) nitrate solution. This keeps the beaker on the left electrically neutral by neutralizing the charge on the copper(II) ions that are produced in the solution as the copper metal is oxidized. At the same time, the nitrate ions are moving to the left, sodium ions (cations) move to the right, through the porous plug, and into the silver nitrate solution on the right. These added cations “replace” the silver ions that are removed from the solution as they were reduced to silver metal, keeping the beaker on the right electrically neutral.

When the electrochemical cell is constructed in this fashion, a positive cell potential (0.46 V) indicates a spontaneous reaction *and* that the electrons are flowing from the left to the right.

Key Features of a Standard Galvanic Cell (as shown in Figure 7.6.2

- Electrons flow from the anode to the cathode: left to right in the standard galvanic cell.
- The electrode in the left half-cell is the anode because oxidation occurs here. The name refers to the flow of anions in the salt bridge toward it.
- The electrode in the right half-cell is the cathode because reduction occurs here. The name refers to the flow of cations in the salt bridge toward it.
- Oxidation occurs at the anode (the left half-cell in the figure).
- Reduction occurs at the cathode (the right half-cell in the figure).
- The salt bridge must be present to close (complete) the circuit and both an oxidation and reduction must occur for current to flow.

Example 7.6.1: Chromium-Copper Galvanic Cell

Consider a galvanic cell consisting of

$$2\,\mathrm{Cr}(s) + 3\,\mathrm{Cu}^{2+}(aq) \longrightarrow 2\,\mathrm{Cr}^{3+}(aq) + 3\,\mathrm{Cu}(s)$$

Write the oxidation and reduction half-reactions. Which reaction occurs at the anode? The cathode?

Solution

By inspection, Cr is oxidized when three electrons are lost to form Cr^{3+}, and Cu^{2+} is reduced as it gains two electrons to form Cu. Balancing the charge gives

$$\begin{aligned}&\text{oxidation: } 2\,\mathrm{Cr}(s) \longrightarrow 2\,\mathrm{Cr}^{3+}(aq) + 6\,\mathrm{e}^-\\&\text{reduction: } 3\,\mathrm{Cu}^{2+}(aq) + 6\,\mathrm{e}^- \longrightarrow 3\,\mathrm{Cu}(s)\\&\overline{\text{overall: } 2\,\mathrm{Cr}(s) + 3\,\mathrm{Cu}^{2+}(aq) \longrightarrow 2\,\mathrm{Cr}^{3+}(aq) + 3\,\mathrm{Cu}(s)}\end{aligned}$$

Cell notation uses the simplest form of each of the equations, and starts with the reaction at the anode. No concentrations were specified so:

$$\mathrm{Cr}(s)\,|\,\mathrm{Cr}^{3+}(aq)\,\|\,\mathrm{Cu}^{2+}(aq)\,|\,\mathrm{Cu}(s).$$

Oxidation occurs at the anode and reduction at the cathode.

Exercise 7.6.1 Copper-Zinc Galvanic Cell

Consider a galvanic cell consisting of

$$\mathrm{Zn}(s) + \mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Cu}(s)$$

Write the oxidation and reduction half-reactions. Which reaction occurs at the anode? The cathode?

Answer available at the end of this section.

Batteries

A battery is an electrochemical cell or series of cells that produces an electric current. In principle, any galvanic cell could be used as a battery. An ideal battery would never run down, produce an unchanging voltage, and be capable of withstanding environmental extremes of heat and humidity. Real batteries strike a balance between ideal characteristics and practical limitations. For example, the mass of a car battery is about 18 kg or about 1% of the mass of an average car or light-duty truck. This type of battery would supply nearly unlimited energy if used in a smartphone, but would be rejected for this application because of its mass. Thus, no single battery is “best” and batteries are selected for a particular application, keeping things like

the mass of the battery, its cost, reliability, and current capacity in mind. There are two basic types of batteries: primary and secondary. A few batteries of each type are described next.

Visit this site to learn more about batteries.

Dry Cells (Primary Batteries)

Primary batteries are single-use batteries because they cannot be recharged. A common primary battery is the **dry cell** (Figure 7.6.1). The dry cell is a zinc-carbon battery. The zinc can serves as both a container and the negative electrode. The positive electrode is a rod made of carbon that is surrounded by a paste of manganese(IV) oxide, zinc chloride, ammonium chloride, carbon powder, and a small amount of water. The reaction at the anode can be represented as the ordinary oxidation of zinc:

$$\mathrm{Zn}(s) \longrightarrow \mathrm{Zn}^{2+}(aq) + 2\,\mathrm{e}^-$$

[Original Image Removed]

Figure 7.6.3 A diagram of a cross section of a dry cell battery is shown. The overall shape of the cell is cylindrical. The lateral surface c cylinder, indicated as a thin red line, is labeled “zinc can (electrode).” Just beneath this is a slightly thicker dark grey surface that cove lateral surface, top, and bottom of the battery, which is labeled “Porous separator.” Inside is a purple region with many evenly spaced darker purple dots, labeled “Paste of M n O subscript 2, N H subscript 4 C l, Z n C l subscript 2, water (cathode).” A dark grey rod, lab “Carbon rod (electrode),” extends from the top of the battery, leaving a gap of less than one-fifth the height of the battery below the rod bottom of the cylinder. A thin grey line segment at the very bottom of the cylinder is labeled “Metal bottom cover (negative).” The very the cylinder has a thin grey surface that curves upward at the center over the top of the carbon electrode at the center of the cylinder. upper surface is labeled “Metal top cover (positive).” A thin dark grey line just below this surface is labeled “Insulator.” Below this, abo purple region, and outside of the carbon electrode at the center is an orange region that is labeled “Seal.”

The reaction at the cathode is more complicated, in part because more than one reaction occurs. The series of reactions that occurs at the cathode is approximately

$$2\,\mathrm{MnO_2}(s) + 2\,\mathrm{NH_4Cl}(aq) + 2\,\mathrm{e}^- \longrightarrow \mathrm{Mn_2O_3}(s) + 2\,\mathrm{NH_3}(aq) + \mathrm{H_2O}(l) + 2\,\mathrm{Cl}^-$$

The overall reaction for the zinc–carbon battery can be represented as

$$2\,\mathrm{MnO_2}(s) + 2\,\mathrm{NH_4Cl}(aq) + \mathrm{Zn}(s) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Mn_2O_3}(s) + 2\,\mathrm{NH_3}(aq) + \mathrm{H_2O}(l) + 2\,\mathrm{Cl}^-$$

with an overall cell potential which is initially about 1.5 V, but decreases as the battery is used. It is important to remember that the voltage delivered by a battery is the same regardless of the size of a battery. For this reason, D, C, A, AA, and AAA batteries all have the same voltage rating. However, larger batteries can deliver more moles of electrons.

Take Note

The dry cell is not very efficient in producing electrical energy because only the relatively small fraction of the MnO_2 that is near the cathode is actually reduced and only a small fraction of the zinc cathode is actually consumed as the cell discharges. In addition, dry cells have a limited shelf life because the Zn anode reacts spontaneously with NH_4Cl in the electrolyte, causing the case to corrode and allowing the contents to leak out (Figure 7.6.2).



Figure 7.6.2 A dry cell with a leak. Source: Photo courtesy of Mitchclanky2008, http://www.flickr.com/photos/25597837@N05/2422765479/.

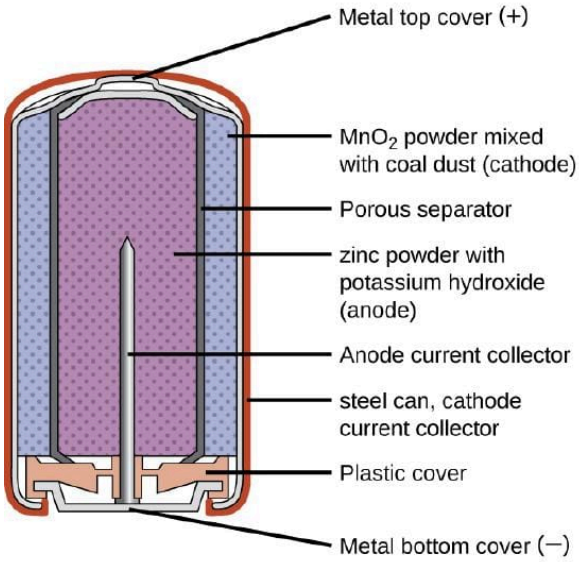


Figure 7.6.4 A cross section of an alkaline battery.

Alkaline batteries (Figure 7.6.4) were developed in the 1950s partly to address some of the performance issues with zinc–carbon dry cells. They are manufactured to be exact replacements for zinc-carbon dry cells. As their name suggests, these types of batteries use alkaline electrolytes, often potassium hydroxide.

An alkaline battery can deliver about three to five times the energy of a zinc-carbon dry cell of similar size. Alkaline batteries are prone to leaking potassium hydroxide, so these should also be removed from devices for long-term storage. While some alkaline batteries are rechargeable, most are not.

Warning

Attempts to recharge an alkaline battery that is not rechargeable often leads to rupture of the battery and leakage of the potassium hydroxide electrolyte.

Lead Storage Batteries (Secondary Batteries)

The lead acid battery (Figure 7.6.5) is the type of secondary battery used in your automobile. Secondary batteries are rechargeable. The lead acid battery is inexpensive and capable of producing the high current required by automobile starter motors. The reactions for a lead acid battery are

$$\begin{aligned}&\text{anode: } \mathrm{Pb}(s) + \mathrm{HSO_4}^-(aq) \longrightarrow \mathrm{PbSO_4}(s) + \mathrm{H}^+(aq) + 2\,\mathrm{e}^-\\ &\underline{\text{cathode: } \mathrm{PbO_2}(s) + \mathrm{HSO_4}^-(aq) + 3\,\mathrm{H}^+(aq) + 2\,\mathrm{e}^- \longrightarrow \mathrm{PbSO_4}(s) + 2\,\mathrm{H_2O}(l)}\\ &\text{overall: } \mathrm{Pb}(s) + \mathrm{PbO_2}(s) + 2\,\mathrm{H_2SO_4}(aq) \longrightarrow 2\,\mathrm{PbSO_4}(s) + 2\,\mathrm{H_2O}(l)\end{aligned}$$

Each cell produces 2 V, so six cells are connected in series to produce a 12-V car battery. Lead acid batteries are heavy and contain a caustic liquid electrolyte, but are often still the battery of choice because of their high current density. The lead acid battery in your automobile consists of six cells connected in series to give 12 V. Their low cost and high current output makes these excellent candidates for providing power for automobile starter motors.

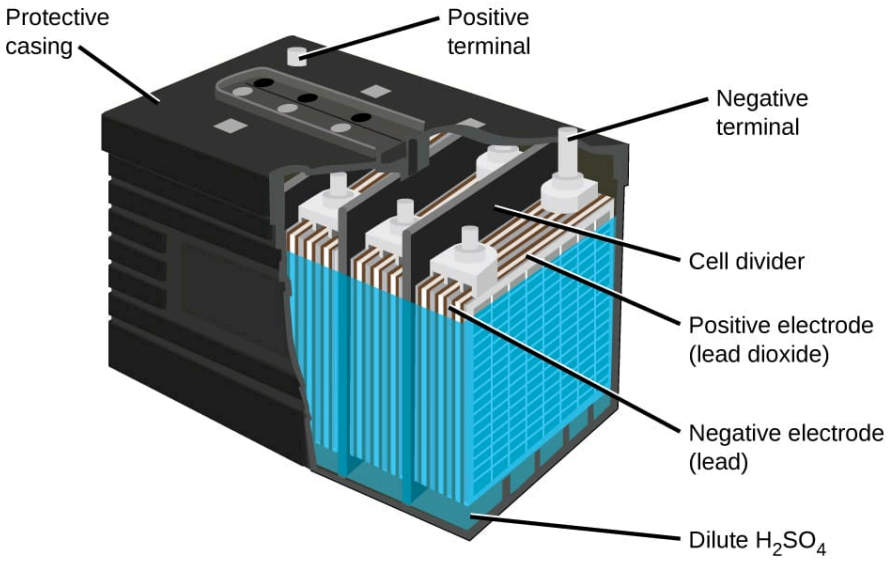


Figure 7.6.5 A lead (acid) storage battery.

As mentioned earlier, unlike a dry cell, the lead storage battery is rechargeable. Note that the forward redox reaction generates solid lead (II) sulfate which slowly builds up on the plates. Additionally, the concentration of sulfuric acid decreases. When the car is running normally, its generator **recharges** the battery by forcing the above reactions to run in the opposite, or non spontaneous direction.

$$2\mathrm{PbSO_4}\,(s) + 2\mathrm{H_2O}\,(l) \rightarrow \mathrm{Pb}\,(s) + \mathrm{PbO_2}\,(s) + 4\mathrm{H}^+\,(aq) + 2\mathrm{SO_4}^{2-}\,(aq)$$

This reaction regenerates the lead, lead (IV) oxide, and sulfuric acid needed for the battery to function properly. Theoretically, a lead storage battery should last forever. In practice, the recharging is not 100% efficient because some of the lead (II) sulfate falls from the electrodes and collects on the bottom of the cells.

Warning

Lead storage batteries contain a significant amount of lead so they must always be disposed of properly.

Other Batteries

The **nickel–cadmium**, or NiCad, battery (Figure 7.6.6) is used in small electrical appliances and devices like drills,

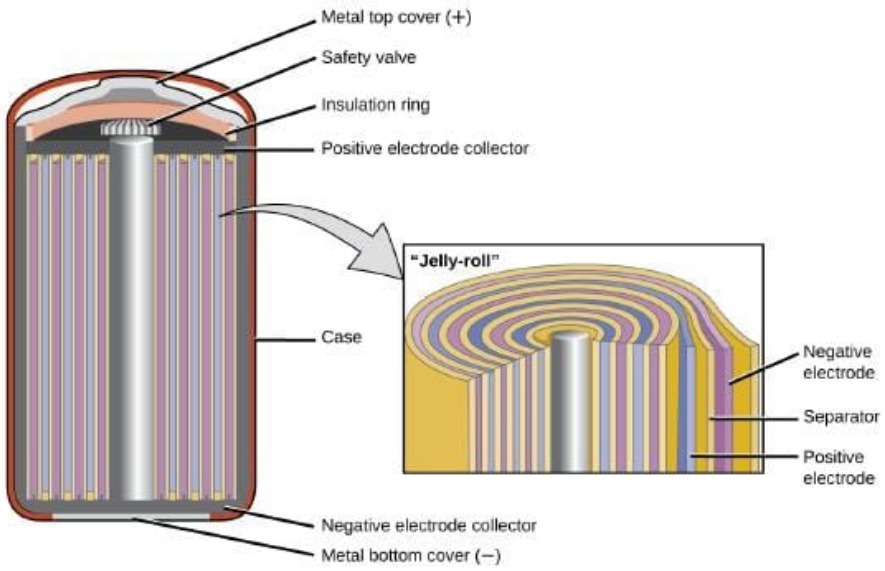


Figure 7.6.6 NiCd battery with “jelly-roll” design.

portable vacuum cleaners, and AM/FM digital tuners. It consists of a nickel-plated cathode, cadmium-plated anode, and a potassium hydroxide electrode. The positive and negative plates, which are prevented from shorting by the separator, are rolled together and put into the case. This is a “jelly-roll” design and allows the NiCd cell to deliver much more current than a similar-sized alkaline battery.

The voltage is about 1.2 V to 1.25 V as the battery discharges. When properly treated, a NiCd battery can be recharged about 1000 times.

Warning

Cadmium is a toxic heavy metal so NiCd batteries should never be opened or put into the regular trash.

Lithium ion batteries are among the most popular rechargeable batteries and are used in many portable electronic devices. The battery voltage is about 3.7 V. Lithium batteries are popular because they can provide a large amount current, are lighter than comparable batteries of other types, produce a nearly constant voltage as they discharge, and only slowly lose their charge when stored. Specialized **lithium-iodide (polymer) batteries** find application in many long-life, critical devices, such as pacemakers and other implantable electronic medical devices. These devices are designed to last 15 or more years.

Disposable primary lithium batteries must be distinguished from secondary lithium-ion or a lithium-polymer. The term "lithium battery" refers to a family of different lithium-metal chemistries, comprising many types of cathodes and electrolytes but all with metallic lithium as the anode. Lithium batteries are widely used in portable consumer electronic devices, and in electric vehicles ranging from full sized vehicles to radio controlled toys.



A **watch battery, coin** or **button cell** (Figure 7.6.7) is a small single cell battery shaped as a squat cylinder typically 5 to 25 mm (0.197 to 0.984 in) in diameter and 1 to 6 mm (0.039 to 0.236 in) high — like a button on a garment, hence the name. A metal can forms the bottom body and positive terminal of the cell. An insulated top cap is the negative terminal. Button cells are single cells, usually disposable primary cells. Common anode materials are zinc or lithium. Common cathode materials are manganese dioxide, silver oxide, carbon monofluoride, cupric oxide or oxygen from the air. Mercuric oxide button cells were formerly common, but are no longer available due to the toxicity and environmental effects of mercury.

Warning

Button cells are very dangerous for small children. Button cells that are swallowed can cause severe internal burns and significant injury or death.

Fuel Cells

A fuel cell is a device that converts chemical energy into electrical energy. Fuel cells are similar to

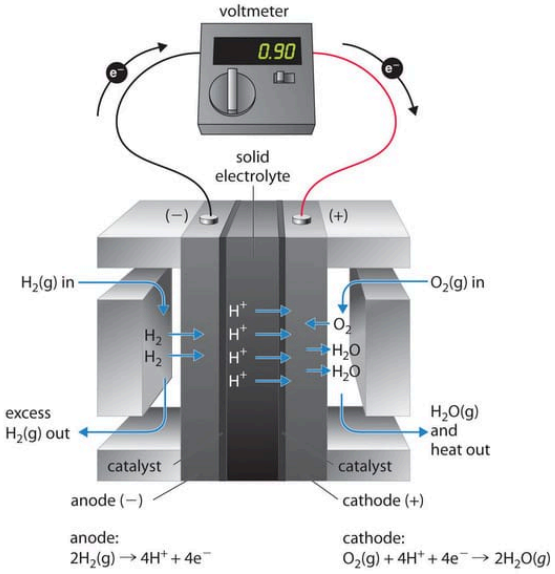


Figure 7.6.8 A hydrogen fuel cell.

batteries but require a continuous source of fuel, often hydrogen. They will continue to produce electricity as long as fuel is available. Hydrogen fuel cells have been used to supply power for satellites, space capsules, automobiles, boats, and submarines (Figure 7.6.8).

The voltage is about 0.9 V. The efficiency of fuel cells is typically about 40% to 60%, which is higher than the typical internal combustion engine (25% to 35%) and, in the case of the hydrogen fuel cell, produces only water as exhaust. Although fuel cells are an essentially pollution-free means of obtaining electrical energy, their expense and technological complexity have thus far limited their applications.

Electrolysis

In this chapter, we have described various galvanic cells in which a spontaneous chemical reaction is used to generate electrical energy. In an electrolytic cell, however, the opposite process, called electrolysis, occurs: an external voltage is applied to drive a nonspontaneous reaction. Electrolysis has many commercial and industrial applications. Electrometallurgy is the process of reduction of metals from metallic compounds to obtain the pure form of metal using electrolysis. Aluminium, lithium, sodium, potassium, magnesium, calcium, and in some cases copper, are produced in this way. Chlorine and sodium hydroxide are the major products of the electrolysis of brine, a water/sodium chloride mixture. Water electrolysis is used to generate oxygen for spacecraft and nuclear submarines.

Electrolysis is also used in the cleaning and preservation of old artifacts. Because the process separates the non-metallic particles from the metallic ones, it is very useful for cleaning a wide variety of metallic objects, from old coins to even larger objects including rusted cast iron cylinder blocks and heads when rebuilding automobile engines. Rust removal from small iron or steel objects by electrolysis can be done in a home workshop using simple materials such as a plastic bucket, tap water, lengths of rebar, washing soda, baling wire, and a battery charger.

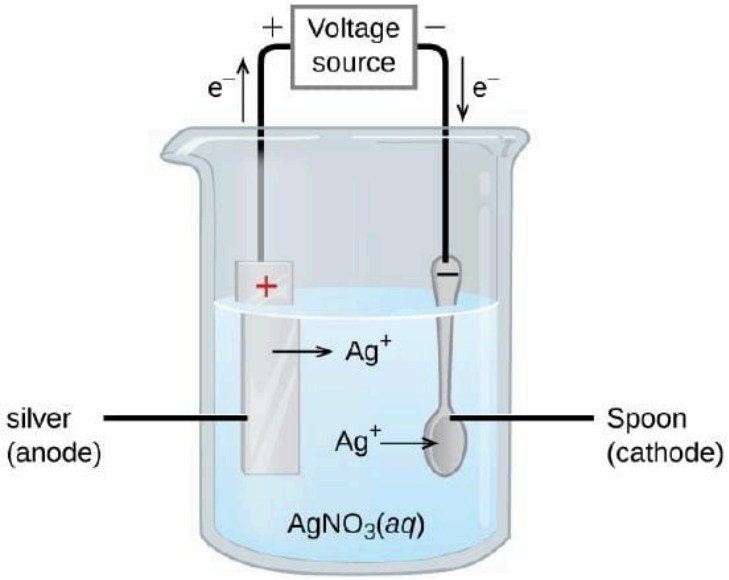


Figure 7.6.9: Silver plating of a spoon.

An important use for electrolytic cells is in **electroplating**. Electroplating results in a thin coating of one metal on top of a conducting surface. Reasons for electroplating include making the object more corrosion resistant, strengthening the surface, producing a more attractive finish, or for purifying metal. The metals commonly used in electroplating include cadmium, chromium, copper, gold, nickel, silver, and tin. Common consumer products include silver-plated or gold-plated tableware, chrome-plated automobile parts, and jewelry. We can get an idea of how this works by investigating how silver-plated tableware is produced (Figure 7.6.9).

Summary

- Electrochemistry is a branch of chemistry that deals with the interconversion of chemical energy and electrical energy.
- Batteries are galvanic cells, or a series of cells, that produce an electric current.

- There are two basic types of batteries: primary and secondary.
- Primary batteries are “single use” and cannot be recharged. Dry cells and (most) alkaline batteries are examples of primary batteries.
- The second type is rechargeable and is called a secondary battery. Examples of secondary batteries include nickel-cadmium (NiCd), lead acid, and lithium ion batteries.
- Fuel cells are similar to batteries in that they generate an electrical current, but require continuous addition of fuel and oxidizer. The hydrogen fuel cell uses hydrogen and oxygen from the air to produce water, and is generally more efficient than internal combustion engines.
- Electrolysis occurs when an external voltage is applied to drive a nonspontaneous reaction.

Contributors and Attributions

-

- Paul Flowers (University of North Carolina - Pembroke), Klaus Theopold (University of Delaware) and Richard Langley (Stephen F. Austin State University) with contributing authors. Textbook content produced by OpenStax College is licensed under a Creative Commons Attribution License 4.0 license. Download for free at http://cnx.org/contents/85abf193-2bd...a7ac8df6@9.110).

- TextMap: Chemistry-The Central Science (Brown et al.)

- Marisa Alviar-Agnew (Sacramento City College)

- Wikipedia

Learning Objectives

- Learn the parts of a voltaic cell.
- Combine half reactions to determine the voltage of a voltaic cell.
- Understand how voltaic cells are used as batteries.

Consider this redox reaction:

$$\mathrm{Zn} + \mathrm{Cu}^{2+} \longrightarrow \mathrm{Zn}^{2+} + \mathrm{Cu}$$

If you were to mix zinc metal and copper ions in a container, this reaction would proceed by itself; we say that this reaction is *spontaneous*.

Suppose, however, we set up this reaction in a way depicted in Fig 14.4.1 - A Redox Reaction in Which the Two Half Reactions Are Physically Separated. Zinc and zinc ions are on one side of the system, while copper and copper ions are on the other side of the system. The two parts are connected with a wire.

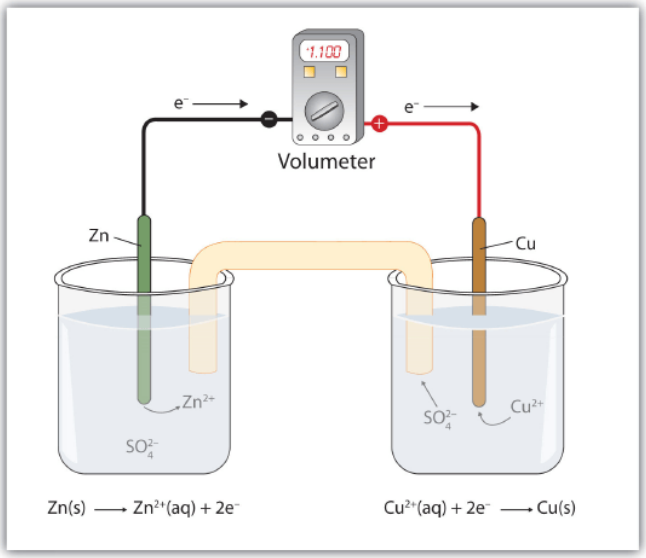


Figure 7.6.1 A Redox Reaction in Which the Two Half Reactions Are Physically Separated. One application of redox reactions requires that they be physically separated.

Even though the two half reactions are physically separated, a spontaneous redox reaction still occurs. However, in this case, the electrons transfer through the wire connecting the two half reactions; that is, this setup becomes a source of electricity. Useful work can be extracted from the electrons as they transfer from one side to the other. For example, a light bulb can be lit, or a motor can be operated. The apparatus as a whole, which allows useful electrical work to be extracted from a redox reaction, is called a **voltaic (galvanic) cell**.

Each individual system that contains a half reaction is called a **half cell**. The half cell that contains the oxidation reaction is called the **anode**, while the half cell that contains the reduction reaction is called the **cathode**. The cathode and anode collectively are the **electrodes** of the voltaic cell. Because electrons are coming from the anode, the anode is considered the *negative* electrode of the cell, while the cathode is considered the *positive* electrode of the cell. Finally, because electrons are moving from one half cell to the other, a charge imbalance builds up as the reaction proceeds. To counter that, a **salt bridge** is used; the salt bridge contains a solution of some ionic compound whose ions migrate to either side of the voltaic cell to maintain the charge balance.

The tendency for electrons to go from one half cell to another is called the **voltage** of the voltaic cell, represented by *E*. Sometimes the term *potential* is used to represent the voltage of a cell. Voltage is expressed in volts (V). The voltage of a voltaic cell is determined by the *difference* in the tendencies of the individual half cells and is characteristic of a given redox reaction when concentrations are specific (1.0 M for dissolved species and 1.0 atm for gases). Because the voltage of a redox reaction is determined by the difference of the tendencies of the individual half reactions, absolute voltages are unnecessary; only relative voltages of each half reaction are needed. The relative voltage of each half cell is represented as $E_{1/2}$ and is based on the standard that the $E_{1/2}$ for the reaction

$$H^+ + e^- \rightarrow \frac{1}{2}H_2$$

is assigned to be exactly 0.000 V under standard conditions of pressure and concentration. Table 7.6.1 - Standard Reduction Potentials of Half Reactions, lists some relative $E_{1/2}$ values for some half reactions. Note that all half reactions are listed as reduction reactions, so these values are called the **standard reduction potentials** of each half reaction.

Reduction Half Reaction	$E_{1/2}$ (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.61
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51

Reduction Half Reaction	$E_{1/2}$ (V)
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23
$Br_2 + 2e^- \rightarrow 2Br^-$	1.06
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96
$Ag^+ + e^- \rightarrow Ag$	0.80
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77
$I_2 + 2e^- \rightarrow 2I^-$	0.54
$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$AgCl + e^- \rightarrow Ag + Cl^-$	0.222
$Sn^{4+} + 2e^- \rightarrow Sn^{2+}$	0.15
$2H^+ + 2e^- \rightarrow H_2$	0.000
$Pb^{2+} + 2e^- \rightarrow Pb$	−0.126
$Ni^{2+} + 2e^- \rightarrow Ni$	−0.25
$Cr^{3+} + e^- \rightarrow Cr^{2+}$	−0.41
$Fe^{2+} + 2e^- \rightarrow Fe$	−0.44
$Cr^{3+} + 3e^- \rightarrow Cr$	−0.74
$Zn^{2+} + 2e^- \rightarrow Zn$	−0.76
$Cr^{2+} + 2e^- \rightarrow Cr$	−0.91
$Ba^{2+} + 2e^- \rightarrow Ba$	−1.57
$Al^{3+} + 3e^- \rightarrow Al$	−1.66
$Mg^{2+} + 2e^- \rightarrow Mg$	−2.37
$Na^+ + e^- \rightarrow Na$	−2.714
$Li^+ + e^- \rightarrow Li$	−3.045

Table 7.6.1 *Standard Reduction Potentials of Half Reactions*

The above table lists only reduction reactions, but a redox reaction has a reduction *and* an oxidation. To make the oxidation reaction, simply reverse the reduction reaction in the above table and change the sign on the $E_{1/2}$ value. If the reduction potential is negative, make the voltage for the oxidation positive; if the reduction potential is positive, make the voltage for the oxidation negative.

Example 7.6.1

What is the value of $E_{1/2}$ for this half reaction?

$$Ag + Cl^- \rightarrow AgCl + e^-$$

Solution

The given reaction is the reverse of this reaction:

$$AgCl + e^- \rightarrow Ag + Cl^- \quad E_{1/2} = 0.222 \text{ V}$$

Therefore, the $E_{1/2}$ of the given reaction is −0.222 V.

Exercise 7.6.1

What is the value of $E_{1/2}$ for this half reaction?

$$Na \rightarrow Na^+ + e^-$$

Answer

2.714 V

To determine the overall voltage of a particular voltaic cell, simply combine the voltages of the oxidation and reduction half reactions. Even if you need to take a multiple of a half reaction for the electrons to cancel, do not take the multiple of the $E_{1/2}$. Use the values directly as is from Table 7.6.1 - Standard Reduction Potentials of Half Reactions.

Spontaneous redox reactions have positive overall voltages. If the voltage of the reaction as written is negative, it is not spontaneous in that direction. Rather, the reverse reaction is the spontaneous redox reaction.

Example 7.6.2

What is the voltage of a voltaic cell based on this reaction? Is the reaction spontaneous as written?

$$2NO_3^- + 8H^+ + 3Cu \rightarrow 2NO + 4H_2O + 3Cu^{2+}$$

Solution

The overall redox reaction is formed from these two half reactions:

$$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O \quad E_{1/2} = 0.96 \text{ V} \qquad Cu^{2+} + 2\,e^- \rightarrow Cu \quad E_{1/2} = 0.34 \text{ V}$$

The second reaction is reversed in the overall redox reaction, so its voltage changes sign from the reduction reaction:

$$Cu \rightarrow Cu^{2+} + 2\,e^- \quad E_{1/2} = -0.34 \text{ V}$$

To obtain the voltage of the voltaic cell based on the overall reaction, we simply combine the two voltages of the half reactions:

$$E = 0.96 + (-0.34) = 0.62 \text{ V}$$

Because the overall voltage is positive, the reaction is spontaneous as written.

Exercise 7.6.2

What is the voltage of a voltaic cell based on this reaction? Is the reaction spontaneous as written?

$$5Ni + 2MnO4^{-} + 16H^{+} \rightarrow 3Mn^{2+} + 8H2O + 5Ni^{2+}$$

Answer

1.76 V; spontaneous

Technically, any redox reaction can be set up to make a voltaic cell. In modern society, however, only certain redox reactions are put to practical use. A portable voltaic cell that generates electricity to power devices for our convenience is called a **battery**. All batteries are based on redox reactions.

The first battery (called a "voltaic pile") was constructed by the Italian scientist Alessandro Volta in 1800 and was based on the copper/zinc reaction depicted in Figure 7.6.1 - A Redox Reaction in Which the Two Half Reactions Are Physically Separated. Unfortunately, it was messy, requiring quantities of copper and zinc salts dissolved in water. In 1866, the French scientist Georges Leclanché invented the **dry cell,** a precursor to today's modern battery. A schematic of a dry cell is shown in Figure 7.6.2 - Dry Cell. The zinc case and the central carbon rod serve as the anode and cathode, respectively. The other reactants are combined into a moist paste that minimizes free liquid, so the battery is less messy (hence the name *dry cell*).

The actual redox reaction is complex but can be represented by the following redox reaction:

$$Zn + 2MnO2 + 2NH4^{+} \rightarrow Zn^{2+} + Mn2O3 + 2NH3 + H2O$$

A dry cell has a voltage of about 1.56 V. While common and useful, dry cells have relatively short lifetimes and contain acidic components. They also cannot be recharged, so they are one-use only. Batteries that can be used only once are called **primary batteries**.

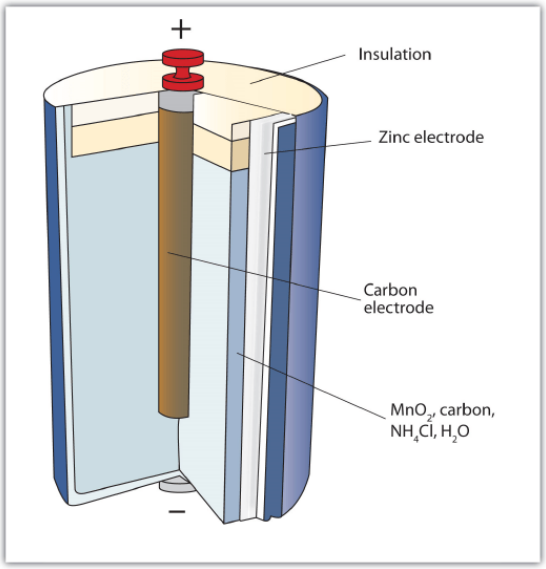


Figure 7.6.2 Dry Cell. The Leclanché dry cell is a common type of battery.

In the late 1950s, Lewis Urry of the Eveready Battery Company in Ohio invented the **alkaline battery** (still marketed today under the trade name *Energizer*). Alkaline batteries are similar to dry cells, but they use a basic moist paste rather than an acidic one. Moreover, the net amount of base does not change during the course of the redox reaction. The overall redox reaction is as follows:

$$Zn + 2MnO_2 \rightarrow ZnO + Mn_2O_3$$

Alkaline batteries have the advantage of being longer lasting and holding their voltage better—about 1.54 V—throughout their lifetime.

A common type of battery, especially with the increased popularity of personal electronic devices, is the button battery (Figure 7.6.3 - Button Battery). A button battery is a small battery that can power small electronic devices; the batteries can be as small as 5 mm across. Two popular redox reactions used for button batteries are the alkaline dry-cell reaction and a silver oxide-based reaction:

$$Zn + Ag_2O \rightarrow ZnO + 2Ag$$

Some button batteries use a lithium-based redox reaction, typified by this anode reaction:

$$Li \rightarrow Li^+ + e^- \quad E_{1/2} = 3.045 \text{ V}$$

The actual redox reaction depends on the composition of the cathode and is variable depending on voltage. Lithium batteries can also be used for applications that require more energy, such as portable computers and electric vehicles. Some lithium-based batteries are rechargeable and can be used over and over again; such batteries are called **secondary batteries**.

An important secondary battery is the lead storage battery, shown in the Figure 7.6.4 - Lead Storage Batteries. The lead storage battery is based on this redox reaction:

$$Pb + PbO_2 + 4H^+ + SO_4^{2-} \rightarrow 2PbSO_4 + 2H_2O$$

The redox reaction produces about 2 V, but it is typical to tie several individual batteries together to generate a larger voltage. The lead storage battery has the distinction that the product of both half reactions is $PbSO_4$, which as a solid that accumulates on the many plates within each cell. The lead storage battery is a secondary battery, as it can be recharged and reused many times. Because it is based on lead, these batteries are rather heavy. They should also be recycled when replaced so that potentially dangerous lead does not escape into the environment. Because of their characteristics, lead storage batteries are used to start large engines in automobiles, boats, and airplanes.

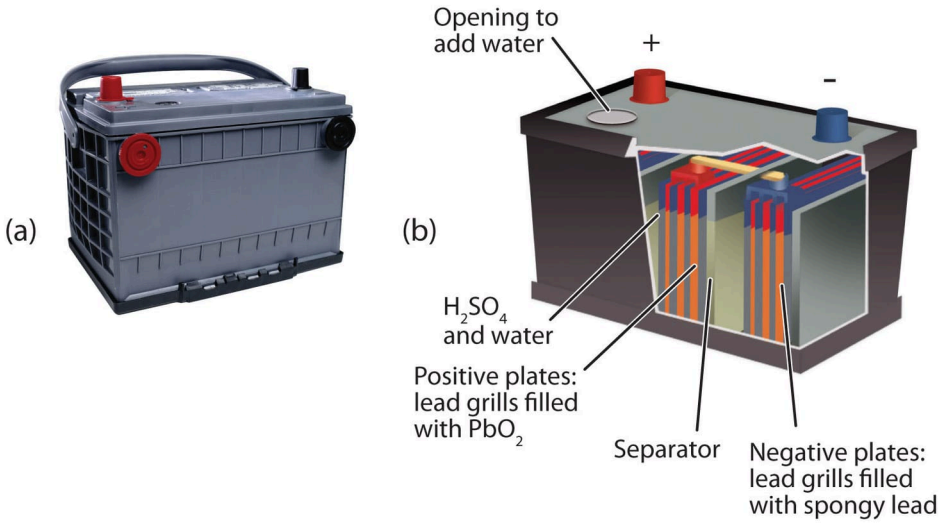


Figure14.4.4 Lead Storage Batteries Photo © Thinkstock. (a) A photo of a lead storage battery. (b) A schematic diagram of a lead storage battery.

Chemistry is Everywhere: Fuel Cells

A *fuel cell* is a type of battery in which reactants flow continuously into a specialized reaction chamber, and products flow out continuously while electrons are extracted from the reaction. Because all reactions in a fuel cell consist of a fuel and an oxidizer undergoing a redox reaction, an introduction of fuel cells is at home in a discussion of redox chemistry.

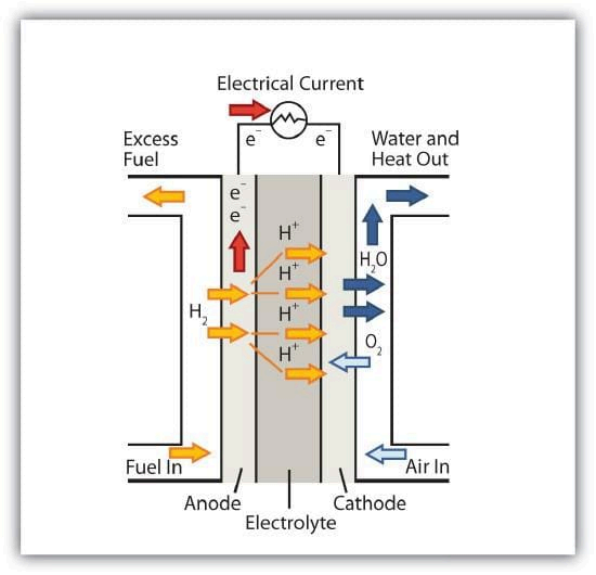


Figure 7.6.5 Fuel Cells. *This fuel cell uses H_2 as the fuel and O_2 as the oxidizer.*

By far the most common fuel cell reaction is based on hydrogen and oxygen:

$$2H_2 + O_2 \rightarrow 2H_2O \quad E = 1.23 \text{ V under standard conditions}$$

However, fuel cells typically do not work under standard nor even optimal conditions, so they typically generate about 0.6–0.7 V. In this fuel cell, the only two products are water and electricity, so the fuel cell not only does not create pollution, but also makes a by-product that in some environments is a valuable commodity (water). Other fuels can be used besides hydrogen; fuel cells have been developed that work on methane, methyl alcohol, ethyl alcohol, carbon-rich materials, and even magnesium metal.

Hydrogen-based fuel cells were and are used to provide electricity for manned space vehicles, partly because their only chemical product is water, which could be used for drinking. However, there has been a recent resurgence in interest in fuel cells because of their potential use in electric cars. Most electric cars run on conventional batteries, which can be very heavy and expensive to replace. It is thought that fuel cells, rather than conventional batteries, might be better sources of electricity for automobiles.

Several current barriers to fuel cell use in electric cars include capacity, cost, and overall energy efficiency. The 2008 Honda FCX, the first production model of a vehicle powered with a fuel cell, can hold 4.1 kg (just under 9 lb) of highly pressured H_2 gas and has a range of 450 km (280 mi). It costs about $120,000–$140,000 to build, making the vehicle beyond the ability of most people to own. Finally, it always requires more energy to produce elemental hydrogen as a fuel than can be extracted from hydrogen as a fuel. As such, hydrogen is described as an energy carrier (like electricity) rather than an energy source (like oil and gas). This distinction points out a fundamental argument against fuel cells as a "better" power source.



Figure 7.6.6 Honda FCX. Source: Photo courtesy of vernieman, commons.wikimedia.org/wiki/File:Honda_FCX_2006_KLIMS_front.jpg.

The 2008 Honda FCX was the first production car to use a fuel cell as a power source. Nonetheless, the car is in very limited service because of its need for relatively large quantities of elemental hydrogen as fuel.

The limitations notwithstanding, there is a lot of interest in fuel cell research. If ways can be found to circumvent their current limitations, fuel cells may become more and more common as power sources.

Key Takeaways

- A voltaic cell produces electricity as a redox reaction occurs.
- The voltage of a voltaic cell can be determined by the reduction potentials of the half reactions.
- Voltaic cells are fashioned into batteries, which are a convenient source of electricity.

Learning Objectives

- Describe electrolysis from a perspective of redox reactions.
- Give examples of electrolysis applications.

Up to this point, we have considered redox reactions for processes that are spontaneous. When set up as a voltaic cell or battery, such reactions can be used as a source of electricity. However, it is possible to go in the other direction. By forcing electricity into a cell, we can make a redox reaction occur that normally would not be spontaneous. Under these circumstances, the cell is called an **electrolytic cell**, and the process that occurs in the cell is called **electrolysis** (Figure 7.6.1).

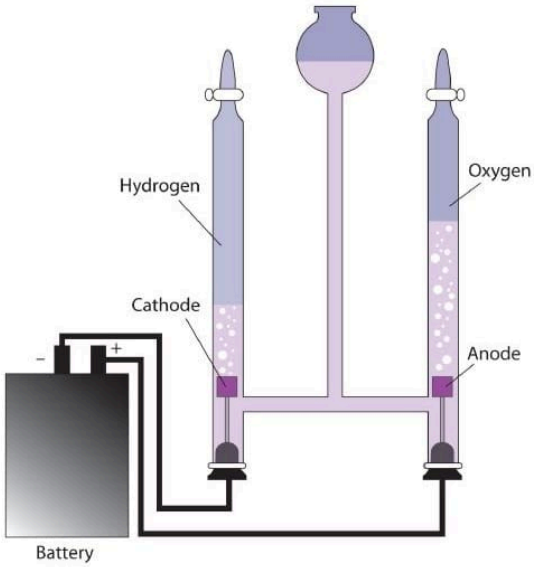


Figure 7.6.1: Electrolysis. In an electrolytic cell, electricity is forced through the cell to induce a non-spontaneous redox reaction. Here, the redox reaction $2H_2O \rightarrow 2H_2 + O_2$ is being caused by the introduction of electricity, which is supplied by the battery.

Electrolysis has many applications. For example, if NaCl is melted at about 800°C in an electrolytic cell and an electric current is passed through it, elemental sodium will appear at the cathode and elemental chlorine will appear at the anode as the following two reactions occur:

$$Na^+ + e^- \rightarrow Na$$

$$2Cl^- \rightarrow Cl_2 + 2e^-$$

Normally we expect elemental sodium and chlorine to react spontaneously to make NaCl. However, by using an input of electricity, we can force the opposite reaction to occur and generate the elements. Lithium, potassium, and magnesium can also be isolated from compounds by electrolysis.

Another element that is isolated by electrolysis is aluminum. Aluminum formerly was a difficult metal to isolate in its elemental form; in fact, the top of the Washington Monument has a 2.8 kg cap of aluminum metal, which at the time (1884) was the largest piece of elemental aluminum ever isolated. However, in 1886, the American Charles Hall and the Frenchman Paul Héroult almost simultaneously worked out an electrolytic process for isolating aluminum from bauxite—an ore of aluminum whose chemical formula is $AlO_x(OH)_{3-2x}$. The basic reactions are as follows:

$$Al^{3+} + 3e^- \rightarrow Al_2O^{2-} \rightarrow O_2 + 4e^-$$

With the development of the Hall-Héroult process, the price of aluminum dropped by a factor of over 200, and aluminum metal became common. So much elemental aluminum is produced in the United States each year that it has been estimated that the electrolysis of aluminum uses 5% of all the electricity in the country. (Recycling aluminum requires about 1/70th the energy of refining aluminum from ore, which illustrates the tremendous energy savings that recycling provides.)

Another application of electrolysis is electroplating, which is the deposition of a thin layer of metal on an object for protective or decorative purposes (Figure 7.6.2). Essentially, a metal object is connected to the cathode of an electrolytic cell and immersed in a solution of a particular metal cation. When the electrolytic cell is operated, a thin coating of the metal cation is reduced to the elemental metal on the surface of the object; the thickness of the coating can be as little as a few micrometers (10^{-6} m). Jewelry, eating utensils, electrical contacts, and car parts like bumpers are common items that are electroplated. Gold, silver, nickel, copper, and chromium are common metals used in electroplating.

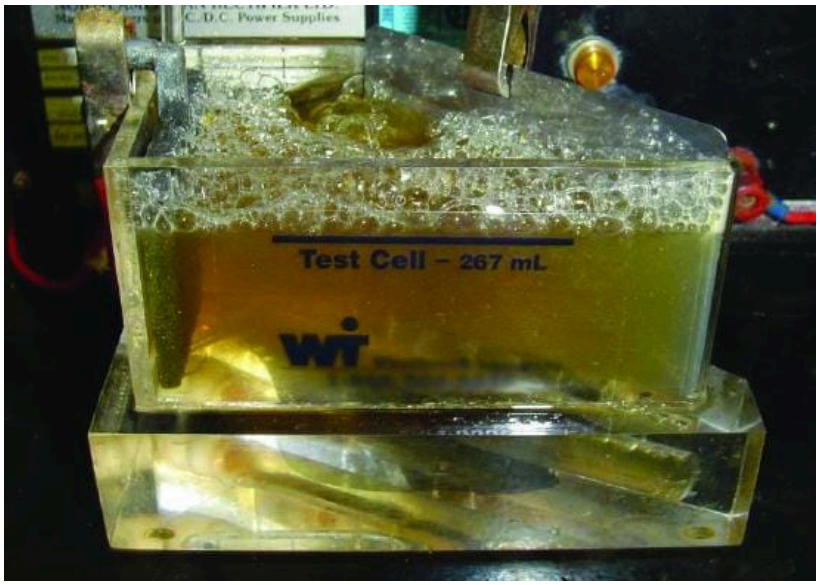


Figure 7.6.2: Electroplating Source. Photo courtesy of Fstep, http://commons.wikimedia.org/wiki/File:Hullcell.jpg(opens in new window).

Key Takeaways

- Electrolysis is a forced non-spontaneous redox occurence by the introduction of electricity into a cell from an outside source.
- Electrolysis is used to isolate elements and electroplate objects.

Answers

Exercise 7.6.1 **Copper-Zinc Galvanic Cell**

From the information given in the problem:

$$\text{anode (oxidation): } \mathrm{Zn}(s) \longrightarrow \mathrm{Zn}^{2+}(aq) + 2\,\mathrm{e}^-$$

$$\text{cathode (reduction): } \mathrm{Cu}^{2+}(aq) + 2\,\mathrm{e}^- \longrightarrow \mathrm{Cu}(s)$$

$$\text{overall: } \mathrm{Zn}(s) + \mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Cu}(s)$$

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7.7: Nuclear Chemistry

Learning Objective

- Define and give examples of the major types of radioactivity.

We saw in Chapter 3 that atoms are composed of subatomic particles—protons, neutrons, and electrons. Protons and neutrons are located in the nucleus and provide most of the mass of an atom, while electrons circle the nucleus in shells and subshells and account for an atom's size.

In Chapter 2, we also introduced the notation for succinctly representing an isotope of a particular atom:

$$^{12}_{6}\mathrm{C}$$

The element in this example, represented by the symbol C, is carbon. Its atomic number, 6, is the subscript next to the symbol and is the number of protons in the atom. The mass number, the superscript next to the symbol, is the sum of the number of protons and neutrons in the nucleus of this particular isotope. In this case, the mass number is 12, which means that the number of neutrons in the atom is 12 – 6 = 6 (that is, the mass number of the atom minus the number of protons in the nucleus equals the number of neurons). Occasionally, the atomic number is omitted in this notation because the symbol of the element itself conveys its characteristic atomic number. The two isotopes of hydrogen—^{2}H and ^{3}H—are given their own names and symbols: deuterium (D) and tritium (T), respectively.

Atomic theory in the nineteenth century presumed that nuclei had fixed compositions. But in 1896, the French scientist Henri Becquerel found that a uranium compound placed near a photographic plate made an image on the plate, even if the compound was wrapped in black cloth. He reasoned that the uranium compound was emitting some kind of radiation that passed through the cloth to expose the photographic plate. Further investigations showed that the radiation was a combination of particles and electromagnetic rays, with its ultimate source being the atomic nucleus. These emanations were ultimately called, collectively, **radioactivity**.

There are three main forms of radioactive emissions. The first is called an **alpha particle**, which is symbolized by the Greek letter α. An alpha particle is composed of 2 protons and 2 neutrons and is the same as a helium nucleus. We often use:

$$^{4}_{2}\mathrm{He}$$

to represent an alpha particle. It has a 2+ charge. When a radioactive atom emits an alpha particle, the original atom's atomic number decreases by 2 (because of the loss of 2 protons), and its mass number decreases by 4 (because of the loss of 4 nuclear particles). We can represent the emission of an alpha particle with a chemical equation. For example, the alpha-particle emission of uranium-235 is as follows:

$$^{235}_{92}\mathrm{U} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{231}_{90}\mathrm{Th}$$

Rather than calling this equation a chemical equation, we call it a **nuclear equation** to emphasize that the change occurs in an atomic nucleus. How do we know that a product of this reaction is \(\ce{^{90}_{231}T}\)? We use the law of conservation of matter, which says that matter cannot be created or destroyed. This means that we must have the same number of protons and neutrons on both sides of the nuclear equation. If our uranium nucleus loses 2 protons, there are 90 protons remaining, identifying the element as thorium. Moreover, if we lose four nuclear particles of the original 235, there are 231 remaining. So we use subtraction to identify the isotope of the Th atom—in this case,

$$^{231}_{90}\mathrm{Th}$$

Chemists often use the names **parent isotope** and **daughter isotope** to represent the original atom and the product other than the alpha particle. In the previous example,

$$^{235}_{92}\mathrm{U}$$

is the parent isotope, and

$$^{231}_{90}\mathrm{Th}$$

is the daughter isotope. When one element changes into another in this manner, it undergoes **radioactive decay**.

Example 7.7.1

Write the nuclear equation that represents the radioactive decay of radon-222 by alpha particle emission and identify the daughter isotope.

Solution

Radon has an atomic number of 86, so the parent isotope is represented as

$$^{222}_{86}\mathrm{Rn}$$

.

We represent the alpha particle as

$$^{4}_{2}\mathrm{He}$$

Use subtraction (222 – 4 = 218 and 86 – 2 = 84) to identify the daughter isotope as polonium:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{218}_{84}\mathrm{Th}$$

Exercise 7.7.2

Write the nuclear equation that represents radioactive decay of polonium-208 by alpha particle emission and identify the daughter isotope.

Answer available at the end of this section.

The second major type of radioactive emission is called a **beta particle**, symbolized by the Greek letter β. A beta particle is an electron ejected from the nucleus (not from the shells of electrons about the nucleus) and has a 1– charge. We can also represent a beta particle as e–10. The net effect of beta particle emission on a nucleus is that a neutron is converted to a proton. The overall mass number stays the same, but because the number of protons increases by one, the atomic number goes up by one. Carbon-14 decays by emitting a beta particle:

$$^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\mathrm{e}$$

Again, the sum of the atomic numbers is the same on both sides of the equation, as is the sum of the mass numbers. (Note that the electron is assigned an "atomic number" of –1, equal to its charge.)

The third major type of radioactive emission is not a particle but rather a very energetic form of electromagnetic radiation called **gamma rays**, symbolized by the Greek letter γ. Gamma rays themselves do not carry an overall electrical charge, but they may knock electrons out of atoms in a sample of matter and make it electrically charged (for which gamma rays are termed *ionizing radiation*). For example, in the radioactive decay of radon-222, both alpha and gamma radiation are emitted, with the latter having an energy of 8.2×10^{-14} J per nucleus decayed:

$$^{222}_{86}\mathrm{Rn} \rightarrow {}^{218}_{84}\mathrm{Po} + {}^{4}_{2}\mathrm{He} + \gamma$$

This may not seem like much energy, but if 1 mol of Rn atoms were to decay, the gamma ray energy would be 4.9×10^{7} kJ!

Example 7.7.3

Write the nuclear equation that represents the radioactive decay of boron-12 by beta particle emission and identify the daughter isotope. A gamma ray is emitted simultaneously with the beta particle.

Solution

The parent isotope is B-12,

$$^{12}_{5}\mathrm{B}$$

while one of the products is B-12,

$$^{0}_{-1}\mathrm{e}$$

So that the mass and atomic numbers have the same value on both sides, the mass number of the daughter isotope must be 12, and its atomic number must be 6. The element having an atomic number of 6 is carbon. Thus the complete nuclear equation is as follows:

$$^{12}_{5}\mathrm{B} \rightarrow {}^{12}_{6}\mathrm{C} + {}^{0}_{-1}\mathrm{e} + \gamma$$

The daughter isotope is carbon-12.

Exercise 7.7.4

Answer available at the end of this section.

Alpha, beta, and gamma emissions have different abilities to penetrate matter. The relatively large alpha particle is easily stopped by matter (although it may impart a significant amount of energy to the matter it contacts). Beta particles penetrate slightly into matter, perhaps a few centimeters at most. Gamma rays can penetrate deeply into matter and can impart a large amount of energy into the surrounding matter. Table 7.7.1 summarizes the properties of the three main types of radioactive emissions.

Table 7.7.1 The Three Main Forms of Radioactive Emissions

Characteristic	Alpha Particles	Beta Particles	Gamma Rays
symbols	α, H24	β $^{0}_{-1}\mathrm{e}$	γ
identity	helium nucleus	electron	electromagnetic radiation
charge	2+	1–	none
mass number	4	0	0
penetrating power	minimal (will not penetrate skin)	short (will penetrate skin and some tissues slightly)	deep (will penetrate tissues deeply)

Occasionally, an atomic nucleus breaks apart into smaller pieces in a radioactive process called **spontaneous fission (or fission)**. Typically, the daughter isotopes produced by fission are a varied mix of products, rather than a specific isotope as with alpha and beta particle emission. Often, fission produces excess neutrons that will sometimes be captured by other nuclei, possibly inducing additional radioactive events. Uranium-235 undergoes spontaneous fission to a small extent. One typical reaction is

$$^{235}_{92}\mathrm{Tc} \rightarrow {}^{139}_{56}\mathrm{Ba} + {}^{94}_{36}\mathrm{Kr} + 2\,{}^{1}_{0}\mathrm{n}$$

where $^{1}_{0}\mathrm{n}$ is a neutron. As with any nuclear process, the sums of the atomic numbers and mass numbers must be the same on both sides of the equation. Spontaneous fission is found only in large nuclei. The smallest nucleus that exhibits spontaneous fission is lead-208. (Fission is the radioactive process used in nuclear power plants and one type of nuclear bomb.)

Key Takeaways

- The major types of radioactivity include alpha particles, beta particles, and gamma rays.
- Fission is a type of radioactivity in which large nuclei spontaneously break apart into smaller nuclei.

Learning Objective

- Understand the significance of the discovery of X-rays and radioactivity to the modern view of the structure of the atom.

The discovery of radioactivity took place over several years beginning with the discovery of x-rays in 1895 by Wilhelm Conrad Roentgen and continuing with such people as Henri Becquerel and the Curie family. The application of x-rays and radioactive materials is far reaching in medicine and industry. Radioactive material is used in everything from nuclear reactors to isotope infused saline solutions. These technologies allow us to utilize great amounts of energy and observe biological systems in ways which were unthinkable less than a century ago. As will be discussed in section3.4, the British physicist Ernest Rutherford (1871-1937) conducted experiments (based on the Curies' work) that led to the modern view of the structure of the atom.

Roentgen: The Discovery of X-Rays

From Section 3.1, we saw how J. J. Thomson discovered the electronundefined by using cathode ray tubes (CRT). In November of 1895, a German physicist named Wilhelm Roentgen (Röntgen) used this same technology to discover the x-ray. Roentgen's ray originated from the CRT and reacted with a barium platinocyanide screen to fluoresce. After seeing this, Roentgen placed objects between this ray and the screen. He noted that the ray would still penetrate substances and leave an image on photographic film.

When naming this type of radiation, Roentgen reflected back to his experiments. He could not confidently explain his observations. For this reason, he decided to call this new type of radiation the x-ray. Throughout his lifetime, he continue to explore this technology and even used his wife as a test subject (refer to chapter 1 for her x-ray image). By never patenting this invention, Roentgen freely shared the x-ray with other researchers and medical professionals. In 1901, he was awarded the Nobel Prize in physics for his work with the x-ray.

As with any new technology, people looked for ways to apply the x-ray to day to day life. Thomas Edison, the inventor of the light bulb, thought the average person should have an x-ray machine in their home. He designed x-ray machines to be smaller and portable. Unfortunately, many of his technicians died from radiation poisoning.



Figure 7.7.1

Hand mit Ringen (Hand with Rings): print of Wilhelm Röntgen's first "medical" X-ray, of his wife's hand, taken on 22 December 1895.

Figure 7.7.2***:*** *(left) Photo of Wilhelm Conrad Roengten, the discoverer of x-rays. (right) Surgical operation during World War I using a fluoroscope that used x-rays to find embedded bullets*

During Edison's time, people would host x-ray parties in their homes. The host of these gatherings would allow guests to x-ray different parts of their body. People would leave these events with their framed souvenirs. X-rays were even used to measure foot size. Devices like the fluoroscope were placed in shoe stores. Technicians, consumers, and observers could look in oculars while the foot was being x-rayed. Unnecessary exposure to radiation was not considered as a hazard during these times. In this era, most believed this was the most accurate way to measure the foot.

The harmful effects and different uses of X-rays

In November 1896, an article entitled "the harmful effects of X-rays" was published in Nature. The witness was an X-ray demonstrator during the summer in London. He therefore paid for himself throughout the summer at the rate of several hours per day of exposure. He testified: "In the first two or three weeks I felt no inconvenience, but after a while appeared on the fingers of my right hand many dark spots which pierced under the skin, and gradually they became very painful; the rest of the skin was red and strongly inflamed My hand was so bad that I was constantly forced to bathe it in very cold water An ointment momentarily calm the pain but the epidermis had dried up had become hard and yellow like parchment and completely insensible, so I was not surprised when my hand began to peel. "

"Soon the skin and nails fall off and the fingers swelled, the pain remaining constant. I lost the skin of my right and left hands, and four of my nails have disappeared from the right hand and two left and three others were ready to fall off. During more than six weeks I was unable to hold anything in my right hand and I can not hold a pen since the loss of my nails ..."

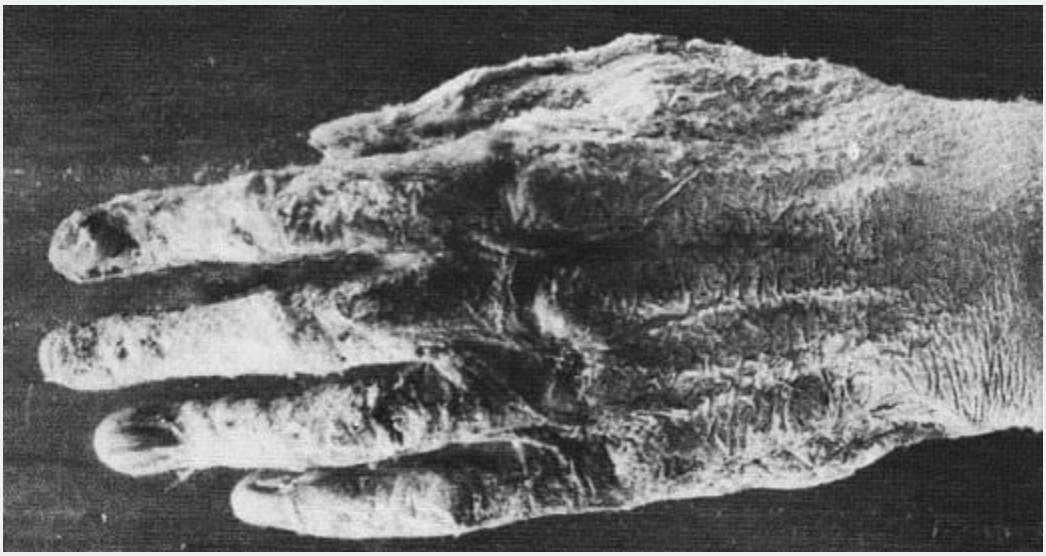


Figure 7.7.3: Radiation poisoning due to over exposure of x-rays.

Different applications use different parts of the X-ray spectrum.

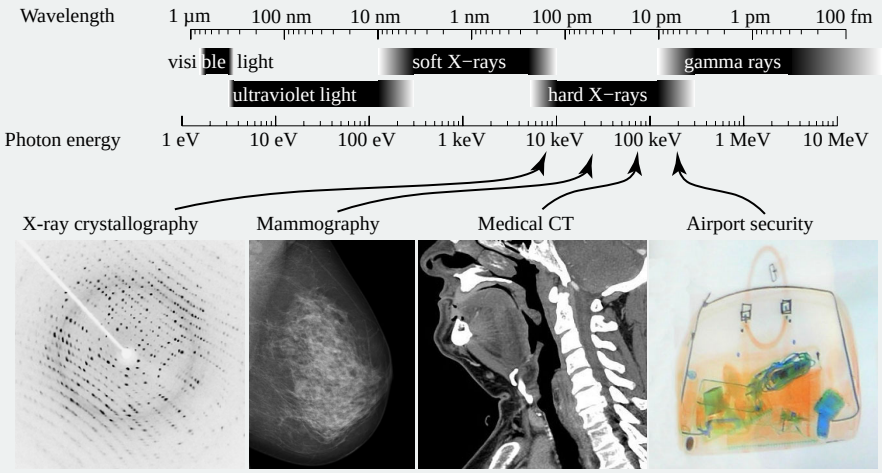
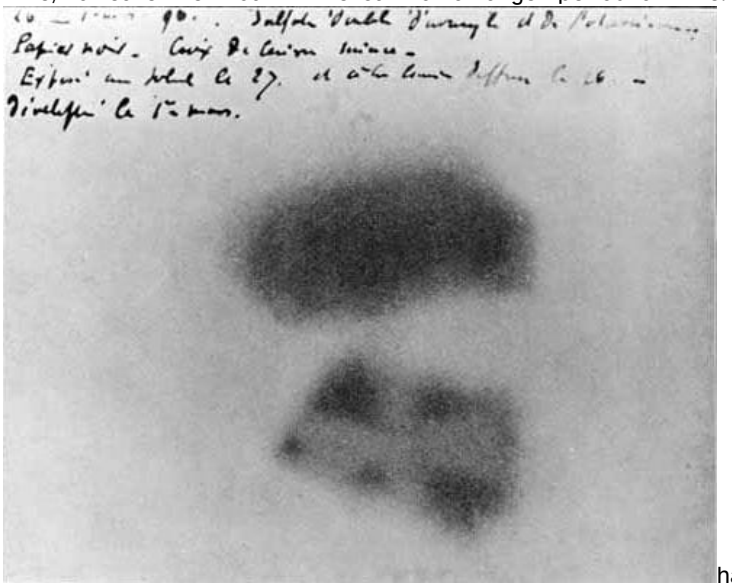


Figure 7.7.4: X-rays are part of the electromagnetic spectrum, with wavelengths shorter than visible light.

The Discovery of Radioactivity

When Becquerel heard about Roentgen's discovery, he wondered if his fluorescent minerals would give the same x-rays. Becquerel placed some of his rock crystals on top of a well-covered photographic plate and sat them in the sunlight. The sunlight made the crystals glow with a bright fluorescent light, but when Becquerel developed the film he was very disappointed. He found that only one of his minerals, a uranium salt, had fogged the photographic plate. He decided to try again, and this time, to leave them out in the sun for a longer period of time. Fortunately, the weather did not cooperate and Becquerel



had to leave the crystals and film stored in a drawer for several cloudy days. Before continuing his experiments, Becquerel decided to check one of the photographic plates to make sure the chemicals were still good. To his amazement, he found that the plate had been exposed in spots where it had been near the uranium containing rocks and some of these rocks had not been exposed to sunlight at all. In later experiments, Becquerel confirmed that the radiation from the uranium had no connection with light or fluorescence, but the amount of radiation was directly proportional to the concentration of uranium in the rock. Becquerel had discovered **radioactivity**.

Figure $7.7.5$***:*** *Image of Becquerel's photographic plate which has been fogged by exposure to radiation from a uranium salt. The shadow of a metal Maltese Cross placed between the plate and the uranium salt is clearly visible (Public Domain).*

The Curies and Radium

One of Becquerel's assistants, a young Polish scientist named Maria Sklowdowska (to become Marie Curie after she married Pierre Curie), became interested in the phenomenon of radioactivity. With her husband, she decided to find out if chemicals other than uranium were radioactive. The Austrian government was happy to send the Curies a ton of pitchblende from the mining region of Joachimstahl because it was waste material that had to be disposed of anyway. The Curies wanted the pitchblende because it was the residue of uranium mining. From the ton of pitchblende, the Curies separated $0.10\ \text{g}$ of a previously unknown element, radium, in the form of the compound radium chloride. This radium was many times more radioactive than uranium.

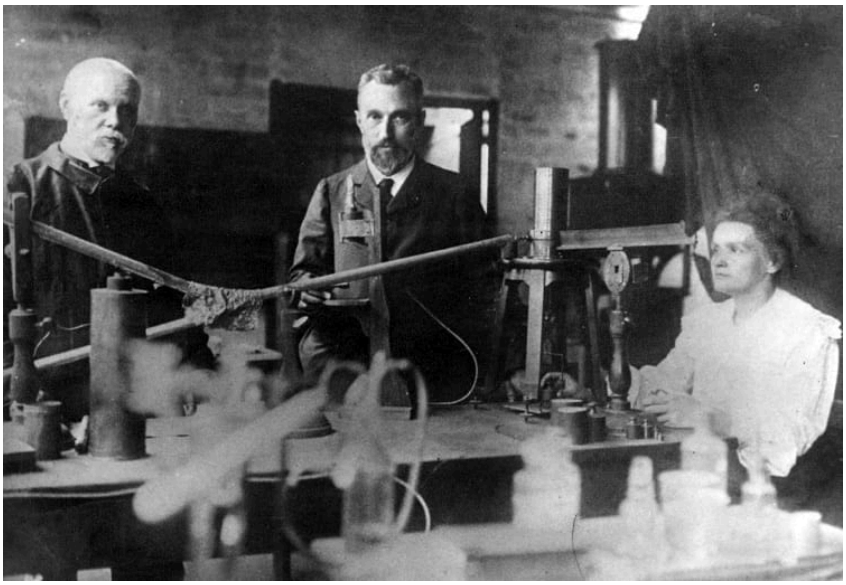


Figure 7.7.6: Marie (left) and Pierre Curie (middle) with Henri Becquerel (right) shared the 1903 Nobel Prize

By 1902, the world was aware of a new phenomenon called radioactivity and of new elements which exhibited natural radioactivity. For this work, Becquerel and the Curies shared the 1903 Nobel Prize in Physics. These three researchers continued to work with hazardous radioactive materials. They experienced various ailments included burns and weakness by touching these substances. In 1906, Pierre Curie accidently fell under a horse drawn cart. He immediately died from his wounds and left Marie with two small children (Irene, 9 years old and Eve, 2 years old). After his death, Marie continued her research by identifying a new element, Polonium. In 1911, she was awarded the Nobel Prize in chemistry for the discoveries of radium and polonium. As of today, Marie Curie is the only person ever to have received two Nobel Prizes in the sciences.

YouTube Video

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video7.7.2: Everyone knows that winning the Nobel Prize is a big deal, but why do we even have a Nobel Prize? And why does it matter?

Summary

- Henri Becquerel, Marie Curie, and Pierre Curie shared the discovery of radioactivity.
- Wilhelm Roengtgen used cathode ray tubes to discover the x-ray.

Contributors and Attributions

- Wikibooks
-
- Elizabeth R. Gordon (Furman University)

Answers

Exercise 7.7.2

$$^{208}_{80}\mathrm{Po} \rightarrow {}^{4}_{2}\mathrm{He} + {}^{204}_{82}\mathrm{Pb}$$

$$^{204}_{82}\mathrm{Pb}$$

Exercise 7.7.4

$$^{133}_{43}\mathrm{Tc} \rightarrow {}^{133}_{44}\mathrm{Ru} + {}^{0}_{-1}\mathrm{e} + \gamma$$

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7.8: Uses and Impacts of Radiation

Learning Objectives

- Know the different sources of background radiation.
- Describe the biological impact of ionizing radiation
- List common sources of radiation exposure in the US.

Background Radiation

We are all exposed to a small amount of radiation in our daily lives. This **background radiation** comes from naturally occurring sources and from human-produced radiation. Exposure to x-rays and nuclear medicine isotopes, ground sources, and cosmic radiation account for almost half of the background exposure of the average American. Radon gas, formed from the decay of uranium and thorium isotopes, is responsible for a little over half the total amount of background radiation. See the table below for background sources.

Table 7.8.1 *Sources of Background Radiation*

radon	54%
consumer products	3%
nuclear medicine	4%
cosmic radiation	8%
terrestrial	8%
internal	11%
x-rays	11%
other	1%

The Problem of Radon

For many people, one of the largest sources of exposure to radiation is from radon gas (Rn-222). Radon-222 is an α emitter with a half–life of 3.82 days. It is one of the products of the radioactive decay series of U-238, which is found in trace amounts in soil and rocks. The radon gas that is produced slowly escapes from the ground and gradually seeps into homes and other structures above. Since it is about eight times more dense than air, radon gas accumulates in basements and lower floors, and slowly diffuses throughout buildings (Figure 7.8.1).

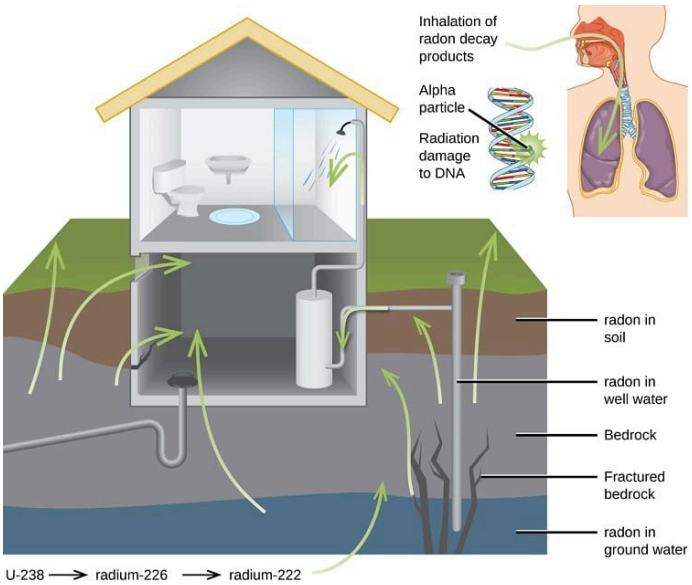


Figure 7.8.1: Radon-222 seeps into houses and other buildings from rocks that contain uranium-238, a radon emitter. The radon enters through cracks in concrete foundations and basement floors, stone or porous cinderblock foundations, and openings for water and gas pipes.

Radon is found in buildings across the country, with amounts depending on where you live. The average concentration of radon inside houses in the US (1.25 pCi/L) is about three times the levels found in outside air, and about one in six houses have radon levels high enough that remediation efforts to reduce the radon concentration are recommended. Exposure to radon increases one's risk of getting cancer (especially lung cancer), and high radon levels can be as bad for health as smoking a carton of cigarettes a day. Radon is the number one cause of lung cancer in nonsmokers and the second leading cause of lung cancer overall. Radon exposure is believed to cause over 20,000 deaths in the US per year.

Measuring Radiation Exposure

Several different devices are used to detect and measure radiation, including Geiger counters, scintillation counters (scintillators), and radiation dosimeters (Figure 7.8.2). Probably the best-known radiation instrument, the Geiger counter (also called the Geiger-Müller counter) detects and measures radiation. Radiation causes the ionization of the gas in a Geiger-Müller tube. The rate of ionization is proportional to the amount of radiation. A scintillation counter contains a scintillator—a material that emits light (luminesces) when excited by ionizing radiation—and a sensor that converts the light into an electric signal. Radiation dosimeters also measure ionizing radiation and are often used to determine personal radiation exposure. Commonly used types are electronic, film badge, thermoluminescent, and quartz fiber dosimeters.

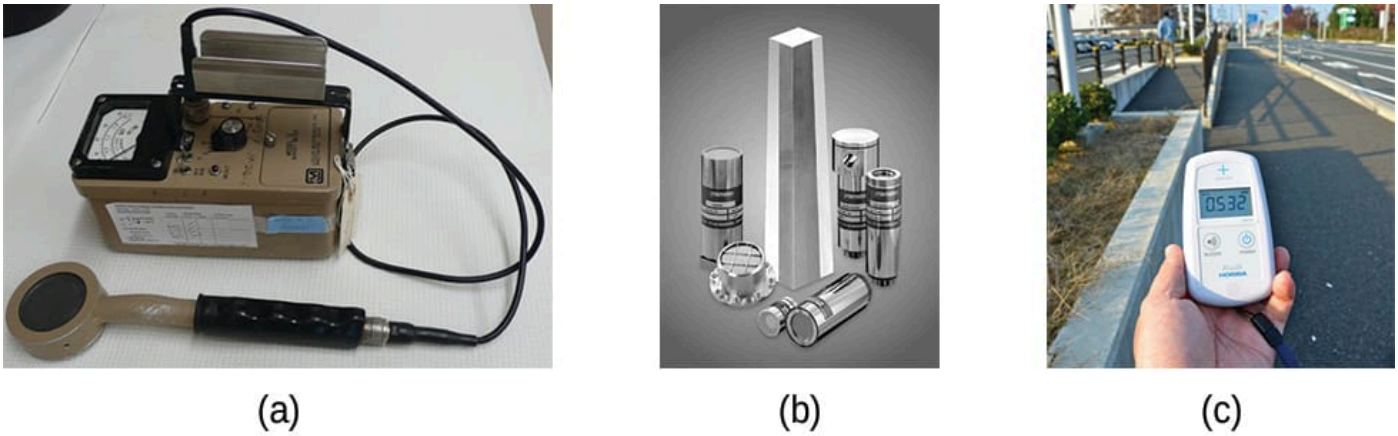


Figure 7.8.2: Devices such as (a) Geiger counters, (b) scintillators, and (c) dosimeters can be used to measure radiation. (credit c: modification of work by "osaMu"/Wikimedia commons).

Radiation Damage to Cells

The increased use of radioisotopes has led to increased concerns over the effects of these materials on biological systems (such as humans). All radioactive nuclides emit high-energy particles or electromagnetic waves. When this radiation encounters living cells, it can cause heating, break chemical bonds, or ionize molecules. The most serious biological damage results when these radioactive emissions fragment or ionize molecules. For example, alpha and beta particles emitted from nuclear decay reactions possess much higher energies than ordinary chemical bond energies. When these particles strike and penetrate matter, they produce ions and molecular fragments that are extremely reactive. The damage this does to biomolecules in living organisms can cause serious malfunctions in normal cell processes, taxing the organism's repair mechanisms and possibly causing illness or even death (Figure 7.8.3).

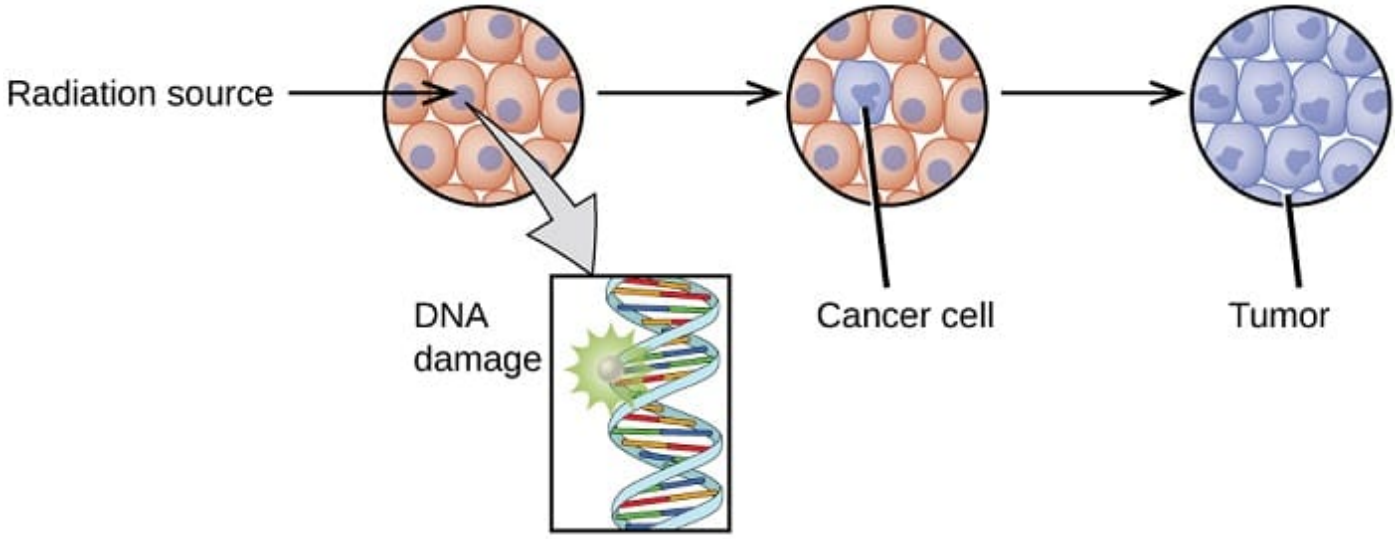


Figure 7.8.3: Radiation can harm biological systems by damaging the DNA of cells. If this damage is not properly repaired, the cells may divide in an uncontrolled manner and cause cancer.

Radiation can harm either the whole body (somatic damage) or eggs and sperm (genetic damage). Its effects are more pronounced in cells that reproduce rapidly, such as the stomach lining, hair follicles, bone marrow, and embryos. This is why patients undergoing radiation therapy often feel nauseous or sick to their stomach, lose hair, have bone aches, and so on, and why particular care must be taken when undergoing radiation therapy during pregnancy.

Ionizing vs. Nonionizing Radiation

There is a large difference in the magnitude of the biological effects of nonionizing radiation (for example, light and microwaves) and ionizing radiation, emissions energetic enough to knock electrons out of molecules (for example, α and β particles, γ rays, X-rays, and high-energy ultraviolet radiation) (Figure 7.8.4).

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Figure 7.8.4: Lower frequency, lower-energy electromagnetic radiation is nonionizing, and higher frequency, higher-energy electromagnetic radiation is ionizing.

Energy absorbed from nonionizing radiation speeds up the movement of atoms and molecules, which is equivalent to heating the sample. Although biological systems are sensitive to heat (as we might know from touching a hot stove or spending a day at the beach in the sun), a large amount of nonionizing radiation is necessary before dangerous levels are reached. Ionizing radiation, however, may cause much more severe damage by breaking bonds or removing electrons in biological molecules, disrupting their structure and function. The damage can also be done indirectly, by first ionizing H_2O (the most abundant molecule in living organisms), which forms a H_2O^+ ion that reacts with water, forming a hydronium ion and a hydroxyl radical:

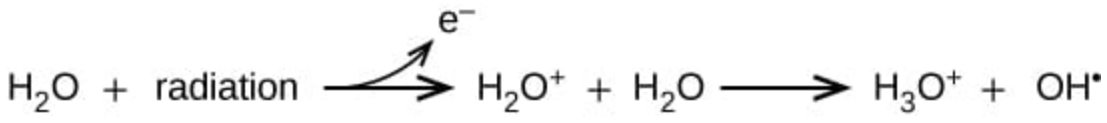


Figure 7.8.5.

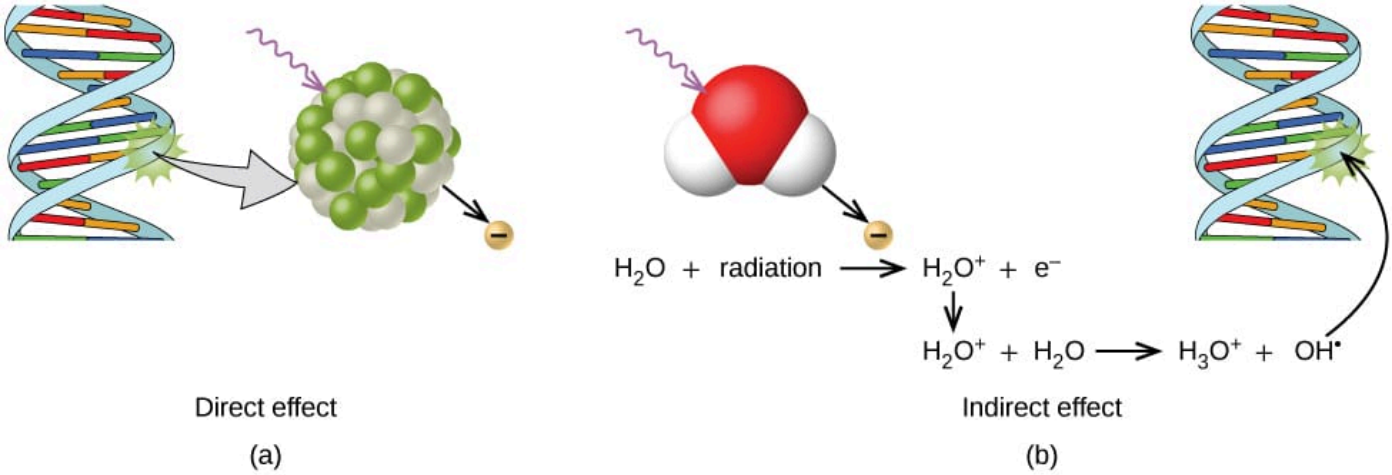


Figure 7.8.5: Ionizing radiation can (a) directly damage a biomolecule by ionizing it or breaking its bonds, or (b) create an H_2O^+ ion, which reacts with H_2O to form a hydroxyl radical, which in turn reacts with the biomolecule, causing damage indirectly.

Effects of Long-term Radiation Exposure on the Human Body

The effects of radiation depend on the type, energy, and location of the radiation source, and the length of exposure. As shown in Figure 7.8.6, the average person is exposed to background radiation, including cosmic rays from the sun and radon from uranium in the ground; radiation from medical exposure, including CAT scans, radioisotope tests, X-rays, and so on; and small amounts of radiation from other human activities, such as airplane flights (which are bombarded by increased numbers of cosmic rays in the upper atmosphere), radioactivity from consumer products, and a variety of radionuclides that enter our bodies when we breathe (for example, carbon-14) or through the food chain (for example, potassium-40, strontium-90, and iodine-131).

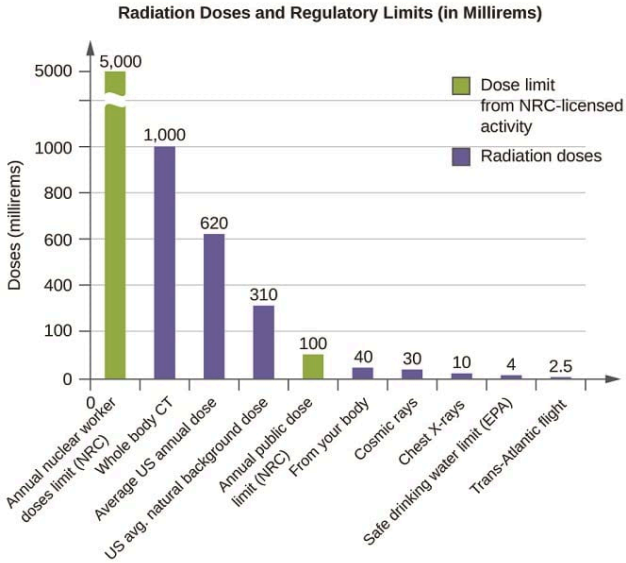


Figure 7.8.6 The total annual radiation exposure for a person in the US is about 620 mrem. The various sources and their relative amounts are shown in this bar graph. (source: U.S. Nuclear Regulatory Commission).

A short-term, sudden dose of a large amount of radiation can cause a wide range of health effects, from changes in blood chemistry to death. Short-term exposure to tens of rems of radiation will likely cause very noticeable symptoms or illness; a

dose of about 500 rems is estimated to have a 50% probability of causing the death of the victim within 30 days of exposure. Exposure to radioactive emissions has a cumulative effect on the body during a person's lifetime, which is another reason why it is important to avoid any unnecessary exposure to radiation. Health effects of short-term exposure to radiation are shown in Table 7.8.2.

Table 7.8.2 Health Effects of Radiation

Exposure (rem)	Health Effect	Time to Onset (without treatment)
5–10	changes in blood chemistry	—
50	nausea	hours
55	fatigue	—
70	vomiting	—
75	hair loss	2–3 weeks
90	diarrhea	—
100	hemorrhage	—
400	possible death	within 2 months
1000	destruction of intestinal lining	—
	internal bleeding	—
	death	1–2 weeks
2000	damage to central nervous system	—
	loss of consciousness;	minutes
	death	hours to days

It is impossible to avoid some exposure to ionizing radiation. We are constantly exposed to background radiation from a variety of natural sources, including cosmic radiation, rocks, medical procedures, consumer products, and even our own atoms. We can minimize our exposure by blocking or shielding the radiation, moving farther from the source, and limiting the time of exposure.

Summary

- Background radiation is defined and different sources of background radiation are listed.
- We are constantly exposed to radiation from a variety of naturally occurring and human-produced sources.
- Various devices, including Geiger counters, scintillators, and dosimeters, are used to detect and measure radiation, and monitor radiation exposure.
- Ionizing radiation is the most harmful because it can ionize molecules or break chemical bonds, which damages the molecule and causes malfunctions in cell processes. It can also create reactive hydroxyl radicals that damage biological molecules and disrupt physiological processes.

Contributors and Attributions

-

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- Marisa Alviar-Agnew (Sacramento City College)

Learning Objectives

- Understand how radiation is used in agriculture and various industries.
- Understand the difference between diagnostic and therapeutic radiation.

Radioactive isotopes have the same chemical properties as stable isotopes of the same element, but they emit radiation, which can be detected. If we replace one (or more) atom(s) with radioisotope(s) in a compound, we can track them by monitoring their radioactive emissions. This type of compound is called a radioactive tracer (or radioactive label). Radioisotopes are used to follow the paths of biochemical reactions or to determine how a substance is distributed within an organism.

Radioisotopes in Industry and Agriculture

Radioisotopes (radioactive isotopes or radionuclides or radioactive nulcides) are used in two major ways: either for their radiation alone (irradiation, nuclear batteries) or for the combination of chemical properties and their radiation (tracers, biopharmaceuticals).

Tracer Applications

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A **tracer** is a substance that can be used to follow the pathway of that substance through some structure. For instance, leaks in underground water pipes can be discovered by running some **tritium (H-3)**-containing water through the pipes and then using a Geiger counter to locate any radioactive tritium subsequently present in the ground around the pipes. (Recall that tritium is a radioactive isotope of hydrogen.)

Tracers are used in diverse ways to study the mechanisms of chemical reactions in plants and animals. These include labeling fertilizers in studies of nutrient uptake by plants and crop growth, investigations of digestive and milk-producing processes in cows, and studies on the growth and metabolism of animals and plants.

For example, the radioisotope **C-14** was used to elucidate the details of how photosynthesis occurs. The overall reaction is:

$$6\,\mathrm{CO_2}(g) + 6\,\mathrm{H_2O}(l) \longrightarrow \mathrm{C_6H_{12}O_6}(s) + 6\,\mathrm{O_2}(g),$$

but the process is much more complex, proceeding through a series of steps in which various organic compounds are produced. In studies of the pathway of this reaction, plants were exposed to CO_2 containing a high concentration of $^{14}_{6}\mathrm{C}$. At regular intervals, the plants were analyzed to determine which organic compounds contained carbon-14 and how much of each compound was present. From the time sequence in which the compounds appeared and the amount of each present at given time intervals, scientists learned more about the pathway of the reaction.

Phosphorus-32 is used in plant sciences for tracking a plant's uptake of fertilizer from the roots to the leaves. The phosphorus-32-labelled fertilizer is given to the plant hydroponically or via water in the soil and the usage of the phosphorus can be mapped from the emitted beta radiation. The information gathered by mapping the fertilizer uptake show how the plant takes up and uses the phosphorus from fertilizer.

Irradiation of Food and Mail

Co-60 (gamma source) and x-rays are use to irradiate many foods in the United States. Ionizing radiation can be used to kill food borne illnesses like salmonella and e coli. Irradiating food can also prolong shelf-life, delay ripening, and destroy insects. In addition, gamma/x-ray can sterilize foods making refrigeration unnecessary. Often, sterilized food is served to hospital patients who have impaired immune systems.

In the United States, irradiation of food is regulated by the FDA (Food and Drug Administration). Foods that have been irradiated must display the symbol called the radura on their packaging. Some foods that could display this symbol are: meats (beef, pork, chicken), shrimp, lobster, fruits, vegetables, shellfish, and spices. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.



Figure 7.8.1 The preservation of strawberries with Ionizing Radiation.

In the fall of 2001, anthrax laced letters (Figure 7.8.2) were sent to various new agencies and two United States senators. Anthrax is an infectious disease caused by bacteria. It exists naturally in some soils and can be isolated in a laboratory.



Figure 7.8.2 Anthrax-laced letter that was sent to Senator Tom Daschle. Image taken from: https://upload.wikimedia.org/wikiped...hle_letter.jpg

To combat this form of bioterrorism, the United States Postal Service(USPS)and the Federal Bureau of Investigation (FBI) installed x-ray generators to irradiate suspicious looking mail sent to some governmental facilities. X-rays will kill the majority of this bacteria and some viruses as well. Once irradiated, mail (packages and letters) might change in color, texture, and smell. This ionizing radiation chemically alters the composition of the mail's paper component. The x-rays do not leave the mail radioactive.

Smoke Detectors

Americium-241, an α emitter with a half-life of 458 years, is used in tiny amounts in ionization-type smoke detectors (Figure 7.8.3). The α emissions from Am-241 ionize the air between two electrode plates in the ionizing chamber. A battery supplies a potential that causes movement of the ions, thus creating a small electric current. When smoke enters the chamber, the movement of the ions is impeded, reducing the conductivity of the air. This causes a marked drop in the current, triggering an alarm.

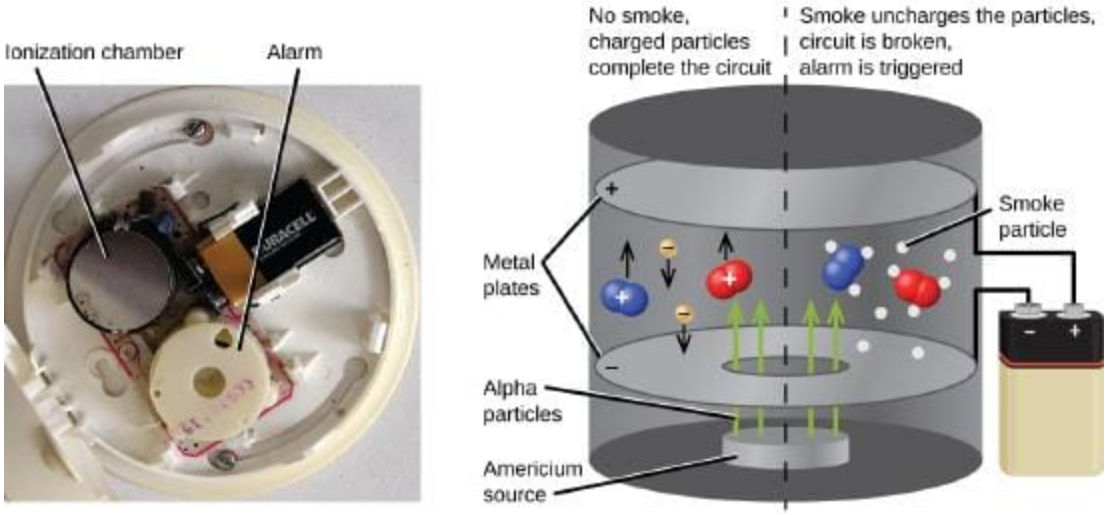


Figure 7.8.3 Inside a smoke detector, Am-241 emits α particles that ionize the air, creating a small electric current. During a fire, smoke particles impede the flow of ions, reducing the current and triggering an alarm. (credit a: modification of work by “Muffet”/Wikimedia Commons)

Other Applications

Commercial applications of radioactive materials are equally diverse . They include determining the thickness of films, paper, and thin metal sheets by exploiting the penetration power of various types of radiation(Figure 7.8.4). Flaws in metals used for structural purposes can be detected using high-energy gamma rays from cobalt-60 in a fashion similar to the way X-rays are used to examine the human body. In one form of pest control, flies are controlled by sterilizing male flies with γ radiation so that females breeding with them do not produce offspring.

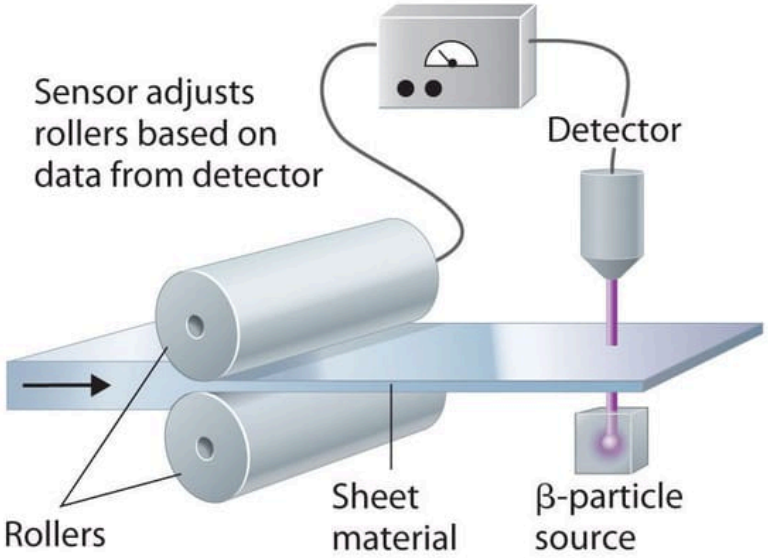


Figure 7.8.4 Using radiation to control the thickness of a material. Because the amount of radiation absorbed by a material is proportional to its thickness, radiation can be used to control the thickness of plastic film, tin foil, or paper. As shown, a beta emitter is placed on one side of the material being produced and a detector on the other. An increase in the amount of radiation that reaches the detector indicates a decrease in the thickness of the material and vice versa. The output of the detector can thus be used to control the thickness of the material.

Radioisotopes in Medicine

Radioactive tracers are also used in many medical applications, including both diagnosis and treatment. They are used to measure engine wear, analyze the geological formation around oil wells, and much more.

Radioisotopes have revolutionized medical practice, where they are used extensively. Over 10 million nuclear medicine procedures and more than 100 million nuclear medicine tests are performed annually in the United States.

Diagnostic Medical Applications

Diagnostic medical applications involve testing for a disease or condition. In nuclear medicine, this could involve using PET scans, or isotopic studies. The radiation involved for each of these types of tools will vary in mrem or mSv amounts.

PET Scanning

Positron Emission Tomography or PET scan is a type of nuclear medicine imaging. Depending on the area of the body being imaged, a radioactive isotope is either injected into a vein, swallowed by mouth, or inhaled as a gas. When the radioisotope is collected in the appropriate area of the body, the gamma ray emissions are detected by a PET scanner (often called a gamma camera) which works together with a computer to generate special pictures, providing details on both the structure and function of various organs. Watch this informational video on how this technique works.

PET Scan

This interactive element is not available in this format. Visit the link or scan the QR code to access it online.

View this YouTube Video online

Video 7.8.1 *PET scans: What to expect.*

PET scans are used to:

- Detect cancer
- Determine the amount of cancer spread
- Assess the effectiveness of treatment plans
- Determine blood flow to the heart muscle
- Determine the effects of a heart attack
- Evaluate brain abnormalities such as tumors and memory disorders
- Map brain and heart function

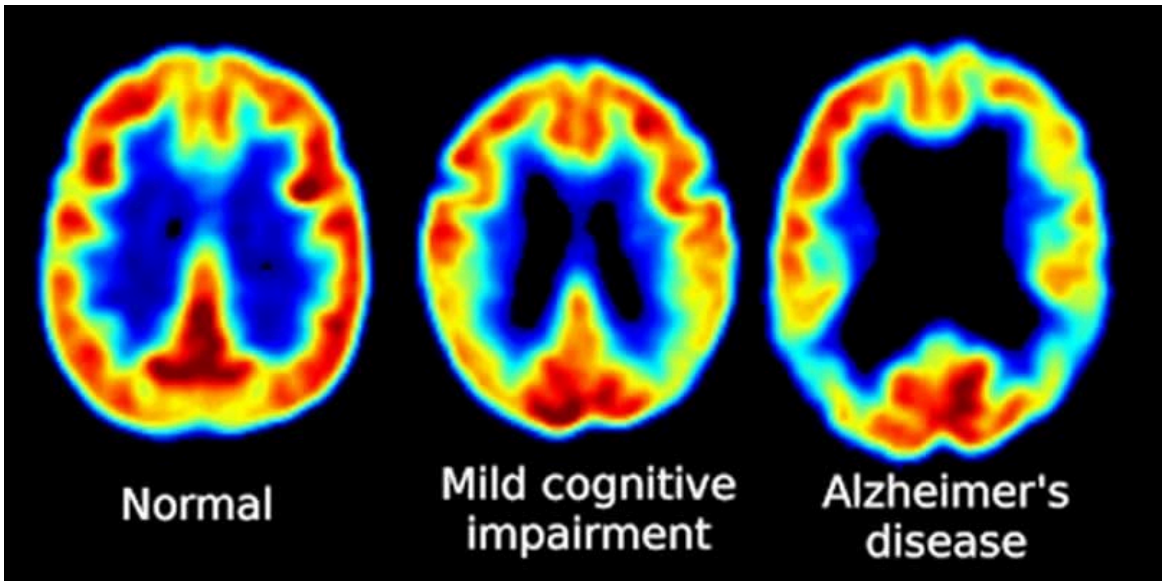


Figure 7.8.5 Mild cognitive impairment (MCI) is a state between normal ageing and dementia, where someone's mind is functioning less well than would be expected for their age. This image is for illustrative purposes only. (Public Domain; Center For Functional Imaging, Lawrence Berkeley National Laboratory. Alzheimer's Disease Neuroimaging Initiative (ADNI).

PET Scanning, is used to image the physiological aspects of the body rather than the anatomy (Figures 7.8.5 and 7.8.6). It images the function of the body rather than the form, such as where tagged molecules go and how they are used. For instance, if you were to image the brain of a deceased person, nothing would show up on a PET scan opposed to a CAT scan, as the brain is no longer functional. Pet Scanning is very useful in imaging tumors, which can be done when patients are injected with certain tracers. Often times PET scanners are used in collaboration with CAT scanners to create a composite image that shows both the function and form of the body. The animation below is a whole-body PET scan using the radioisotope of ^{18}F ($t_{1/2}$ = 110 min). Using this tracer, doctors can determine if cancer has metastasized by looking at the metabolic activity of glucose.

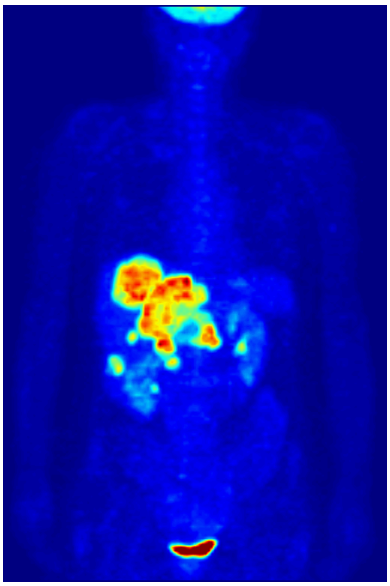


Figure 7.8.6 A whole body analysis by a PET scan. (Public Domain; Jens Maus, http://jens-maus.de/)

Other Isotopic Tests

Radioisotopes have revolutionized medical practice, where they are used extensively. Over 10 million nuclear medicine procedures and more than 100 million nuclear medicine tests are performed annually in the United States. Four typical examples of radioactive tracers used in medicine are technetium-99 ($^{99}_{43}\mathrm{Tc}$), thallium-201 ($^{201}_{81}\mathrm{Tl}$), iodine-131 ($^{131}_{53}\mathrm{I}$), and sodium-24 ($^{24}_{11}\mathrm{Na}$). Damaged tissues in the heart, liver, and lungs absorb certain compounds of technetium-99 preferentially. After it is injected, the

location of the technetium compound, and hence the damaged tissue, can be determined by detecting the γ rays emitted by the Tc-99 isotope. Thallium-201 (Figure 7.8.7) becomes concentrated in healthy heart tissue, so the two isotopes, Tc-99 and Tl-201, are used together to study heart tissue. Iodine-131 concentrates in the thyroid gland, the liver, and some parts of the brain. It can therefore be used to monitor goiter and treat thyroid conditions, such as Grave's disease, as well as liver and brain tumors. Salt solutions containing compounds of sodium-24 are injected into the bloodstream to help locate obstructions to the flow of blood.

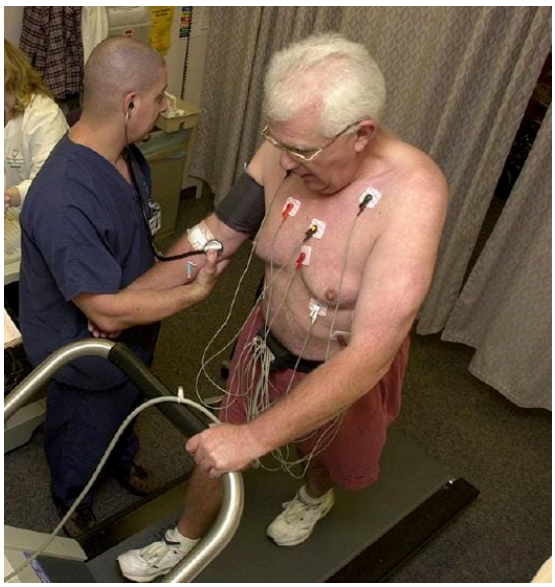


Figure 7.8.7 Administering thallium-201 to a patient and subsequently performing a stress test offer medical professionals an opportunity to visually analyze heart function and blood flow. (credit: modification of work by "BlueOctane"/Wikimedia Commons)

Small doses of I-131 (too small to kill cells) are used for purposes of imaging the thyroid. Once the iodine is concentrated in the thyroid, the patient lays down on a sheet of film and the radiation from the I-131 makes a picture of the thyroid on the film. The half-life of iodine-131 is approximately 8 days so after a few weeks, virtually all of the radioactive iodine is out of the patient's system. During that time, they are advised that they will set off radiation detectors in airports and will need to get special permission to fly on commercial flights.

Some isotopes that are used to diagnose diseases are shown in Table 7.8.1. All of these nuclear isotopes release one form of ionizing radiation (either/add particle or ray). In addition, each isotopic application would involve a specific amount of mrem/mSv radiation.

Table 7.8.1 Selected Radioisotopes Used in Diagnostic Nuclear Medicine

Symbol-mass	Half-Life ($t_{1/2}$)	Application
Xe-133	5.27 days	Lung imaging
H-3	12.26 years	Analyzing total body water
Tl-201	73 hours	Stress tests for heart problems
Fe-59	44.5 days	Detection of anemia
Gd-153	242 days	Analyzing bone density
Cr-51	27.8 days	Determining blood volume
C-11	20.4 minutes	Brain scans
Tc-99m	6.0 hours	Heart, lung, kidney, bone marrow, brain, or bone marrow imaging
Pu-238	86 years	Powering pacemakers
I-131	8.0 days	Imaging Thyroid

Symbol-mass	Half-Life ($t_{1/2}$)	Application
The “m” in Tc-99m stands for “metastable,” indicating that this is an unstable, high-energy state of Tc-99. Metastable isotopes emit γ radiation to rid themselves of excess energy and become (more) stable.		

Keep in mind that x-rays, CT scans, PET scans, and isotopic studies involve ionizing radiation. In contrast, MRI (magnetic resonance imaging) and ultrasounds do not utilize ionizing forms of radiation.

Therapeutic Radiation

There are many techniques used to treat cancer. Surgery can be used to remove cancerous tumors inside or on the body. With chemotherapy, ingested or injected chemicals are used to kill rapidly dividing cells (cancerous and noncancerous). Other cancer treatment methods include immunotherapy, stem cell replacement, hormone therapy, and targeted therapy.

Radiation therapy and Chemotherapy: Two different treatment procedures

Patients diagnosed with cancer might be required to do chemotherapy or radiation therapy. Sometimes, both of these methods are used for a patient. In this picture, a female patient is receiving chemotherapy through an IV. She is participating in cooling therapy while receiving her treatment. By placing her hands and feet in cooling devices, this will lower her chances of losing her finger and toenails. Cold cap therapy is also now available for chemotherapy patients. Wearing this type of device could enable a patient to keep his/her hair during chemotherapy.



Figure 7.8.8 *(CC BY 2.0; Jenny Mealing)*.

Current therapeutic radiation applications involve the use of gamma, x-rays, or protons. Recently, some research facilities are investigating the use of alpha and beta tagged molecules to kill cancer cells. These radioisotopes will first locate a cancer related molecule on a tumor cell. Then, the alpha or beta tagged species will inject its radiation into the tumor. Sr-89 (beta emitter) and Ra-223 (alpha emitter) have been used in clinical research trials of certain types of bone cancers.

Radiation Therapy is used as a treatment to control malignant cells within cancer patients. Oncologists (specialists that deal with cancer) utilize radiation frequently to help slow or cure the spread of cancer within individuals. Radiation is specifically applied to malignant tumors in order to shrink them in size. Medical professionals, mainly radiation oncologists, administer a variety of dosages to patient, contingent to the patients current health, as well as other treatments such as chemotherapy, success of surgery, etc.

External Beam Therapy (Photon and Proton Therapy)

External Beam Therapy (EBT) is a method of delivering a high energy beam of radiation to the precise location of a patient's tumor. These beams can destroy cancer cells and with careful planning, NOT kill surrounding cells. The concept is to have several beams of radiation, each of which is sub-lethal, enter the body from different directions. The only place in the body where the beam would be lethal is at the point where all the beams intersect. Before the EBT process, the patient is three-dimensionally mapped using CT scans and x-rays. The patient receives small tattoos to allow the therapist to line up the beams exactly. Alignment lasers are used to precisely locate the target. The radiation beam is usually generated with a linear accelerator. The video below illustrates the basic preparation and administration of external beam therapy.

Radiation Therapy

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View this YouTube Video online

Video 7.8.2 *Targeting cancer - Radiation therapy treatment process.*

Photon EBT utilizes either x-ray or gamma rays. An x-ray source would require a linear accelerator to produce high energy electrons. In contrast, a gamma source incorporates a radioactive isotope (like Co-60). Keep in mind both of these technologies use ionizing radiation. As a result, cancer patients must be monitored throughout their life to ensure they do not develop other cancers, like leukemia. EBT is used to treat the following diseases as well as others:

- Breast cancer
- Colorectal cancer
- Head and neck cancer
- Lung cancer
- Prostate cancer

The majority of radiation patients receive photon EBT. The smaller size of this machine makes this therapy an option for all sizes of hospitals and cancer treatment centers. Photon EBT equipment costs approximately three million dollars. The size and the price of this technology enables smaller facilities to keep their patients closer to home during treatment

Another method of radiation treatment involving protons is not as commonly used in the United States. **Photon therapy** requires a cyclotron to generate proton beams (recall, a proton is an ionized H-1 isotope). Unlike x-ray or gamma rays (photon therapy), protons are extremely heavy.

At this time, research facilities are working on miniaturizing proton generators. An ideal technology would reduce the cost from hundreds of millions of dollars to about twenty to thirty million per device. This would make proton therapy more available and convenient for patients.

Proton Therapy vs Radiation Therapy

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View this YouTube Video online

Video7.8.3 *Proton therapy versus radiation therapy*

There are many advantages to choosing proton therapy over photon therapy. Unlike photon radiation, proton beams will only penetrate to the depth of the tumor and not pass through the entire body. This reduces the overall toxicity dose. In addition, fewer treatments are required for proton therapy patients than photon therapy.Unfortunately, proton therapy is more expensive than photon therapy and less common. Once approved by a facility and medical insurance, a patient may have to move temporarily to a larger city to receive treatment. Some forms of cancer have never been clinically treated with proton therapy (namely, breast cancer). Patients desiring proton therapy might not be able to receive type of therapy due to lack of research.

Brachtherapy

Brachytherapy involves placing ionizing pellets(seeds) or rods directly at the tumor. Photons (in the form of x-ray or gamma ray) are produced inside the body and will penetrate throughout this particular area localizing the radiation. Pellets are surgically implanted while rods can be temporarily inserted to produce radiation internally. With pellets/seeds, the patient will remain radioactive as long as these devices remain inside the body. People undergoing this type of radiotherapy need to be aware of their constant emission of radiation. Radiation that is administered through rods connected to a photon device will disperse energy immediately and not leave the patient radioactive.

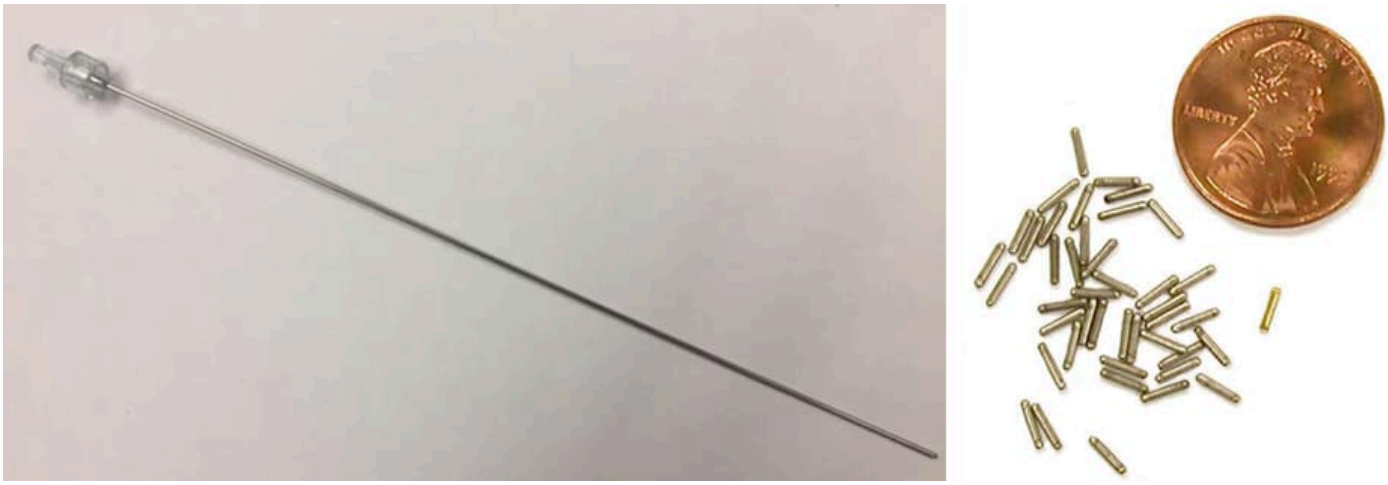


Figure 7.8.9 The picture on the left shows a brachytherapy rod. This device would be connected to a gamma source and connected inside a human body. The second images shows radioactive pellets or seeds that can be surgically implanted near a tumor. Note how small the pellets are when compared to an American penny. Image taken from: https://c1.staticflickr.com/8/7254/7...74d83b7058.jpg

Brachytherapy is widely used in the treatment of cancers involving reproductive organs. Because the radiation is isolated internally, patients are less likely to experience side effects when receiving this type of treatment. Cancers that have been treated with brachytherapy are shown below:

- Prostate
- Breast
- Esophageal
- Lung
- Uterine
- Anal/Rectal
- Sarcomas
- Head and neck

Table 7.8.2 list commonly used radionuclides for brachtherapy.

Table 7.8.2 Commonly Used Radiation Sources (radionuclides) for Brachytherapy.

Radionuclide	Type	Half-life
Cesium-131 (^{131}Cs)	Electron Capture, ε	9.7 days
Cesium-137 (^{137}Cs)	β^-- particles, γ-rays	30.17 years
Cobalt-60 (^{60}Co)	β^-- particles, γ-rays	5.26 years

Radionuclide	Type	Half-life
Iridium-192 (^{192}Ir)	γ-rays	73.8 days
Iodine-125 (^{125}I)	Electron Capture, ε	59.6 days
Palladium-103 (^{103}Pd)	Electron Capture, ε	17.0 days
Ruthenium-106 (^{106}Ru)	β^-- particles	1.02 years
Radium-226 (^{226}Ra)	β^-- particles	1599 years

Side Effects of Radiation Therapy

Patients receiving radiation therapy can experience a variety of side effects. For example, sterility could occur if reproductive organs are irradiated. Skin that has been irradiated can appear dry and feel itchy. Some patients will loose sensation in the irradiated area. Radiation can affect the production of white and red blood cells. A reduction of white blood cells results in immunity disorders. Red blood cell lose causes anemia. Gastrointestinal issues such as diarrhea and nausea are common during radiation therapy. Some patients will lose hair as well. Lastly, dry mouth and tooth decay are prevalent during radiation treatments.

Medications are available to alleviate symptoms of radiation therapy. Narcotics can be prescribed to help alleviate intense pain. Prescription medications like zofran (Figure 7.8.10) and phenergan can help with nausea. Special mouthwashes have been formulated to reduce dry mouth and cavities.



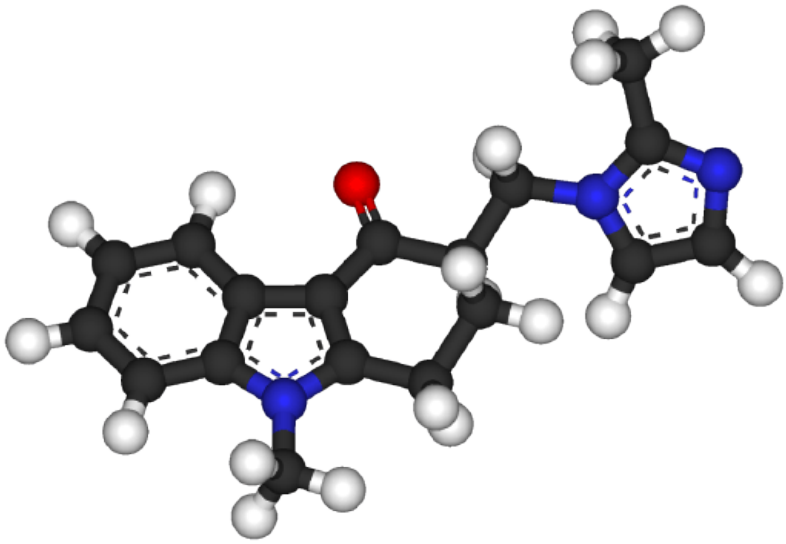


Figure 7.8.10(left) *Zofran is a medication used to prevent nausea and vomiting caused by cancer chemotherapy, radiation therapy, or surgery. (right) A vial of Zofran 4 mg containing ondansetron for intravenous injection. This medicine is a controlled substance and is used for nausea. Images used with permision (left: Public domain; Fvasconcellos and right: CC BY 3.0; Intropin)*

Hair loss is a side effect of radiation, but only locally

Radiation therapy can cause hair loss, but hair is only lost in the area being treated. For instance, radiation to your head may cause you to lose some or all the hair on your head (even eyebrows and lashes), but if you get treatment to your hip, you won't lose the hair on your head.

Summary

- Compounds known as radioactive tracers can be used to follow reactions, track the distribution of a substance, diagnose and treat medical conditions, and much more.
- Other radioactive substances are helpful for controlling pests, visualizing structures, providing fire warnings, and for many other applications.
- Hundreds of millions of nuclear medicine tests and procedures, using a wide variety of radioisotopes with relatively short half-lives, are performed every year in the US.

- Radiation therapy uses high-energy radiation to kill cancer cells by damaging their DNA. The radiation used for this treatment may be delivered externally or internally.

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-
- ◦ Anonymous
- Marisa Alviar-Agnew (Sacramento City College)

- Wikipedia

Learning Objective

- Learn some applications of radioactivity.

Radioactive isotopes have a variety of applications. Generally, however, they are useful because either we can detect their radioactivity or we can use the energy they release.

Radioactive isotopes are effective tracers because their radioactivity is easy to detect. A **tracer** is a substance that can be used to follow the pathway of that substance through some structure. For instance, leaks in underground water pipes can be discovered by running some tritium-containing water through the pipes and then using a Geiger counter to locate any radioactive tritium subsequently present in the ground around the pipes. (Recall that tritium is a radioactive isotope of hydrogen.)

Tracers can also be used to follow the steps of a complex chemical reaction. After incorporating radioactive atoms into reactant molecules, scientists can track where the atoms go by following their radioactivity. One excellent example of this is the use of carbon-14 to determine the steps involved in photosynthesis in plants. We know these steps because researchers followed the progress of carbon-14 throughout the process.

Radioactive Dating

Radioactive isotopes are useful for establishing the ages of various objects. The half-life of radioactive isotopes is unaffected by any environmental factors, so the isotope acts like an internal clock. For example, if a rock is analyzed and is found to contain a certain amount of uranium-235 and a certain amount of its daughter isotope, we can conclude that a certain fraction of the original uranium-235 has radioactively decayed. If half of the uranium has decayed, then the rock has an age of one half-life of uranium-235, or about $4.5 \times 10^9 y$. Many analyses like this, using a wide variety of isotopes, have indicated that age of the earth itself is over $4 \times 10^9 y$.

In another interesting example of radioactive dating, hydrogen-3 dating has been used to verify the stated vintages of some old fine wines.

One isotope, carbon-14, is particularly useful in determining the age of once-living artifacts. A tiny amount of carbon-14 is produced naturally in the upper reaches of the atmosphere, and living things incorporate some of it into their tissues, building up to a constant, albeit very low, level. Once a living thing dies, it no longer acquires carbon-14; as time passes the carbon-14 that was in the tissues decays. (The half-life of carbon-14 is 5,370 y.) If a once-living artifact is discovered and analyzed many years after its death and the remaining carbon-14 is compared to the known constant level, an approximate age of the artifact can be determined. Using such methods, scientists determined that the age of the Shroud of Turin (Figure $7.8.1$ purported by some to be the burial cloth of Jesus Christ and composed of flax fibers, a type of plant) is about 600–700 y, not 2,000 y as claimed by some.

Scientists were also able to use radiocarbon dating to show that the age of a mummified body found in the ice of the Alps was 5,300 y.

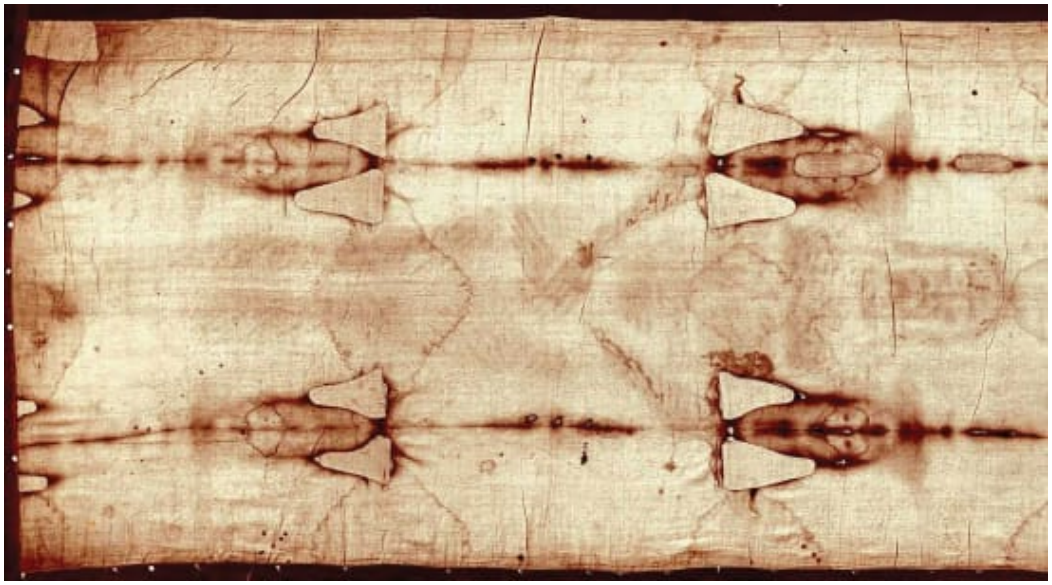


Figure 7.8.1: Shroud of Turin. In 1989, several groups of scientists used carbon-14 dating to demonstrate that the Shroud of Turin was only 600–700 y. Many people still cling to a different notion, despite the scientific evidence.

Irradiation of Food

The radiation emitted by some radioactive substances can be used to kill microorganisms on a variety of foodstuffs, extending the shelf life of these products. Produce such as tomatoes, mushrooms, sprouts, and berries are irradiated with the emissions from coba . This exposure kills a lot of the bacteria that cause spoilage, so the produce stays fresh longer. Eggs and some meat, such as beef, pork, and poultry, can also be irradiated. Contrary to the belief of some people, irradiation of food *does not* make the food itself radioactive.

Medical Applications

Radioactive isotopes have numerous medical applications in diagnosing and treating illness and diseases. One example of a diagnostic application is using radioactive iodine-131 to test for thyroid activity (Figure 7.8.1). The thyroid gland in the neck is one of the few places in the body with a significant concentration of iodine. To evaluate thyroid activity, a measured dose of $^{131}\mathrm{I}$ is administered to a patient, and the next day a scanner is used to measure the amount of radioactivity in the thyroid gland. The amount of radioactive iodine that collects there is directly related to the activity of the thyroid, allowing trained physicians to diagnose both hyperthyroidism and hypothyroidism. Iodine-131 has a half-life of only 8 d, so the potential for damage due to exposure is minimal. Technetium-99 can also be used to test thyroid function. Bones, the heart, the brain, the liver, the lungs, and many other organs can be imaged in similar ways by using the appropriate radioactive isotope.

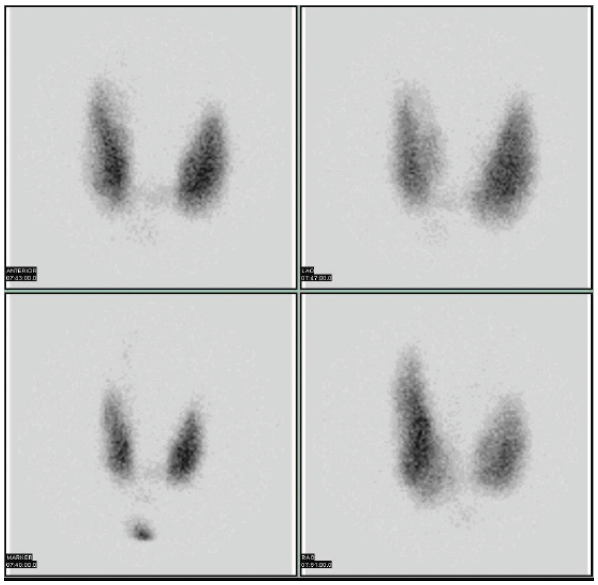


Figure 7.8.2: Medical Diagnostics. Radioactive iodine can be used to image the thyroid gland for diagnostic purposes. Source: Scan courtesy of Myo Han, en.Wikipedia.org/wiki/File:Thyroid_scan.jpg.

Very little radioactive material is needed in these diagnostic techniques because the radiation emitted is so easy to detect. However, therapeutic applications usually require much larger doses because their purpose is to preferentially kill diseased tissues. For example, if a thyroid tumor were detected, a much larger infusion (thousands of rem, as opposed to a diagnostic dose of less than 40 rem) of iodine-131 could help destroy the tumor cells. Similarly, radioactive strontium is used to not only detect, but also to ease the pain of bone cancers. Table 7.8.1 lists several radioactive isotopes and their medical uses.

Table 7.8.1: Some Radioactive Isotopes with Medical Applications

Isotope	Use
$^{32}_{P}$	cancer detection and treatment, especially in eyes and skin
$^{59}_{Fe}$	anemia diagnosis
$^{60}_{Co}$	gamma ray irradiation of tumors
$^{99m}_{Tc*}$	brain, thyroid, liver, bone marrow, lung, heart, and intestinal scanning; blood volume determination
$^{131}_{I}$	diagnosis and treatment of thyroid function
$^{133}_{Xe}$	lung imaging
$^{198}_{Au}$	liver disease diagnosis
***The "m" means that it is a metastable form of this isotope of technetium.**	

In addition to the direct application of radioactive isotopes to diseased tissue, the gamma ray emissions of some isotopes can be directed toward the tissue to be destroyed. Cobalt-60 is a useful isotope for this kind of procedure.

Food and Drink Application: Radioactivity in Wines

Wine lovers put some stock in *vintages*, or the years in which the wine grapes were grown before they were turned into wine. Wine can differ in quality depending on the vintage. Some wine lovers willingly pay much more for a bottle of wine with a certain vintage. But how does one verify that a bottle of wine was in fact part of a certain vintage? Is the label a fake? Is that stash of wine found in the basement of a French chateau really from the 1940s, or was it made in 2009?



Figure 7.8.3: A wine label. This wine label from a bottle of wine claims a vintage of 1991. Is the wine really from this vintage, or is it a fake? Radioactivity can help determine the answer. Source: Used by permission of Ralph E. Wermuth.

Cesium-137 is a radioactive isotope that has a half-life of 30.1 y. It was introduced into the atmosphere in the 1940s and 1950s by the atmospheric testing of nuclear weapons by several countries after World War II. A significant amount of cesium-137 was released during the Chernobyl nuclear disaster in 1986. As a result of this atmospheric contamination, scientists have precise measurements of the amount of cesium-137 available in the environment since 1950. Some of the isotope of cesium is taken up by living plants, including grape vines. Using known vintages, oenologists (wine scientists) can construct a detailed analysis of the cesium-137 of various wines through the years.

The verification of a wine's vintage requires the measurement of the activity of cesium-137 in the wine. By measuring the current activity of cesium-137 in a sample of wine (the gamma rays from the radioactive decay pass through glass wine bottles easily, so there's no need to open the bottle), comparing it to the known amount of cesium-137 from the vintage, and taking into account the passage of time, researchers can collect evidence for or against a claimed wine vintage.

Before about 1950, the amount of cesium-137 in the environment was negligible, so if a wine dated before 1950 shows any measurable cesium-137 activity, it is almost surely a fake, so do not shell out lots of money for it! It may be a good wine, but it is almost definitely not over 60 years old.

Summary

Radioactivity has several practical applications, including tracers, medical applications, dating once-living objects, and preservation of food.

Learning Objective

- Compare qualitatively the ionizing and penetration power of alpha particles (α), beta particles (β), and gamma rays (γ).

With all the radiation from natural and man-made sources, we should quite reasonably be concerned about how all the radiation might affect our health. The damage to living systems is done by radioactive emissions when the particles or rays strike tissue, cells, or molecules and alter them. These interactions can alter molecular structure and function; cells no longer carry out their proper function and molecules, such as DNA, no longer carry the appropriate information. Large amounts of radiation are very dangerous, even deadly. In most cases, radiation will damage a single (or very small number) of cells by breaking the cell wall or otherwise preventing a cell from reproducing.

The ability of radiation to damage molecules is analyzed in terms of what is called **ionizing power**. When a radiation particle

interacts with atoms, the interaction can cause the atom to lose electrons and thus become ionized. The greater the likelihood that damage will occur by an interaction is the ionizing power of the radiation.

Much of the threat from radiation is involved with the ease or difficulty of protecting oneself from the particles. How thick of wall do you need to hide behind to be safe? The ability of each type of radiation to pass through matter is expressed in terms of **penetration power.** The more material the radiation can pass through, the greater the penetration power and the more dangerous they are. In general, the greater mass present the greater the ionizing power and the lower the penetration power.

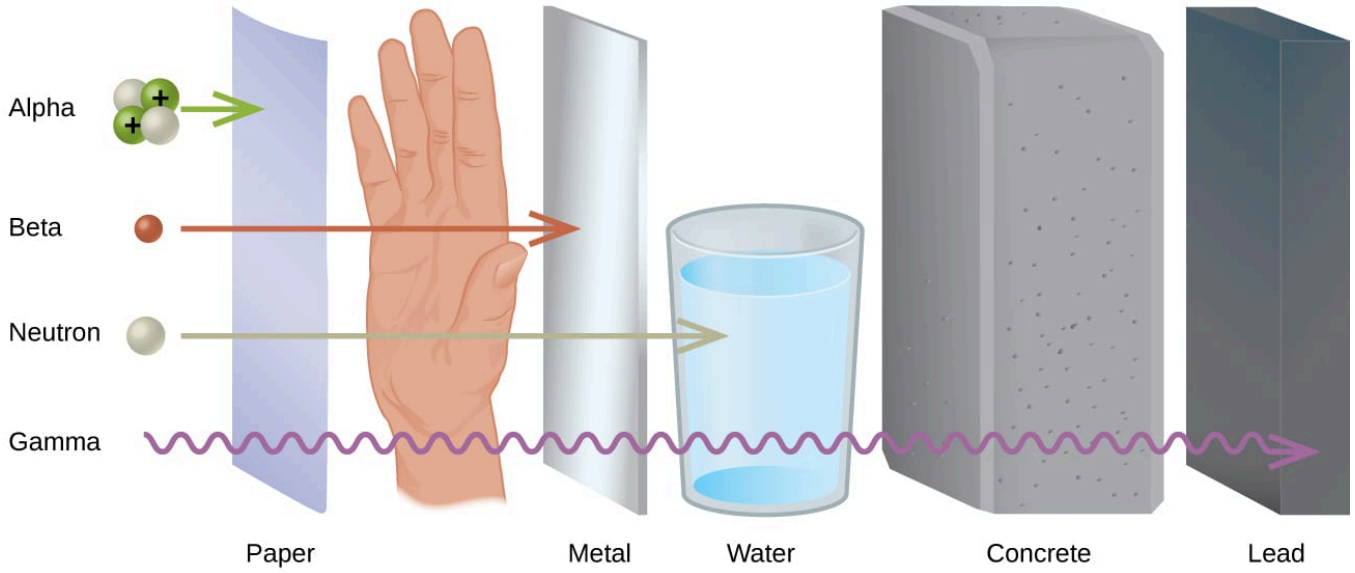


Figure 7.8.1: The ability of different types of radiation to pass through material is shown. From least to most penetrating, they are alpha < beta < neutron < gamma. (CC BY-SA 4.0; OpenStax).

Comparing only the three common types of ionizing radiation, **alpha particles** have the greatest mass. Because of the large mass of the alpha particle, it has the highest ionizing power and the greatest ability to damage tissue. That same large size of alpha particles, however, makes them less able to penetrate matter. Alpha particles have the least penetration power and can be stopped by a thick sheet of paper or even a layer of clothes. They are also stopped by the outer layer of dead skin on people. This may seem to remove the threat from alpha particles but only from external sources. In a situation like a nuclear explosion or some sort of nuclear accident where radioactive emitters are spread around in the environment, the emitters can be inhaled or taken in with food or water and once the alpha emitter is inside you, you have no protection at all.

Beta particles are much smaller than alpha particles and therefore, have much less ionizing power (less ability to damage tissue), but their small size gives them much greater penetration power. Most resources say that beta particles can be stopped by a one-quarter inch thick sheet of aluminum. Once again, however, the greatest danger occurs when the beta emitting source gets inside of you.

Gamma rays are not particles but a high energy form of electromagnetic radiation (like x-rays except more powerful). Gamma rays are energy that has no mass or charge. Gamma rays have tremendous penetration power and require several inches of dense material (like lead) to shield them. Gamma rays may pass all the way through a human body without striking anything. They are considered to have the least ionizing power and the greatest penetration power.

A comparison of Alpha particles, beta particles and gamma rays is given in Table 7.8.1.

Table 7.8.1 Comparison of Penetrating Power,Ionizing Power and Shielding of Alpha and Beta Particles, and Gamma Rays.

Particle	Symbol	Mass	Penetrating Power	Ionizing Power	Shielding
Alpha	α	4amu	Very Low	Very High	Paper Skin
Beta	β	$1/2000\text{amu}$	Intermediate	Intermediate	Aluminum
Gamma	γ	0 (energy only)	Very High	Very Low	2 inches lead

The safest amount of radiation to the human body is zero. It isn't possible to be exposed to no ionizing radiation so the next best

goal is to be exposed to as little as possible. The two best ways to minimize exposure is to limit time of exposure and to increase distance from the source.

Summary

- Types of radiation differ in their ability to penetrate material and damage tissue, with alpha particles the least penetrating but potentially most damaging and gamma rays the most penetrating.
- The two best ways to minimize exposure is to limit time of exposure and to increase distance from the source.

Learning Objectives

- Explain where nuclear energy comes from.
- Describe the difference between fission and fusion.
- Know key examples of nuclear fission and nuclear fusion.

A combination of radiochemistry and radiation chemistry is used to study nuclear reactions such as fission and fusion. Some early evidence for nuclear fission was the formation of a short-lived radioisotope of barium which was isolated from neutron irradiated uranium (^{139}Ba, with a half-life of 83 minutes and ^{140}Ba, with a half-life of 12.8 days, are major fission products of uranium). **Nuclear fission** is the splitting of an atomic nucleus. In nuclear weapons and reactors, neutrons hit unstable nuclei to form smaller atoms. **Nuclear Fusion** is the bringing together of two atomic nuclei to form a larger atom. Figure $7.8.1$ illustrates the difference between nuclear fission and nuclear fusion.

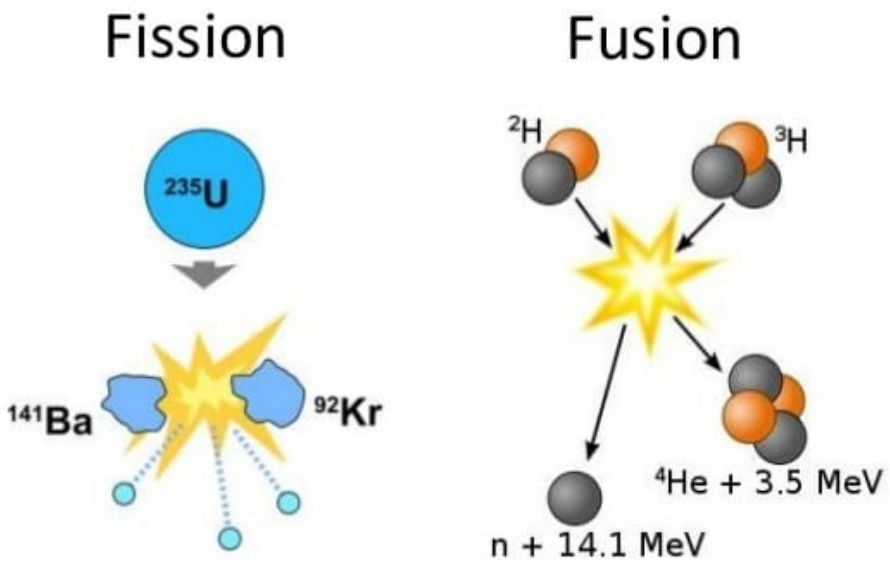


Figure $7.8.1$ Nuclear fission occurs with one large nuclear is split into two or more smaller nuclei (left) . Nuclear fusion happens when two small nuclei combine to make a larger nucleus (right) .

Einstein and the Equivalence of Mass and Energy

Nuclear changes occur with a simultaneous release of energy. Where does this energy come from? If we could precisely measure the masses of the reactants and products of a nuclear reaction, we would notice that the amount of mass drops slightly in the conversion from reactants to products. Consider the following nuclear equation, in which the molar mass of each species is indicated to four decimal places:

$$^{235}_{235.0439}\mathrm{U} \rightarrow ^{139}_{138.9088}\ \mathrm{Ba} + ^{94}_{93.9343}\ \mathrm{Kr} + ^{2^1}_{2\times 1.0087}\ \mathrm{n}$$

If we compare the mass of the reactant (235.0439) to the masses of the products (sum=234.8605) we notice a mass difference of -0.1834 g or -0.0001834 kg. Where did this mass go?

According to Albert Einstein's theory of relativity, energy (*E*) and mass (*m*) are related by the following equation:

$$E = mc^2$$

where *c* is the speed of light, or

$$c = 3.00 \times 10^8\, m/s$$

In the course of the chemical reaction for uranium, the mass difference is converted to energy, which is given off by the reaction:

$$E = (-0.0001834\, kg)(3.00 \times 10^8\, m/s)^2 = -1.65 \times 10^{13} J = -1.65 \times 10^{10} kJ$$

(For the units to work out, mass must be expressed in units of kilograms.) That is, 16.5 billion kJ of energy is given off every time 1 mol of uranium-235 undergoes this nuclear reaction. This is an extraordinary amount of energy. Compare it to combustion reactions of hydrocarbons, which give off about 650kJ/mol of energy for every CH_2 unit in the hydrocarbon-on the order of *hundreds* of kilojoules per mole. Nuclear reactions give off *billions* of kilojoules per mole.

If this energy could be properly harvested, it would be a significant source of energy for our society. **Nuclear energy** involves the controlled harvesting of energy from fission reactions. The reaction can be controlled because the fission of uranium-235 (and a few other isotopes, such as plutonium-239) can be artificially initiated by injecting a neutron into a uranium nucleus. The overall nuclear equation, with energy included as a product, is then as follows:

$$^{235}\mathrm{U} + {}^{1}\mathrm{n} \rightarrow {}^{139}\mathrm{Ba} + {}^{94}\mathrm{Kr} + 3\,{}^{1}\mathrm{n}$$

Thus by the careful addition of extra neutrons into a sample of uranium, we can control the fission process and obtain energy that can be used for other purposes. (Artificial or induced radioactivity, in which neutrons are injected into a sample of matter that subsequently cause fission, was first demonstrated in 1934 by Irène Joliot-Curie and Frédéric Joliot, the daughter and son-in-law of Marie Curie.)

Binding Energy

The forces that bind nucleons together in an atomic nucleus are much greater than those that bind an electron to an atom through electrostatic attraction. This is evident by the relative sizes of the atomic nucleus and the atom (10^{-15} and 10^{-10}m, respectively). The energy required to pry a nucleon from the nucleus is therefore much larger than that required to remove (or ionize) an electron in an atom. In general, all nuclear changes involve large amounts of energy per particle undergoing the reaction. This has numerous practical applications.

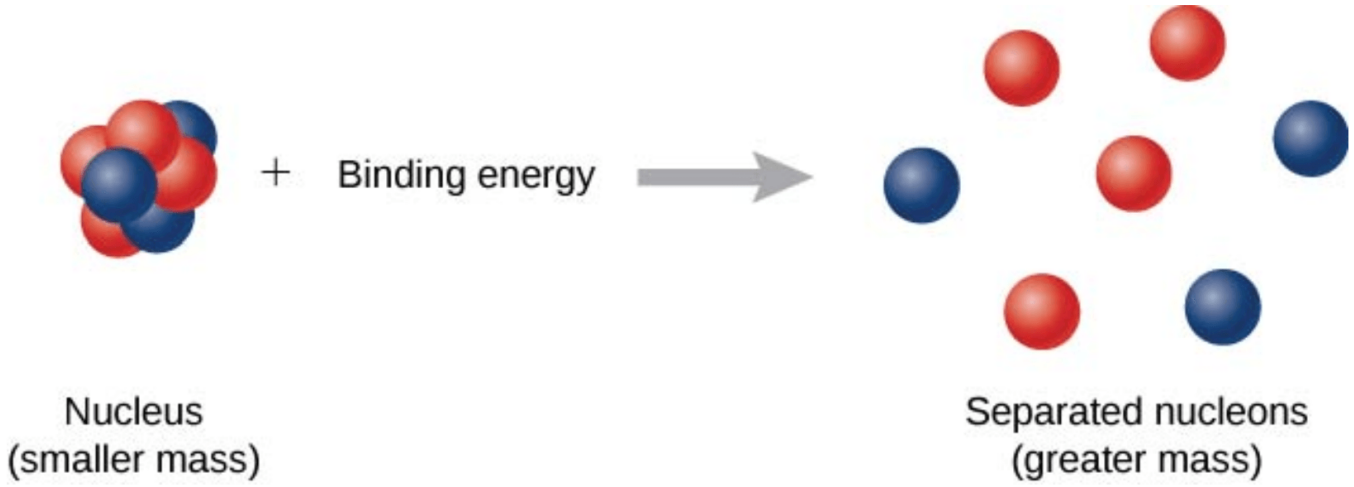


Figure 7.8.2*The binding energy is the energy required to break a nucleus into its constituent protons and neutrons. A system of separated nucleons has a greater mass than a system of bound nucleons.*

As shown in figure 7.8.2, energy is put into the system to break apart the nucleus. The amount of energy required is called the total **binding energy** (BE), E_b. The binding energy is equal to the amount of energy released in forming the nucleus, and is therefore given by

$$E_b = (\Delta m)c^2. \tag{7.8.1}$$

In nuclear physics, one of the most important experimental quantities is the *binding energy per nucleon (BEN)*, which is defined by

$$BEN = \frac{E_b}{A} \tag{7.8.2}$$

This quantity is the average energy required to remove an individual nucleon (proton or neutron) from a nucleus—analogous to the ionization energy of an electron in an atom. If the BEN is relatively large, the nucleus is relatively stable. BEN values are estimated from nuclear scattering experiments.

Figure 7.8.3 shows the relative binding energies for various isotopes. Of these elements, fission requires heavy, unstable nuclei. This means selected atoms would have low binding energies and would have large atomic masses. Nuclei that are larger than Fe-56 may undergo fission. Uranium-238 and uranium-235 both have lower binding energies with heavy masses-.

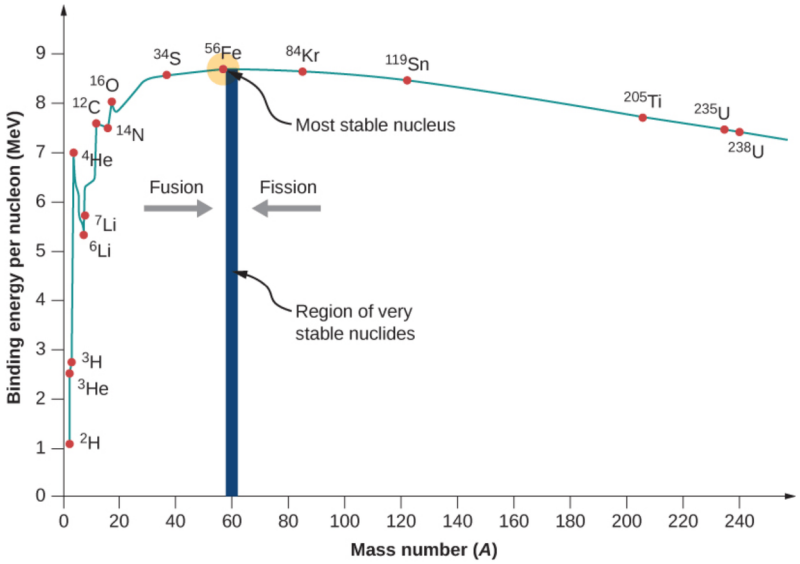


Figure 7.8.3 In this graph of binding energy per nucleon for stable nuclei, the BEN is greatest for nuclei with a mass near ^{56}Fe. Therefore, fusion of nuclei with mass numbers much less than that of Fe, and fission of nuclei with mass numbers greater than that of Fe, are exothermic processes. (CC BY. OpenStax).

These two isotopes would be suitable for splitting based on these requirements. Of these two isotopes, only uranium-235 is readily fissionable. This is due to odd neutron count contributing to additional instability. Another fissionable isotope not shown in Figure 7.8.3 is plutonium-239. This is a synthetic isotope produced by transmutation and decay reactions. Like uranium-235, it has low binding energy, high mass, and an odd number of neutrons. U-235 and Pu-239 are used in atomic bombs (fission based) and nuclear reactors.

Positively charged centers of atoms make nuclear fusion extremely difficult. For this reason, smaller atoms are suitable for fusion reactions. In addition, these particular isotopes need to have low binding energies in order to undergo fusion. Atoms that possess these two qualities are H-2 and H-3. For fusion to occur, extreme temperatures are required to fuse these deuterium and tritium together.

As we will see, the BEN-versus-*A* graph implies that nuclei divided or combined release an enormous amount of energy. This is the basis for a wide range of phenomena, from the production of electricity at a nuclear power plant to sunlight.

Nuclear Fission

In both fission and fusion, large amounts of energy are given off in the form of heat, light, and gamma radiation. Italian physicist, Enrico Fermi, performed the first fission reaction in 1934. He was unaware that he had split a uranium atom into two smaller nuclei. Nuclear fission of heavy elements was discovered on December 17, 1938 by German Otto Hahn and his assistant Fritz Strassmann, and explained theoretically in January 1939 by Lise Meitner (Figure 7.8.4) and her nephew Otto Robert Frisch. Frisch

named the process "fission" by analogy with biological fission of living cells.



Figure 7.8.4 *Austrian physicist, Lise Meitner of U-238.*

Image take from: https://upload.wikimedia.org/wikiped...ner12crop2.J

Since then, fission has been observed in many other isotopes, including most actinide isotopes that have an odd number of neutrons. A typical nuclear fission reaction is shown in Figure 7.8.5.

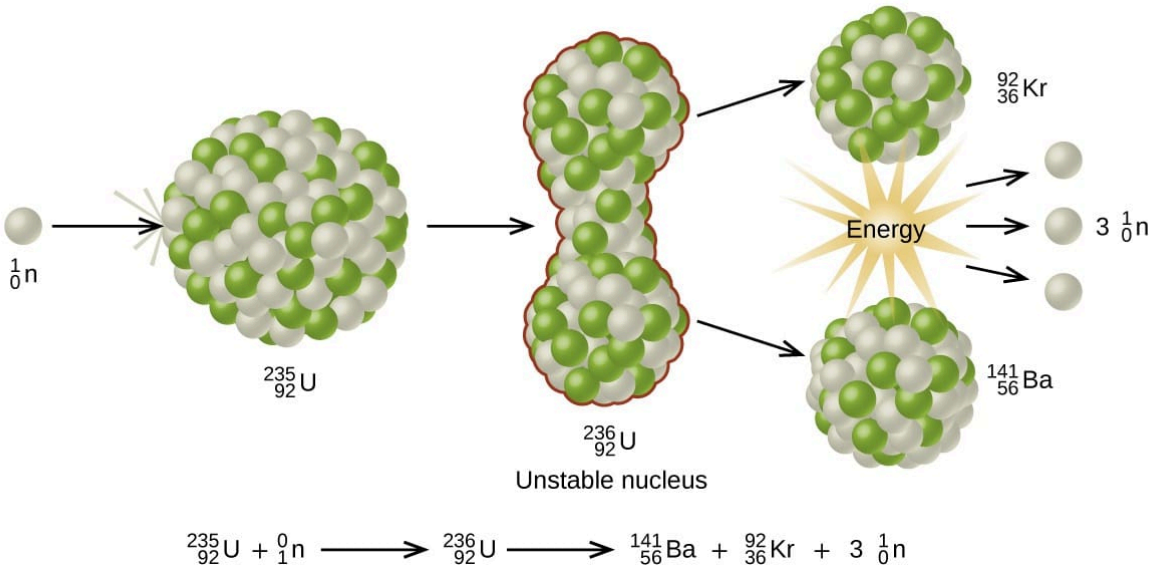


Figure 7.8.5 When a slow neutron hits a fissionable U-235 nucleus, it is absorbed and forms an unstable U-236 nucleus. The U-236 nucleus then rapidly breaks apart into two smaller nuclei (in this case, Ba-141 and Kr-92) along with several neutrons (usually two or three), and releases a very large amount of energy.

Among the products of Meitner, Hahn, and Strassman's fission reaction were barium, krypton, lanthanum, and cerium, all of which have nuclei that are more stable than uranium-235. Since then, hundreds of different isotopes have been observed among the products of fissionable substances.

Fission has been used in nuclear weapons and powers all nuclear reactors. Approximately fifty-five countries worldwide possess fission technology in the form of research or energy reactors. Less than ten of these countries have fission weapons.

A tremendous amount of energy is produced by the fission of heavy elements. For instance, when one mole of U-235 undergoes fission, the products weigh about 0.2 grams less than the reactants; this "lost" mass is converted into a very large amount of energy, about 1.8×10^{10} kJ per mole of U-235. Nuclear fission reactions produce incredibly large amounts of energy compared to chemical reactions. The fission of 1 kilogram of uranium-235, for example, produces about 2.5 million times as much energy as is produced by burning 1 kilogram of coal.

Nuclear Chain Reaction

As described earlier, when undergoing fission U-235 produces two “medium-sized” nuclei, and two or three neutrons. These neutrons may then cause the fission of other uranium-235 atoms, which in turn provide more neutrons that can cause fission of even more nuclei, and so on. If this occurs, we have a nuclear chain reaction (Figure 7.8.6). On the other hand, if too many neutrons escape the bulk material without interacting with a nucleus, then no chain reaction will occur.

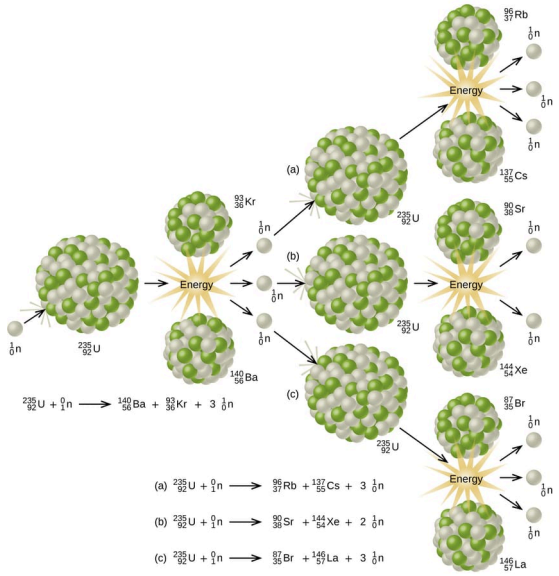


Figure 7.8.6 The fission of a large nucleus, such as U-235, produces two or three neutrons, each of which is capable of causing fission of another nucleus by the reactions shown. If this process continues, a nuclear chain reaction occurs.

Material that can sustain a nuclear fission chain reaction is said to be fissile or fissionable. (Technically, fissile material can undergo fission with neutrons of any energy, whereas fissionable material requires high-energy neutrons.) Nuclear fission becomes self-sustaining when the number of neutrons produced by fission equals or exceeds the number of neutrons absorbed by splitting nuclei plus the number that escape into the surroundings. The amount of a fissionable material that will support a self-sustaining chain reaction is a critical mass. An amount of fissionable material that cannot sustain a chain reaction is a subcritical mass. An amount of material in which there is an increasing rate of fission is known as a supercritical mass. The critical mass depends on the type of material: its purity, the temperature, the shape of the sample, and how the neutron reactions are controlled (Figure 7.8.7).

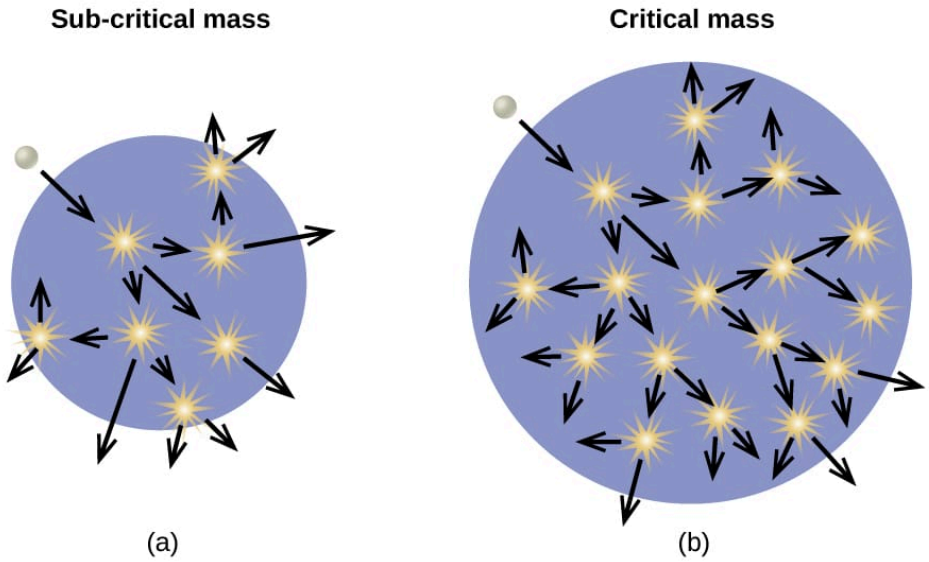


Figure 7.8.7 (a) In a subcritical mass, the fissile material is too small and allows too many neutrons to escape the material, so a chain reaction does not occur. (b) In a critical mass, a large enough number of neutrons in the fissile material induce fission to create a chain reaction.

Thermonuclear Reactions

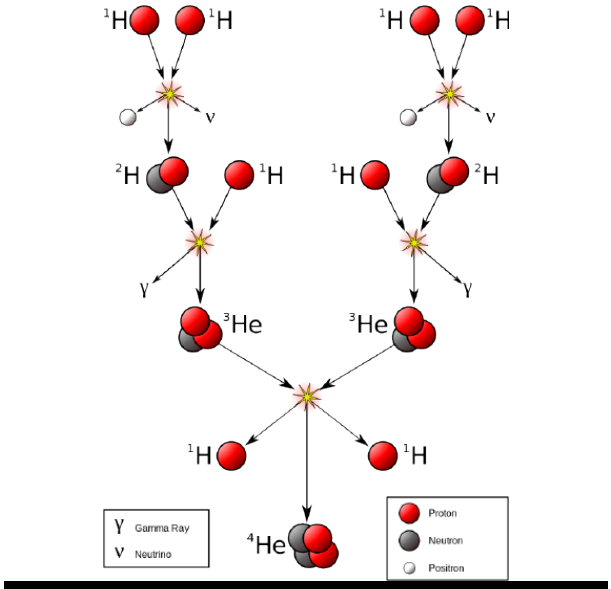
Thermonuclear fusion is a way to achieve nuclear fusion by using extremely high temperatures. There are two forms of thermonuclear fusion: *uncontrolled*, in which the resulting energy is released in an uncontrolled manner, as it is in thermonuclear weapons ("hydrogen bombs") and in most stars; and *controlled*, where the fusion reactions take place in an environment allowing some or all of the energy released to be harnessed for constructive purposes. The discussion below is limited to the fusion process that powers the sun and the stars.

The prime energy producer in the sun is the fusion of hydrogen to form helium, which occurs at a solar-core temperature of 14 million kelvin. The net result is the fusion of four protons into one alpha particle, with the release of two positrons, two neutrinos (which changes two of the protons into neutrons), and energy (Figure 7.8.8).

This is a net reaction of a more complicated series of events:

$$4\,{}^{1}_{1}\mathrm{H} \longrightarrow {}^{4}_{2}\mathrm{He} + 2\,{}^{1}_{0}\mathrm{n}$$

A helium nucleus has a mass that is 0.7% less than that of four hydrogen nuclei; this lost mass is converted into energy during the fusion. This reaction produces about 3.6×10^{11} kJ of energy per mole of ${}^{4}_{2}\mathrm{He}$ produced. This is somewhat larger than the energy produced by the nuclear fission of one mole of U-235 (1.8×10^{10} kJ), and over 3 million times larger than the energy produced by the (chemical) combustion of one mole of octane (5471 kJ).



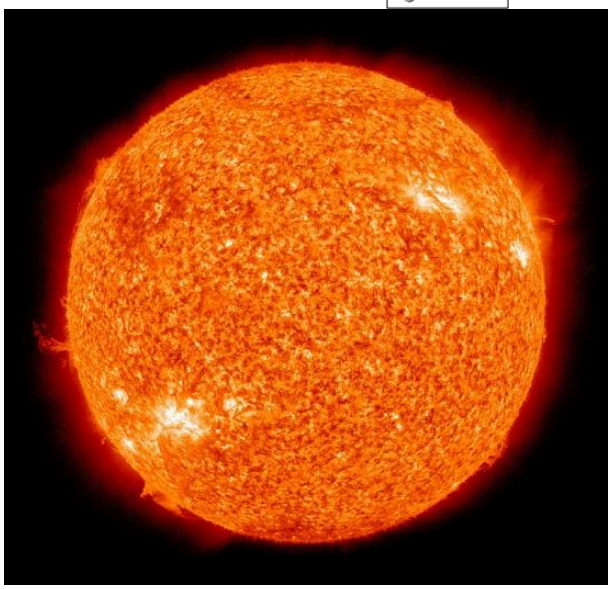


Figure 7.8.8 Nuclear fusion producing helium (left). The Sun is a main-sequence star, and thus generates its energy by nuclear fusion of hydrogen nuclei into helium (right).

Summary

- Nuclear energy comes from tiny mass changes in nuclei as radioactive processes occur.
- In fission, large nuclei break apart and release energy; in fusion, small nuclei merge together and release energy.
- The continues process whereby neutrons produced from the initial fission of a large nucleus like U-235 cause further fission of another nucleus.
- In a critical mass, a large enough number of neutrons in the fissile material induce fission to create a chain reaction.
- The most important fusion process in nature is the one that powers the stars. The net result is the fusion of four protons into one alpha particle, with the release of two positrons, two neutrinos (which changes two of the protons into neutrons), and energy

Contributors and Attributions

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- TextMap: Beginning Chemistry (Ball et al.)
- Marisa Alviar-Agnew (Sacramento City College)
- Wikipedia

Learning Objectives

- Describe the key features of an atomic bomb.
- Define nuclear fallout.

World war two brought about many advances in weaponry. The Nazis used the immense science ability of their newly conquered countries to attempt to create super weapons to keep the new territory. The weapons included new tanks like the feared Tiger Tank, or the highly advanced Luftwaffe air force. However, the one weapon the Nazi's really wanted to hold was the atomic bomb. The problem that the Nazi's began to have though was that many of the great minds they once had left the country and came to the United States. Also, in the end Allied troops got to Hitler and his scientists just before they could complete any atomic bombs. The U.S. also wanted to create an atom bomb, and for the same reasons. The Manhattan Project was started to form this new weapon.

The Manhattan Project was successful in the United States because of the vast resources that could be pulled together for the project. The great minds worked under the University of Chicago football field to try and create fission. A plant in Tennessee had the sole purpose of producing enough uranium for atomic bomb testing. Finally, New Mexico was the epicenter of the whole operation. This is where the bombs were constructed and tested. The resources needed were immense. There had to be enough fuel for the bomb, and everything required to make it. There had to be housing and food for all the scientists. Maybe most importantly, there had to be enough money to support the project. The Unites States was able to effectively use all of these resources, while the Germans did not have the resources, and didn't use them as effectively. Finally, after years of hard work an atomic bomb test took place of July 16, 1945. The bomb was held together in some parts by just tape, yet it was a huge success. The problem was that the war in Europe was already over, so President Truman had to decide whether or not to use the bomb on Japan.

The war in the Pacific was a lot different than the war in Europe. In Europe, you landed once and fought your way through to Berlin to defeat the Nazi's. However, in the Pacific Japan controlled hundreds of Islands before you got to the mainland. The invasion would have cost huge amounts of American lives and resources. As a result, when the bomb had a successful test President Truman said he hardly hesitated to approve of the usage of the atomic bomb on Japan.

The first atomic bomb was dropped on Hiroshima on August 6, 1945 by the Enola Gay. The bomb destroyed the city, killing almost 100,000 people. The city of Hiroshima was essentially leveled. Following this attack, on August 9, 1945 another atomic bomb was dropped on Nagasaki. The same effects were felt in this city. The United States was ready to drop another atomic bomb in the third week of August, 3 more in September, and 3 more in October if necessary. However, on August 15 Japan surrendered to the allies and World War Two was officially over.



Figure 7.8.1 (left) Hiroshima before the atomic bomb (right) Hiroshima after the atomic bomb

The Workings of the Atomic Bomb

A crucial feature of the fission of uranium without which an atom bomb is impossible is that fission *produces more neutrons than it consumes*. As can be seen from Eqs. (1), for every neutron captured by a $^{235}_{92}\mathrm{U}$ nucleus, between two and four neutrons are produced. Suppose now that we have a very large sample of the pure $^{235}_{92}\mathrm{U}$ isotope and a stray neutron enters this sample. As soon as it hits a $^{235}_{92}\mathrm{U}$ nucleus, fission will take place and about three neutrons will be produced. These in turn will fission three more $^{235}\mathrm{U}$ nuclei, producing a total of nine neutrons. A third repetition will produce 27 neutrons. a fourth 81. and so on. This process (which is called a *chain reaction*) escalates very rapidly. Within a few microseconds a very large number of nuclei fission, with the release of a tremendous amount of energy, and an atomic explosion results.

$$
{}^{235}_{92}\mathrm{U} + {}^{1}_{0}n \longrightarrow \left[{}^{236}_{92}\mathrm{U}\right] \begin{cases} \longrightarrow {}^{140}_{56}\mathrm{Ba} + {}^{93}_{36}\mathrm{Kr} + 3{}^{1}_{0}n \\ \longrightarrow {}^{144}_{55}\mathrm{Cs} + {}^{90}_{37}\mathrm{Rb} + 2{}^{1}_{0}n \\ \longrightarrow {}^{144}_{54}\mathrm{Xe} + {}^{90}_{38}\mathrm{Sr} + 2{}^{1}_{0}n \\ \longrightarrow {}^{146}_{57}\mathrm{La} + {}^{87}_{35}\mathrm{Br} + 3{}^{1}_{0}n \\ \longrightarrow {}^{160}_{62}\mathrm{Sm} + {}^{72}_{30}\mathrm{Zn} + 4{}^{1}_{0}n \end{cases}
$$

There are two reasons why a normal sample of uranium metal does not spontaneously explode in this way. In the first place natural uranium consists mainly of the isotope $^{238}_{92}\mathrm{U}$, while the fissionable isotope $^{235}_{92}\mathrm{U}$ comprises only 0.7 percent of the total. Most of the neutrons produced in a given fission process are captured by $^{238}_{92}\mathrm{U}$ nuclei without any further production of neutrons. The escalation of the fission process thus becomes impossible. However, even a sample of pure $^{235}\mathrm{U}$ will not always explode spontaneously. If it is sufficiently small, many of the neutrons will escape into the surroundings without causing further fission. The sample must exceed a **critical mass** before an explosion results. In an atomic bomb several pieces of fissionable material, all of which are below the critical mass, are held sufficiently far apart for no chain reaction to occur. When these are suddenly

brought together, an atomic explosion results immediately.

Two types of atomic bombs were developed concurrently during the war: a relatively simple and a more complex . The gun-type design proved impractical to use with , and therefore a simpler gun-type called was developed that used , an that makes up only 0.7 percent of natural . Chemically identical to the most common isotope, , and with almost the same mass, it proved difficult to separate the two.

In parallel with the work on uranium was an effort to produce , which was discovered at the in 1940. After the feasibility of the world's first artificial nuclear reactor, the , was demonstrated in 1942 at the in the , the Project designed the at Oak Ridge and the production reactors at the in , in which uranium was irradiated and into plutonium. The plutonium was then chemically separated from the uranium, using the . The plutonium implosion-type weapon was developed in a concerted design and development effort by the Los Alamos Laboratory.

Isotopic Enrichment

A great deal of the five years of the Manhattan Project was spent in separating the 0.7 percent of ^{235}U from the more abundant ^{238}U. Most of this work was performed at the at . This was done by preparing the gaseous compound UF_6 and allowing it to effuse through a porous screen. Each effusion resulted in a gas which was slightly richer in the lighter isotope. Repeating this process eventually produced a compound rich enough in ^{235}U for the purposes of bomb manufacture.

Synthesis of Plutonium

In parallel with the work on uranium was an effort to produce , which was discovered at the in 1940. Only the first bomb dropped on Japan used uranium. The second bomb used the artificial element *plutonium*, produced by the neutron bombardment of ^{238}U:

$$^{235}_{92}\mathrm{U} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{239}_{94}\mathrm{Pu} + 2\,{}^{0}_{-1}\beta$$

Fission of Pu-239 occurs in much the same way as for U-235, giving a variety of products; for example,

$$^{239}_{94}\mathrm{Pu} + {}^{1}_{0}\mathrm{n} \rightarrow {}^{90}_{38}\mathrm{Pu} + {}^{147}_{56}\mathrm{Ba} + 3\,{}^{1}_{0}\mathrm{n}$$

Again this is a highly exothermic reaction yielding about the same energy per mole (20 000 GJ mol^{-1}) as ^{235}U.

This immense amount of energy released is the reason why plutonium and uranium were used. By comparison, 1 ton of TNT yields about 4 GJ mol^{-1} of energy. This means that the atomic bombs had about 5,000 times more energy than 1 ton of TNT, that is a huge amount of energy, and explains why the bomb was so effective and destructive.

Nuclear explosion

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Video 7.8.1 *Nuclear explosion power comparison.*

Although the Fat Man atomic bomb was extremely powerful and destructive, it is not even close to the power of subsequent nuclear bombs. The atomic age was set forth by the United States on those two fateful days in August. In the subsequent years many other countries harnessed the power of nuclear weapons, including Russia, China, Britain, and France. The "high point" of the atomic age came with the explosion of the *Tsar Bomba* by Russia. This bomb was equal to 50 megatons of TNT, or 2,500 times more powerful than the Fat Man bomb. It was said that when the bomb was detonated the shock wave in the ground could be felt over 600 miles away (the bomb wasn't even detonated on the ground either, but rather in mid air!). The advance and

use of nuclear weapons seemed to be growing at an extraordinarily fast pace. However, countries stepped up and signed various treaties, including the Nuclear Non-Proliferation Treaty, to stop the increase in nuclear weapons and hopefully the use of any more in the future.

Bomb Construction

At Los Alamos, New Mexico, J. Robert Oppenheimer, and his research group were given the task of assembling and testing the fission bomb. Both types of fuel that were manufactured (U-235 and Pu-239) would be placed separately into two different bomb designs. The smaller of the two bombs would contain U-235 and was code-named "Little Boy." The larger bomb, which housed Pu-239, was named "Fat Man." Of the two, only the Pu-239 bomb would be tested before it was dropped in wartime.

The basic design of the U-235 bomb is shown in the figure below. To prevent the spontaneous detonation of an atomic bomb, the fissile U-235 is kept in a subcritical configuration. It is then rapidly assembled into a supercritical mass using conventional explosives. Once the bomb has achieved this mass, any neutron introduced into it will be likely to initiate a chain reaction. The mechanism for "Little Boy" was a gun that fired one subcritical piece of U-235 into another to form a supercritical mass (Figure 7.8.2). The pieces had to be assembled within a time less than the average time between appearances of spontaneous neutrons from either U-235 or cosmic radiation. A conventional explosive in an artillery barrel could fire the U-235 mass at speeds of a few millimeters per second, fast enough to prevent a fizzle caused by a spontaneous neutron setting off a premature chain reaction.

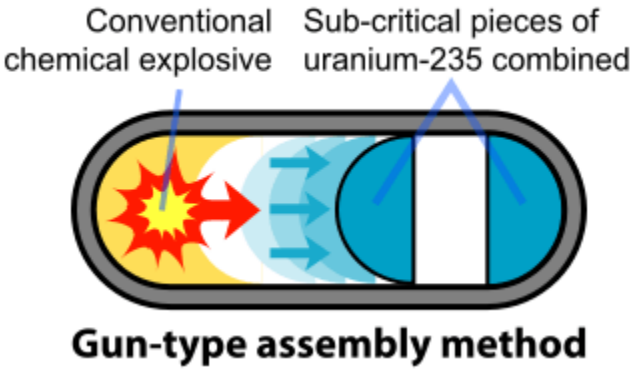


Figure 7.8.2: Little Boy: Gun-Type Device. (Public Domain).

Originally, the gun-type mechanism was planned for both the U-235 and Pu-239 weapons. However, a problem arose with the Pu-239 bomb that required a different assembly mechanism because of the small amount of Pu-240 that is produced with the Pu-239 in the reactor. Pu-240 emits large numbers of neutrons spontaneously: 1,030 neutrons per gram per second compared with 0.0004 neutrons per gram per second for U-235. Even at a concentration of 1% Pu-240 in the fissile Pu-239, the required mass of Pu emits 52,000 neutrons per second or one neutron every 20 microseconds. Thus, it is very probable that a neutron from Pu-240 will initiate a premature chain reaction during the critical last 100 microseconds of the critical mass in a gun-type assembly. This problem was discovered in mid-1944, well after the start of the construction of the massive Hanford plutonium production facilities.

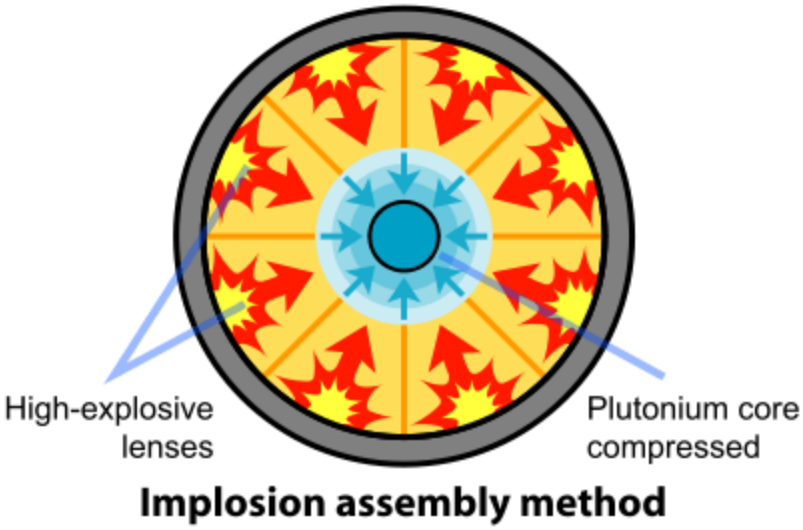


Figure 7.8.3: Fat Man: Implosion-type bomb. (Public Domain).

The removal of the Pu-240 was impractical. So the scientists and engineers looked for a faster method of assembling the plutonium. A mechanism based on implosion provided the solution to this problem. In this design, the fissile material is shaped into a single sphere with a mass slightly less than critical (Figure 7.8.3). Layers of carefully shaped high explosives surround the sphere. When the explosives are detonated, the force of the shock wave compresses the fissile material into a smaller volume, forming a supercritical mass. This method of assembly is much faster than the gun-type mechanism and thus eliminates the problems resulting from spontaneous neutron emission of Pu-240. The spherical mass resulted in a pumpkin-shaped weapon called "Fat Man".

Example 7.8.1

Watch the from the 23-minute mark through the 32-minute mark of the third installment of the History Channel's Manhattan Project and answer the questions below.

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1. What were the problems with the two different types of fuel that were to be used in the fission weapons?
2. What was the configuration or set-up of the Pu-239 bomb?
3. The combination of electromagnetic separation and gaseous diffusion enriched enough U-235 to produce how many bombs?
4. Why was the U-235 bomb never tested?
5. When was the construction of the U-235 completed?
6. How much did Truman know about the nuclear bombs when FDR was alive?
7. Who was selected to pilot the plane for the nuclear weapons?
8. What island would serve as a base to house the nuclear bombs before delivering them to their targets?

In July 1945, the United States had enough fissile material for one uranium and two plutonium weapons. The scientists and engineers felt confident that the gun-type assembly mechanism for Little Boy would function properly. Besides, they did not have the material for a test device. They were less confident about the implosion mechanism on the plutonium weapon and felt that a test was necessary. On July 16, 1945, the first nuclear device, known as "The Gadget", was placed on a 100-foot tower and successfully detonated in the Alamogordo Desert, 200 miles south of Los Alamos (Figure 7.8.4).

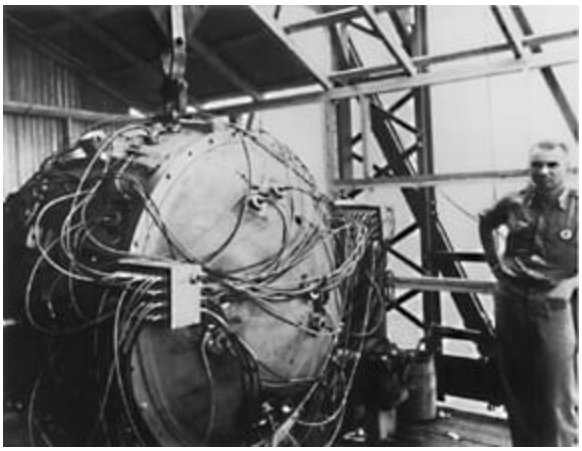


Figure 7.8.4: The Gadget at the Trinity Test Site, 1945. Courtesy of the U.S. Department of Defense

As the war with Japan continued and a costly allied invasion loomed as a real possibility, President Truman approved the use of nuclear weapons against selected Japanese targets. The U.S. Army Air Force received orders to use these weapons anytime after August 3, 1945. On August 6, "Little Boy" was dropped on Hiroshima (Figure 7.8.5).



Figure 7.8.5: Little Boy - The only U-235 weapons constructed during WWII. Courtesy of the U.S. Department of Defense

Little Boy was a uranium weapon containing 141.4 pounds of fissionable material containing 82.7% U-235. Only about two pounds fissioned, releasing an energy equivalent to 15-16,000 tons of TNT. The immediate effects of the blast killed an estimated 70,000 people, and by the end of 1945 an additional 20,000 to 70,000 deaths occurred, many due to lack of adequate medical resources.

Three days later, "Fat Man" destroyed a large part of Nagasaki (Figure 7.8.6). Fat Man contained 13.6 pounds of Pu-239, of which only 2 pounds underwent fission. The explosive yield was equivalent to about 22,000 tons of TNT. Fat man resulted in 35,000 immediate deaths. By the end of 1945, at least 70,000 people perished in this event.



Figure 7.8.6***:*** *Fat Man – One of many Pu-239 nuclear weapons constructed during WWII. Courtesy of the U.S. Department of Defense*

Example 7.8.2

Start the video below at the 23:20 mark and stop it at the 29-minute setting. Then, answer the questions below:

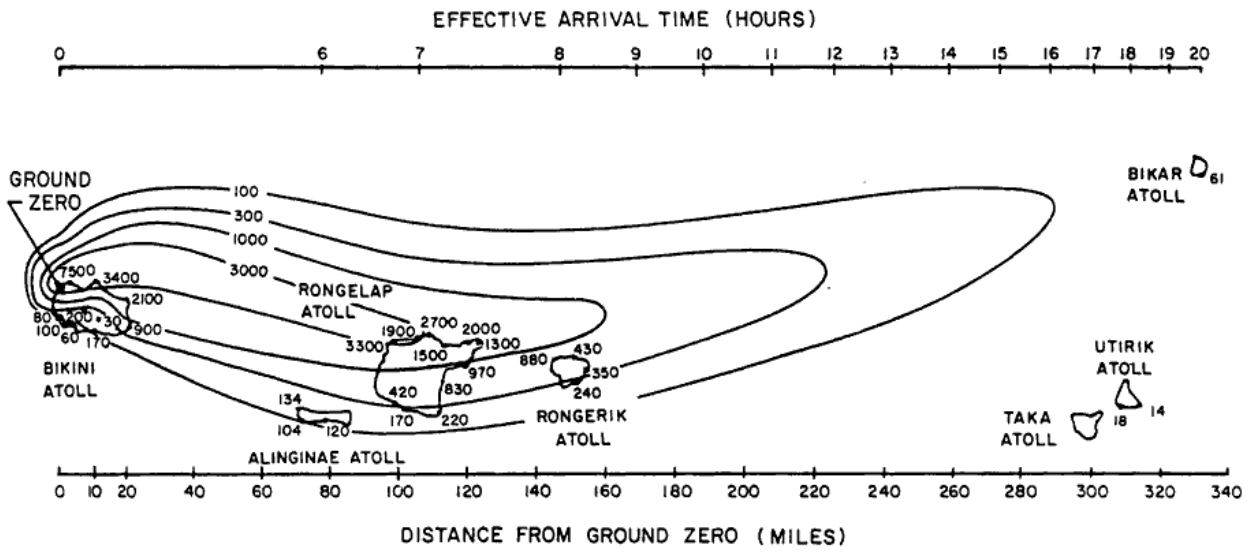
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9. When and where was the first nuclear bomb tested?
10. Was the bomb placed on the ground?
11. What were scientists afraid the first nuclear bomb would do?
12. What did Edward Teller (father of the fusion bomb) apply to his skin before the bomb was ignited? Why did he do this?
13. What did the bomb do to the tower and the surrounding sand?
14. In the photograph, where did General Grove focus his attention?

Nuclear Fallout

Nuclear fallout refers to the radioactive material that "falls out" of the atmosphere after a nuclear explosion (reactor or bomb). It consists of dust and radioactive particles that can contaminate an area with radioactivity and pose a huge health hazard to biological organisms. It can contaminate the animal food chain which can have drastic effects on the affected region. Weather has a huge impact on fallout, wind currents can spread radioactive fallout either over a large area, such as in the case of *Castle Bravo*, or not.



alt

Figure 7.8.2: The 450 km (280 mi) fallout plume from 15 Mt shot Castle Bravo, 1954. (Pubic Domain; United States Department of Energy via Wikipedia)

In regards to nuclear reactions, fission produces more fallout products that fusion. Intense heat of the latter type of reactions reduces the amount and type of radioisotopes produced. Although more powerful, fusion is a "cleaner" type of nuclear reaction. By increasing the fission component of a fusion weapon, bombs can be designed to generate more fallout products.Of the different types of isotopes that fission produces, the most worrisome are Sr-90, Cs-137,and I-131. The first of these isotopes has a half-life of 28.8 years and decays by gamma and beta emission. Once ingested or inhaled, Sr-90 deposits in teeth, bone,and bone marrow. Sr-90 exposure can lead to bone cancer and/or leukemia. Cs-137 ($t_{1/2}$ = 30.17 years) emits beta and gamma radiation. This ionizing radiation deposits in soft tissues of the body. In addition, this particular radioisotope is water soluble and can easily contaminant food and water supplies. I-131 had a half-life over 8.1 days and undergoes gamma/beta decay. Once this fallout product enters the body, it attacks the thyroid gland. Extensive damage could affect heart rate, blood pressure, body temperature, and childhood growth. In addition, I-131 exposure could result in thyroid cancer.

Nuclear Fallout

Watch the fallout video and then answer the questions below:

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Summary

- In an atomic bomb several pieces of fissionable material are held sufficiently far apart for no chain reaction to occur. When these are suddenly brought together, the critical mass is reached and an atomic explosion results immediately.
- The two fission bombs assembled at Los Alamos, New Mexico, by J. Robert Oppenheimer, and his research group were manufactured with two types of fuel namely, U-235 and Pu-239.
  - The smaller of the two bombs would contain U-235 and was code-named "Little Boy." This bomb was dropped in Hirsohima, Japan on Aug.6, 1945.
  - The larger bomb, which housed Pu-239, was named "Fat Man." This bomb was dropped in Nagazaki, Japan on Aug.9,

1945.
- Nuclear fallout refers to the radioactive material that "falls out" of the atmosphere after a nuclear explosion (reactor or bomb).

Contributors and Attributions

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Learning Objective

- Know the benefits and consequences of nuclear energy.

Nuclear Power Generation

The generation of electricity is critical for operation of businesses, health care delivery, schools, homes, and other areas requiring the use of electrical power. According to 2011 statistics, coal is used for 42% of the total power generated, with natural gas being employed for another 25%. Nuclear power plants are employed in about 19% of the cases, with renewable energy sources supplying the last 13%. All of these fuels are used to heat water to generate steam. The steam then turns a turbine to generate electricity.

Nuclear power plants use the energy from nuclear fission to produce electricity. **U-235** is the preferred nuclear fuel because when its atoms are split (fissioned), they not only emit heat and high energy radiation but also enough neutrons to maintain a chain reaction and provide energy to power a nuclear power plant. Uranium is found in rocks all over the world but is relatively rare and the supply is finite making it a nonrenewable energy source.

Uranium usually occurs in combination with small amounts of other elements and once it is mined, the U-235 must be extracted and processed before it can be used as a fuel in a nuclear power plant to generate electricity. The processed uranium is formed into fuel rods and then bundled into fuel assemblies. Fuel assemblies are stored onsite until they are needed by the reactor operators. At this stage, the uranium is only mildly radioactive, and essentially all radiation is contained within the metal tubes. When needed, the fuel is loaded into a reactor core (Figure $7.8.1$). Typically, about one third of the reactor core (40 to 90 fuel assemblies) is changed out every 12 to 24 months.

The most common type of reactors are the **pressurized water reactors** (PWR) (Figure $7.8.1$) in which water is pumped through the reactor core and heated by the fission process. The water is kept under high pressure inside the reactor so it does not boil. The heated water from the reactor passes through tubes inside the steam generator where the heat is transferred to water flowing around the tubes in the steam generator. The water in the steam generator boils and turns to steam. The steam is piped to the turbines. The force of the expanding steam drives the turbines, which spin a magnet in coil of wire – the generator– to produce electricity.

After passing through the turbines, the steam is converted back to water by circulating it around tubes carrying cooling water in the condenser. The condensed steam – now water – is returned to the steam generators to repeat the cycle.

The three water systems (condenser, steam generator, and reactor) are separate from each other and are not permitted to mix. Water in the reactor is radioactive and is contained within the containment structure whereas water in the steam generator and condenser is nonradioactive.

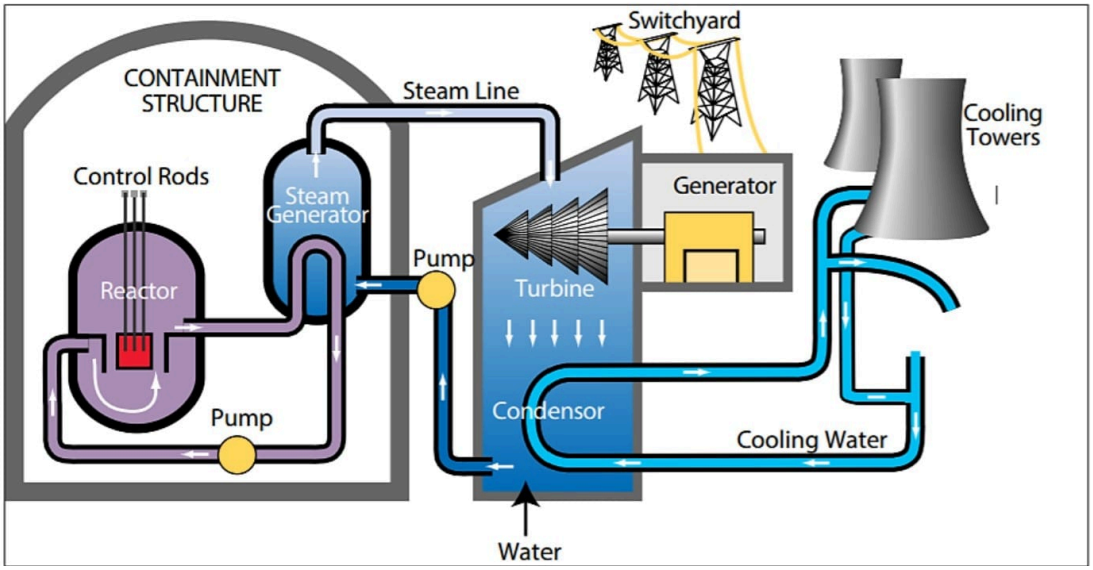


Figure 7.8.1**:** *A schematic diagram of a pressurized water reactor (PWR), the most common type of nuclear reactor. Diagram from Tennessee Valley Authority (public domain). www.tva.com*

Benefits of Nuclear Energy

By using fission, nuclear power plants generate electricity without emitting air pollutants like those emitted by fossil fuel-fired power plants. This means that financial costs related to chronic health problems caused by air pollutants such as particulate material, carbon monoxide, nitrogen oxides and ozone among others are significantly reduced. In addition nuclear reactors do not produce carbon dioxide while operating which means that nuclear energy does not contribute to the global warming problem.

Another benefit of nuclear energy over fossil fuels especially coal is that uranium generates far more power per unit weight or volume. This means that less of it needs to be mined and consequently the damage to the landscapes is less especially when compared to the damage that results from coal mining such as mountaintop removal.

Nuclear power is also used to propel ships. The turbine can be connected to a propeller system. The rotating turbine shaft will turn the propeller to move the ship.



Figure 7.8.2 *Nuclear submarine.*

The Drawbacks of Nuclear Energy

The main environmental concern related to nuclear power is the creation of radioactive wastes such as uranium mill tailings, spent (used) reactor fuel, and other radioactive wastes. These materials can remain radioactive and dangerous to human health for thousands of years. Radioactive wastes are classified as low-level and high-level. By volume, most of the waste related to the nuclear power industry has a relatively low-level of radioactivity. **Uranium mill tailings** contain the radioactive element radium, which decays to produce **radon**, a radioactive gas. Most uranium mill tailings are placed near the processing facility or mill where they come from. Uranium mill tailings are covered with a barrier of material such as clay to prevent radon from escaping into the atmosphere, and they are then covered by a layer of soil, rocks, or other materials to prevent erosion of the sealing barrier.

Spent fuel rods contain a variety of products, consisting of unstable nuclei ranging in atomic number from 25 to 60, some transuranium elements, including plutonium and americium, and unreacted uranium isotopes. The unstable nuclei and the transuranium isotopes give the spent fuel a dangerously high level of radioactivity. The long-lived isotopes require thousands of years to decay to a safe level. The ultimate fate of the nuclear reactor as a significant source of energy in the United States probably rests on whether or not a politically and scientifically satisfactory technique for processing and storing the components of spent fuel rods can be developed.

The other types of **low-level radioactive waste** are the tools, protective clothing, wiping cloths, and other disposable items that get contaminated with small amounts of radioactive dust or particles at nuclear fuel processing facilities and power plants. These materials are subject to special regulations that govern their handling, storage, and disposal so they will not come in contact with the outside environment.

High-level radioactive waste consists of spent nuclear reactor fuel (i.e., fuel that is no longer useful for producing electricity). The spent reactor fuel is in a solid form consisting of small fuel pellets in long metal tubes called rods. Spent reactor fuel assemblies are initially stored in specially designed pools of water, where the water cools the fuel and acts as a radiation shield. Spent reactor fuel assemblies can also be stored in specially designed dry storage containers. An increasing number of reactor operators now store their older spent fuel in dry storage facilities using special outdoor concrete or steel containers with air cooling. There is currently no permanent disposal facility in the United States for high-level nuclear waste.

When a nuclear reactor stops operating, it must be **decommissioned**. This involves safely removing the reactor and all equipment that has become radioactive from service and reducing radioactivity to a level that permits other uses of the property. The U.S. Nuclear Regulatory Commission has strict rules governing nuclear power plant decommissioning that involve cleanup of radioactively contaminated plant systems and structures, and removal of the radioactive fuel.

A **nuclear meltdown**, or uncontrolled nuclear reaction in a nuclear reactor, can potentially result in widespread contamination of air and water. Some serious nuclear and radiation accidents have occurred worldwide. The most severe accident was the Chernobyl accident of 1986 in the then Soviet Union (now Ukraine) which killed 31 people directly and sickened or caused cancer in thousands more. The Fukushima Daiichi nuclear disaster (2011) in Japan (Figure 7.8.3) was caused by a 9.0 magnitude earthquake that shut down power supply and a tsunami that flooded the plant's emergency power supply. This resulted in the release of radioactivity although it did not directly result in any deaths at the time of the disaster. Another nuclear accident was the Three Mile Island accident (1979) in Pennsylvania, USA. This accident resulted in a near disastrous core meltdown that was due to a combination of human error and mechanical failure but did not result in any deaths and no cancers or otherwise have been found in follow up studies of this accident. While there are potentially devastating consequences to a nuclear meltdown, the likelihood of one occurring is extremely small. After every meltdown, including the 2011 Fukushima Daiichi disaster, new international regulations were put in place to prevent such an event from occurring again.



Figure 7.8.3 (a) After the accident, contaminated waste had to be removed, and (b) an evacuation zone was set up around the plant in areas that received heavy doses of radioactive fallout. (credit a: modification of work by "Live Action Hero"/Flickr)

The processes for mining and refining uranium ore and making reactor fuel require large amounts of energy. Nuclear power plants have large amounts of metal and concrete, which also require large amounts of energy to manufacture. If fossil fuels are used for mining and refining uranium ore or in constructing the nuclear plant, then the emissions from burning those fuels could be associated with the electricity that nuclear power plants generate.

Nuclear Reactor Design and Operation

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Video 7.8.1 *The design and safe operation of a nuclear reactor.*

Summary

- The importance of nuclear power in generating electricity is described.
- The operation of a nuclear power plant is described.
- The drawbacks of nuclear energy is described.

Contributors and Attributions

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