

THE COMPLETE
IDIOT'S
GUIDE[®] TO

Chemistry

Third Edition

by Ian Guch



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Firing Up the Bunsen Burner

Part

1

Welcome to chemistry! You're probably shaking in your boots, wondering what terrible things you'll find in this book.

Don't worry! We're going to ease into this chemistry thing by learning the basic vocabulary—things like the metric system, data manipulation, and the atom. By the time you're done with this part, you'll feel a lot more comfortable with chemistry than you do now.

Okay, take a deep breath and turn the page.

Measuring Up

1

In This Chapter

- The metric system
- How to perform unit conversions
- Accuracy and precision
- Significant figures
- Experimental error

We all know how to find the lengths of our thumbs, and we probably know how to figure out our weights (though many of us also wish we didn't). These are simple everyday measurements that we all learned to do as little kids.

Of course, things get more complicated when you take a chemistry class. For those of us used to thinking in inches, pounds, and degrees Fahrenheit, we may balk at having to use the mysterious centimeters, kilograms, and degrees Celsius. Not only that, but it's no longer enough to look at a ruler and just read the length of our thumb: we now have to worry about the mysterious term *significant figures* and whether we have the right number of them.

Calculations are worse. Instead of putting numbers into a calculator and writing down the result, we now include only *some* of the digits. Again, the concept of significant figures rears its ugly head in unusual and unpredictable ways. What's a chemistry student to do?

Not to worry. All these terms are simple to understand and use in the laboratory. The problem lies not in the terms or ideas; it lies in the way they're explained. In this chapter, you arm yourself with the necessary tools for collecting and understanding chemical data.

Quantity	Unit	Symbol
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

Sometimes it's not handy to use the units described previously. For example, if you're trying to find the distance from your house to the northernmost point in Canada, the unit meter is difficult to use because of the lengthy distance. As a result, the metric system uses a series of prefixes that we can add to the SI units to make them easier to use. The following table shows the most commonly used prefixes.

Selected Prefixes for Metric Units

Prefix	Symbol	Meaning
giga	G	one billion, or 1,000,000,000 (10^9)
mega	M	one million, or 1,000,000 (10^6)
kilo	k	one thousand, or 1,000 (10^3)
deci	d	one tenth, or 0.1 (10^{-1})
centi	c	one hundredth, or 0.01 (10^{-2})
milli	m	one thousandth, or 0.001 (10^{-3})
micro	μ	one millionth, or 0.000001 (10^{-6})
nano	n	one billionth, or 0.000000001 (10^{-9})

The magic of these prefixes is that this system allows us to convert from really big quantities (like the distance from New York to L.A., in miles) to much smaller quantities (like the length of your driveway, in inches) with greater ease by moving the decimal point around. There's no messy dividing by 5,280, or 36, or 12, or whatever, like you've been doing most of your life.



CHEMISTRIVIA

Although one millionth of a meter is officially a micrometer in SI units, it's far more commonly referred to as a micron.

Let's say that you live in Washington, D.C., and you're commuting to Baltimore. Instead of saying that you have to travel 60,000 meters, you can say that we have to travel 60 kilometers, or 60 km. Because *kilo* multiplies the unit by a thousand, 60 kilometers

Step 2: Write a times sign after this unit, followed by a straight, horizontal line.

71 in × _____

Step 3: In the space below the line, write the unit of the number that you already wrote on the paper.

Because “in” is after the number in this unit, write this below the line:

71 in × _____
in

Step 4: Write the unit of what you’re trying to find on the top of the line.

In the example, you’re trying to find your height in centimeters, so put “cm” above the line:

71 in × $\frac{\text{cm}}{\text{in}}$

Step 5: Write the *conversion factor* in front of the units on the line.

Uh oh. You have a problem. You don’t know what a conversion factor is, which makes this an extremely difficult problem. Fortunately, the guy behind you in line taps you on the shoulder to explain that conversion factors are just numbers that allow you to convert from one unit to another. Even better, he knows that the conversion factor between inches and centimeters is 2.54 centimeters = 1 inch. As you turn to thank him, he melts away into the crowd.



DEFINITION

A **conversion factor** is a number that allows you to convert one set of units to another.

The conversion factor tells you that there are 2.54 centimeters in 1 inch, so write 2.54 in front of “cm” and 1 in front of “in”:

71 in × $\frac{2.54 \text{ cm}}{1 \text{ in}}$

Step 3: Below the line, write the unit of the number you are trying to convert.

$$42 \text{ days} \times \frac{\quad}{\text{days}}$$

So far, so good. No problems yet!

Step 4: Write the unit of what you're trying to find on top of the line.

Here lies a big problem. If you write the unit that you're eventually trying to find, "fortnights," on the top of the line, you can't solve the problem because you don't know how many days are in a fortnight.

However, you know how many weeks are in a fortnight because Officer Friendly told you. Even better, you know that there are seven days in a week. Solving this problem, then, requires two calculations. In the first calculation, you convert days to weeks. In the second, you convert weeks to fortnights. Let's focus on the first calculation, from days to weeks:

$$42 \text{ days} \times \frac{\text{weeks}}{\text{days}}$$



BAD REACTIONS

A common mistake people make when doing unit conversions is to assume that all unit conversion problems can be done in one step. Always keep an eye on the conversion factors you're given and figure out how many steps will be needed before starting your work.

Step 5: Write the conversion factor next to the appropriate unit.

$$42 \text{ days} \times \frac{1 \text{ week}}{7 \text{ days}}$$

Step 6: Solve the math problem you've written, making sure you cancel out any appropriate units.

$$42 \text{ days} \times \frac{1 \text{ week}}{7 \text{ days}} = 6 \text{ weeks}$$

weigh 10.0 grams, which would make the repeated measurements of 8.8 grams precise (repeatable) but not accurate (correct).



DEFINITION

Accuracy refers to a measurement that's close to the actual value of the item being measured. **Precision** relates to repeated measurements of the same object always resulting in the same value. Accurate data is always precise, but precise data is not necessarily always accurate.

If you've ever played darts, you already have a pretty good idea of the difference between accuracy and precision. When people play darts for the first time, the darts typically fly all over the room, hitting the wall, shattering windows, and terrifying the parakeet. At this point, the darts are flying with neither precision nor accuracy.

After practicing for a few hours, people frequently find that their darts are all grouped closely, but not in the center of the dartboard. We would say that these darts were thrown precisely (they're in the same location) but not accurately (they're not in the center of the target). Only with extended practice do the darts eventually become grouped in the center of the board, showing that they have been thrown both precisely and accurately.

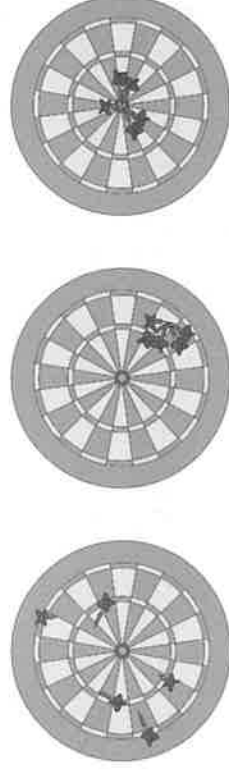


Figure 1.1: The dartboard on the left shows darts thrown with neither precision nor accuracy, the one in the middle shows darts thrown precisely but not accurately, and the one on the right shows darts thrown both accurately and precisely.



YOU'VE GOT PROBLEMS

Problem 7: Devise a scenario in which a precise measuring instrument is not also accurate.

Significant Figures

Now that you've learned a bit about accuracy and precision, it's time to tackle the sticky subject of significant figures.

When asked, how heavy should you say the paper clip is? Examining the previous data, you can see that each measurement agrees that the paper clip weighs about 1 gram, so this digit is probably valid. Looking at the tenths of a gram, it looks as if all the measured numbers round to 1.1 grams. After that, the digits appear random, so they probably don't have any real meaning. As a result, you should say that the paper clip weighs 1.1 grams. Because this answer has two digits, this measurement is said to have two significant figures.

It's important to write the correct number of significant figures when collecting data because it gives people a solid idea of how good your data is. To get the most out of your measurements, use the following rules when collecting data:

- When using analog instruments (such as rules, which don't have digital displays) always write the number of digits that you can directly measure with the equipment, plus an extra digit that you estimate.

Take a look at Figure 1.2. Normally, you don't spend much time thinking about the number of digits you use to write down the length of a paper clip. However, to do this correctly, you need to write down the number of digits that can be directly measured, plus an extra digit that you estimate.

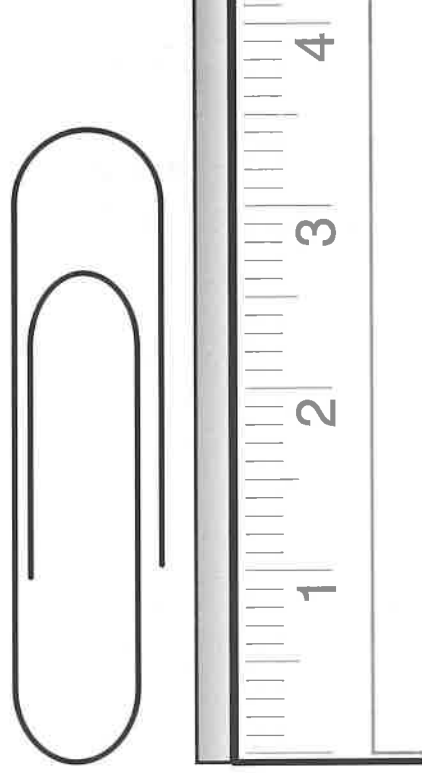


Figure 1.2: As you can see, this ruler has markings that show millimeters. As a result, you can estimate the length of the paper clip to the nearest tenth of a millimeter—in this case, a length of 3.44 cm.

Let's look at another example (see Figure 1.3).



YOU'VE GOT PROBLEMS

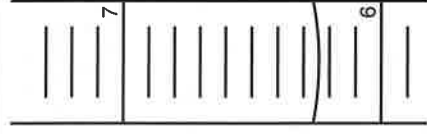


Figure 1.4: Problem 8: What is the volume of the liquid in the graduated cylinder?

Determining Significant Figures in Someone Else's Data

You can tell a lot about the quality of a measurement by the number of significant figures shown. For example, if someone has used poor-quality equipment, you'll usually see only one or two digits, indicating a lack of precision. If someone has used high-quality equipment, you may see four or five digits.

The following rules determine the number of significant figures in a measurement:

- All nonzero digits are significant. If the mass of a paper clip is recorded as 1.21 grams, all three of these digits give you useful information, so they are considered significant figures.
- Zeros between nonzero digits are significant. If the mass of a paper clip is 10.01 grams, all four of these digits are significant.
- All zeros written to the left of all the nonzero digits are *never* significant. This is important for measurements with values less than 1, such as 0.0023 grams. For this measurement, you would say that the 2 and the 3 are significant because of rule #1, but the zeros are not significant because they are to the left of these digits.

This seems counterintuitive. After all, don't those zeros give you useful information? Isn't useful information the point of significant figures? Yes, those digits do give you some useful information. To understand why they're not significant, imagine the following scenario.



THE MOLE SAYS

The data in this table give you a good idea of why you need significant figures. A mathematician would say that all three of these numbers are equivalent, but we scientists know that the difference in these numbers tells us how much we can trust the measurement.

Finally, following are the other rules for determining the number of significant figures in a measurement:

- When using scientific notation, ignore the exponential part (the “ $\times 10^{-4}$ ” part) when finding significant figures. For example, 2.30×10^{-6} grams has three significant figures: the 2, the 3, and the 0.
- Some numbers don’t use significant figures at all and are said to be infinitely precise. These include numbers obtained from counting actual objects (for example, somebody can’t give you a fraction of a penny) and numbers that are defined as having some value (for example, there are *exactly* 1,000 meters in 1 kilometer). Any measured data that you take in the lab, however, follows the rules for significant figures.



YOU’VE GOT PROBLEMS

Problem 9: How many significant figures do the following measured numbers have?

- a) 2.490 grams
- b) 1,010 grams
- c) 0.01010 grams

Using Significant Figures in Calculations

Think back to the example when you found the density of foam rubber to be 0.11111111 g/mL based on fishy data. Though we’ve discussed significant figures and what they mean, we still haven’t discussed how to express calculations in significant figures. After all, it doesn’t do you much good to know how good your data is if you then screw up the meaning when you use it to calculate something!

As it turns out, the significant figure rules for calculations are fairly simple:

- When multiplying or dividing, determine which of the numbers has the smallest number of significant figures. The answer to the calculation should be written with the same number of significant figures as this number.

Types of Errors

Two types of errors are recognized in the collection of data: systematic error and random error.

Systematic error (also known as determinate error) occurs when something goes wrong with an experiment that causes the values to be skewed by the same amount every time. For example, my wife claims that our bathroom scale always measures her weight to be 3 pounds heavier than it really is. To compensate for this error, she subtracts three pounds from whatever value it gives for her weight. Systematic error can be caused by almost anything: equipment failure, human error, demonic possession, shoes worn when being weighed, and so on.



DEFINITION

Systematic error occurs when you get the same mistake every time you do an experiment. **Random error** occurs when the mistake varies randomly. Compensating for systematic error is much easier than compensating for random error.

Random error (a.k.a. indeterminate error) is error that can't be compensated for. This type of error may be caused by just about anything. For example, if I jump on the bathroom scale and find that my weight is 212 pounds the first time, 214 pounds the second time, and 209 pounds the third time, I can safely assume that the differences are caused by random fluctuations in the functioning of the scale and not huge fluctuations in my weight. As a result, my weight is probably close to the average of these values, or 212 pounds. Similarly, it is common—and correct—to average a series of experimental measurements to compensate for random error. The more, the better. A lot of scientists make huge numbers of measurements in an experiment, to compensate for random errors. If these measurements all give the same answer, the scientists do a happy little dance. If the measurements are all over the place, they report this and use the average as their answer.



CHEMISTRIVIA

A friend of mine who works with microscopes that are sensitive enough to measure individual atoms found that he was getting random error from the vibrations of people walking outside the building in which he worked. As a result, he now does most of his work late at night.

The History of the Atom

Chapter

2

In This Chapter

- The ancient Greek view
- Dalton's laws
- Rutherford's "plum pudding" model
- Isotopes

Most of us think we have a pretty good idea of what atoms are. We know that they're tiny and that they're important in chemistry. To demonstrate your knowledge, let's play a game. Define *atom* for me. You have 10 seconds ...

Time's up! Let's hear it. What do you mean it's harder to define than you thought?

The simplest ideas in chemistry are sometimes the hardest to articulate. If defining *atom* gives you trouble, don't feel bad. After all, it took science more than 2,000 years to get a good idea of what the atom looks like—and even now, it still retains some of its mystery.

As a result, it's pretty hard to understand the atom all at once. Instead of diving headfirst into the deep end of quantum mechanics, let's float inside a big goofy inner tube in the shallow end for a while.

What's an Atom?

An *atom* is the smallest chunk of an element that has the same properties as a larger chunk of that element. An element, in turn, is the most basic quantity and type of matter that you use in chemistry. When you look at the periodic table, for example, you find the more than 100 types of atom from which all matter is constructed.

**CHEMISTRIVIA**

Leucippus was the first person to come up with the atomic theory, but his student Democritus became far more famous than his teacher and is sometimes credited with the idea.

The Greeks weren't exactly clear on the nature of these atoms because they had no way of testing their theories. Thus, they were stuck with the idea that there were only four elements (earth, air, fire, and water) and that they differed from each other in their shapes, sizes, and structures.

Because of a complete lack of evidence and because Aristotle had a more popular competing idea in which atoms were replaced with continuous, nonlumpy matter at all levels, this theory fell out of fashion until the late Renaissance. Fortunately, it was a common belief after the fall of the Roman Empire that the ancient Greeks knew pretty much everything, so that knowledge was preserved until scientists were ready to use it again.

Some Random Observations

For nearly two millennia, nobody did much experimenting to discover the nature of matter. However, in the late eighteenth century, this changed with the discoveries of Antoine Lavoisier. Lavoisier found that the weight of the products of a chemical reaction is the same as the weight of the reactants. This is now known as the *law of conservation of mass*.

**DEFINITION**

The **law of conservation of mass** states that the weights of the products in a chemical reaction are equal to the weights of the reactants. No matter what chemical changes occur, matter is neither created nor destroyed.

To us, the law of conservation of mass is old news. After all, it seems intuitive that if we make a sandwich with 100 grams of turkey and 50 grams of bread, the final weight of the sandwich will be 150 grams. However, let's look at an example that confused people back in the old days.

Imagine that your significant other has dumped you, and you're burning all of his/her pictures. When you start the bonfire, the weight of the pictures is 250 grams. However, when you weigh the ashes afterward, you find that they weigh only 50 grams. It looks as if matter was destroyed, perhaps turned into energy that created the fire.

We know now that this isn't the case. When you burn something, the ashes account for some of the weight; soot, smoke, and water vapor account for the rest of it. The weight is still around *somewhere*, but it has moved to a new location where you can't see it.

John Dalton Puts It All Together

In 1808, John Dalton used the existing knowledge about atoms to devise a series of rules describing the behavior of atoms, referred to as Dalton's Laws.

- All matter is made of tiny, indestructible particles called atoms.
- All atoms of a given element are identical. For example, if you purify gold by two different methods, each sample of gold will contain atoms with the same chemical and physical properties.
- Atoms of different elements have different properties. Some elements might share some properties, but no elements have identical sets of chemical and physical properties.
- Atoms are neither created nor destroyed in chemical reactions; atoms obey the law of conservation of mass. When chemical reactions take place, only the arrangements of atoms change, not the weight.
- Atoms of different elements form compounds in whole-number ratios. For this reason, chemical compounds have formulas such as H_2O , not $\text{H}_{2.1}\text{O}_{0.8}$.

Nowadays, we know that some of these rules are true and some are false. Atoms aren't indestructible—we can break them apart in nuclear reactions. And atoms of the same element don't always share identical properties. Still, Dalton didn't do so bad, considering that he never directly saw an atom.



CHEMISTRIVIA

In addition to his work with the atom, John Dalton was the first scientist to write a paper on the subject of color blindness (a condition he personally experienced). In honor of this paper, the type of color blindness he experienced is still referred to as daltonism.

J. J. Thomson: A Man and His Dessert

In 1897, British physicist J. J. Thomson performed an experiment that applied voltage across two wires (called electrodes) in a vacuum tube. A glowing beam of particles was observed to travel from the cathode (the negative electrode) to the anode (the positive electrode). Because these light rays originated at the cathode, they were called, creatively enough, cathode rays.

During the course of his experiments, Thomson determined that the mass of electrons was tiny compared to the mass of the overall atom. This led him to conclude that electrons are small and light compared to whatever contained the positive charge in an atom. As a result, he devised the “plum pudding” model of the atom. In this model, Thomson asserted that the atom was one big blob of positive charge (the pudding) with small particles of negative charge (electrons/plums) embedded in it. Because not many people are familiar with plum pudding, a chocolate chip cookie is probably a more modern example.

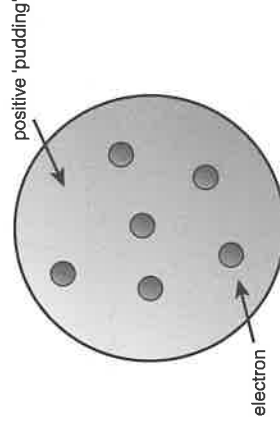


Figure 2.3: Named after Thomson's dessert, the plum pudding model portrays the atom as a big ball of positive charge containing small particles with a negative charge. In this picture of a cookie, the dough represents the positive charge and the chips represent the negatively charged electrons.

Today historians believe that, had Thomson been diabetic, the course of atomic theory would have been vastly different.

Rutherford and His Alpha Particles

When radioactivity was discovered in 1896, several scientists decided that it was the cool new thing to study. During the next decade, scientists studied the decay of radioactive elements into smaller, energetic particles. In time, they discovered three types of radiation: alpha radiation, which consists of helium nuclei with a +2 charge; beta radiation, which consists of electrons; and gamma radiation, which consists of energetic light. (You'll learn more about radiation in Chapter 26.)

One of the big researchers in the field of radiation at the time was Ernest “Radioactive Man” Rutherford. One day in 1909, Rutherford was messing around in the lab, shooting alpha particles at a thin piece of gold foil to see how they interacted with the presumably gooey pudding studded with a few small plums postulated by Thomson. Though most of the particles shot straight through, some of the particles were deflected at large angles. Some of the alpha particles even flew back at the radiation source. This was surprising, because the goo shouldn't have deflected the particles at all, and the plums should have deflected the alphas only a little. This was definitely not what Rutherford expected.

**CHEMISTRIVIA**

The gold foil experiment was performed one year after Rutherford received the Nobel Prize in Chemistry for other work with radiation. This was one smart dude.

Get to the Point, Already: What We Believe Now

We now know that the fundamental parts of an atom are the protons and neutrons, located in the nucleus, and the electrons, located in “orbitals” outside the nucleus:

Particle	Location	Mass	Charge
proton	nucleus	~ 1 amu	+1
neutron	nucleus	~ 1 amu	0
electron	orbitals	$1/1836$ amu (~ 0)	-1

To fully understand this table, you need to know some atomic vocabulary:

- Orbitals are where the electrons reside in an atom. We discuss this at great length in the next chapter.
- The nucleus of an atom is where all the positive charge and most of its mass are located. The neutrons are also found in the nucleus, but they have no charge.
- The term *amu* stands for atomic mass unit. Because nuclear particles are tiny, discussing their weights in kilograms doesn't make much sense (protons and neutrons each weigh approximately 1.67×10^{-27} kg; electrons weigh approximately 9.11×10^{-31} kg).

The number of protons in an atom defines what element is present. For example, if an atom has one proton, it's hydrogen, regardless of how many neutrons or electrons it has. Atoms with the same numbers of protons and electrons are always neutral because the positive and negative charges cancel each other out.

Some of you might wonder why atoms need neutrons at all. Neutrons are like my lazy Uncle Bob. Uncle Bob isn't really much of a nuisance—he just sits around the house in his underwear drinking beer and watching TV. What good is Uncle Bob, anyway?

But how can there be more than one possible number of neutrons in an atom? Let's go back to the example of Uncle Bob to see why.

On weekends, my Uncle Bob invites his unshaven, lazy friends over to the house to watch football. At any one time, there are between 4 and 17 “worthless bums” (to quote my aunt) hanging around the house, throwing Chex Mix at the TV. As is often the case, my aunt and cousins start to fight at some point during the game.

With superhuman speed, Uncle Bob and his loyal sidekicks spring into action. “Shut up, all of you!” they scream. Whether there are 4 or 17 guests in the house, all of them work together like a fine-tuned machine to get my aunt and cousins to be quiet.

The same thing works with neutrons in an atom's nucleus. For many atoms, there can be several different numbers of neutrons that serve to stabilize the positive charge in the nucleus. For example, the three protons in lithium can be stabilized by either three or four neutrons. Because these different numbers of neutrons weigh different amounts, these two types of atoms are isotopes of the same element.

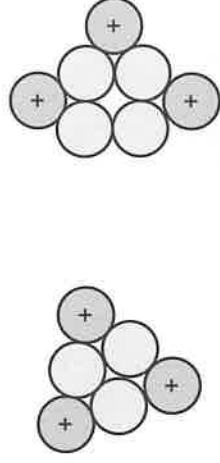


Figure 2.6: *Either three or four neutrons can separate the protons in lithium so that they don't fly apart.*

A, Z, and X Notation

All elements have different isotopes, so we need a system of symbols to tell them apart.

As a result, we have the A, Z, X notation that follows:

A_ZX

In this system, Z stands for the *atomic number* (the number of protons in the atom), A stands for the *atomic mass* (the number of protons plus the number of neutrons), and X stands for the *atomic symbol*, which denotes the element on the periodic table.



DEFINITION

The **average atomic mass** of an element is the weighted average of the masses of all its isotopes. Because different samples of an element might have different proportions of isotopes, the value given for average atomic mass on the periodic table is determined by the estimated isotopic ratios of each element in the earth's crust and atmosphere.

The following equation determines the average atomic mass of an element:

$$\text{Average atomic mass} = (\text{Mass of isotope 1})(\text{Abundance of isotope 1}) + (\text{Mass of isotope 2})(\text{Abundance of isotope 2}) + \dots$$

Now consider a question. What's the average atomic mass of lithium? Naturally occurring lithium contains two isotopes in the following abundances:

Isotope	Isotopic Mass (amu)	Abundance (%)
${}^6\text{Li}$	6.015	7.5
${}^7\text{Li}$	7.016	92.5

To solve this problem, plug the masses and abundances of each isotope into the equation for average atomic mass:

$$\begin{aligned}\text{Average atomic mass} &= (6.015 \text{ amu})(0.075) + (7.016 \text{ amu})(0.925) \\ &= 0.45 \text{ amu} + 6.49 \text{ amu} \\ &= 6.94 \text{ amu}\end{aligned}$$



BAD REACTIONS

When inserting the isotopic abundances into the equation for average atomic mass, convert the percents into decimals by dividing by 100. Otherwise, you'll find that your average atomic masses are very, very large!



YOU'VE GOT PROBLEMS

Problem 2: Find the average atomic mass of boron using the table that follows:

Isotope	Isotopic Mass (amu)	Abundance (%)
${}^{10}\text{B}$	10.013	19.9
${}^{11}\text{B}$	11.009	80.1

The Modern Atom

In This Chapter

- The Bohr planetary model of the atom
- What spectroscopy is and how it's used
- How quantum mechanics improved on the Bohr Model
- How to write electron configurations and orbital filling diagrams

In Chapter 2, you learned that many of the great minds in science have struggled to understand the atom. You also learned that those minds came up with interesting (but wrong) models to explain how the atom works. If these guys were so smart, why did they keep screwing up?

As you'll see in this chapter, one of the problems early atomic theoreticians had was that they thought of the atom in terms they could understand. For example, the Greeks thought in terms of building materials, so they ended up with tiny solid blocks. Though John Dalton described how the atom behaved, he never really changed the basic view of what the atom was. J. J. Thomson liked his desserts and came up with a model of the atom that allowed him to dream of feasts to come. These models were intuitive to the people who came up with them based on the information they had at the time.

However, the main reason these models didn't work out so well was that people assumed that small particles must behave in the same ways as much larger objects. At the beginning of the twentieth century, people began to realize that this wasn't necessarily the case. As a result, they devised some unusual models of the atom.

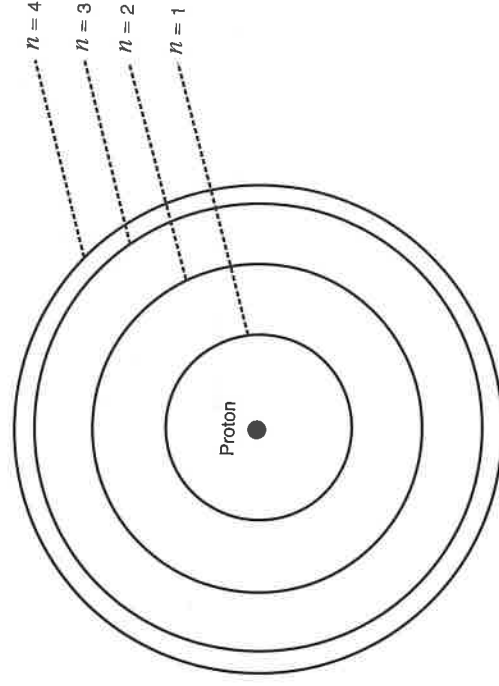


Figure 3.2: Bohr believed that electrons rotated around the nucleus in the same way that planets revolve around the sun.

That wasn't all. In addition, Bohr believed that the farther from the nucleus the electrons circled, the more energy they had. According to this model, the electrons couldn't have just any random energy. Instead, electrons were permitted to circle the nucleus at only certain distances, which corresponded to different energies. These orbits were denoted by the variable n , with the nearest orbital to the nucleus having the value of $n = 1$, the second nearest having a value of $n = 2$, and so on.

In Bohr's model, when energy is added to an atom, electrons move from orbitals close to the nucleus to orbitals that are farther away. The starting orbital of lower energy is called the *ground state*, and the orbitals with higher energy are called *excited states*. However, much like a child who's eaten 10 kg of chocolate on Halloween, eventually the electrons fall back from their excited states and re-enter the ground state. Now, the energy that first excited the electrons has to go *somewhere* when the electron falls back down. After all, it can't just vanish into thin air. When the electron returns to the ground state, this energy is given off as light. The energy of the light given off is equal to the difference in energy between the ground-state orbital and the excited-state orbital. Because electrons can jump into several different excited states, several different energies of light are emitted during this process. Bohr was a happy guy when he saw that the energies of light predicted by his equation matched the energies of light given off in the hydrogen emission spectrum. He was even happier when this discovery won him the 1922 Nobel Prize for Physics.

**DEFINITION**

Spectroscopy is a method of identifying unknown substances from their spectra. Because all elements and compounds have unique spectra, you can think of these spectra as “chemical fingerprints” specific to each material.

Many different types of spectroscopy work under the same general principle, although not all of them correspond to the movement of electrons from one energy state to another. Some of the most important varieties of spectroscopy include these:

- **UV-vis spectroscopy**—Spectroscopy that uses ultraviolet and visible light to observe how electrons jump from one orbital to another.
- **Infrared (IR) spectroscopy**—A form of spectroscopy that measures the vibrations of chemical bonds in chemical compounds using infrared light.
- **Nuclear magnetic resonance (NMR) spectroscopy**—A method of using radio waves to induce nuclear spin, though this only works for some elements.

A Quantum Leap into Quantum Mechanics

Unfortunately for Bohr, his model didn’t properly explain how atoms behave. Fortunately, quantum mechanics came to the rescue!

What Do Orbitals Look Like?

The Bohr model was definitely a step in the right direction. After all, if a model can correctly predict the orbital energies of hydrogen, there must be *something* right about it! Unfortunately, it had a problem. It was unable to predict the orbital energies of any other element. As a result, the hunt to come up with the *real* model of the atom was on. After a whole lot of work by a whole lot of famous guys you’ve probably never heard of, a new model of the atom was born. This new model, called quantum mechanics, sums up our current understanding of how atoms work.

The big problem with Bohr’s model was that there weren’t enough variables in his equation to do a good job of predicting the orbital energies of elements other than hydrogen. Eventually, the Schrödinger equation was written to explain these orbital energies. Because the Schrödinger equation is complicated, we save that for another book.

When the Schrödinger equation was put together, a strange thing happened. Whereas previous models of the atom predicted that electrons with certain energies were particles that zoomed around in predictable paths, the Schrödinger equation predicted that these

- **The angular momentum quantum number, l** —This quantum number determines the shape and type of the orbital. Possible values for l are 0, 1, 2, and so on, up to $(n - 1)$. For example, if $n = 2$, the possible values for l are 0 and 1. A spherical s-orbital is defined as having an l value of 0, a dumbbell-shaped p-orbital has $l = 1$, an oddly shaped d-orbital has $l = 2$, and a weird f-orbital has $l = 3$.

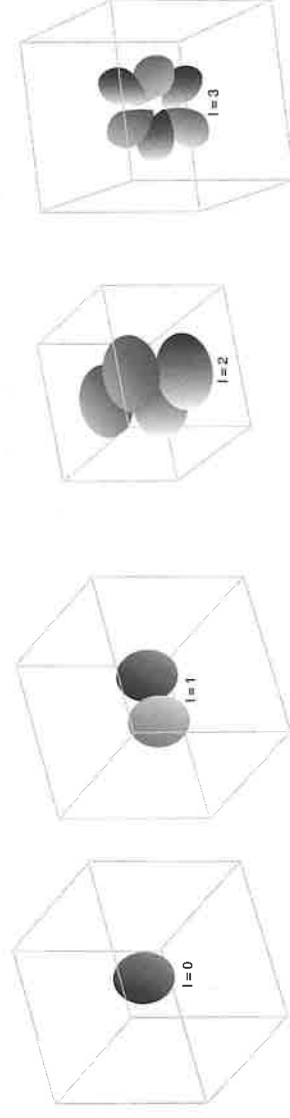


Figure 3.4: The value of l determines the shape and type of orbital being described.

- **The magnetic quantum number, m_l** —The magnetic quantum number determines the direction that the orbital points in space. Possible values for m_l are all the integers from $-l$ through l . For example, p-orbitals (which are denoted by $l = 1$) can point in three possible directions, denoted by $m_l = -1, 0$, and 1 . These three directions lie along the x -, y -, and z -axes:

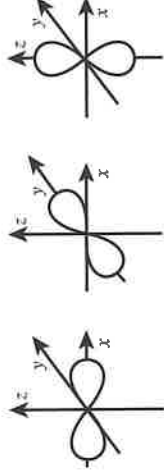
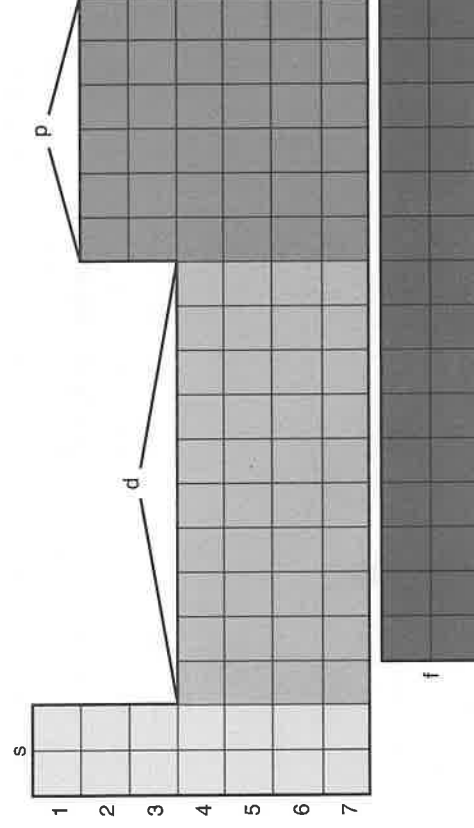


Figure 3.5: These three p-orbitals are referred to as *degenerate*, meaning they have identical energies.

- Similarly, there's one s-orbital per energy level—because $l = 0$, the only value for $m_l = 0$. There are also five d-orbitals ($l = 2$, so $m_l = -2, -1, 0, 1, 2$) and seven f-orbitals ($l = 3$, so $m_l = -3, -2, -1, 0, 1, 2, 3$) per energy level.
- **The spin quantum number, m_s** —Possible values for m_s are $+1/2$ and $-1/2$. The reason we need a spin quantum number comes from the *Pauli exclusion principle*, which states that no two electrons can have the same set of quantum numbers. If we had only the first three quantum numbers, the Pauli exclusion principle



For the purposes of determining electron configurations, we move helium so that it's next to hydrogen, as shown.

Figure 3.6: *Counting through the periodic table to the element you're looking for makes electron configurations much easier!*

Let's look at some examples:

- Hydrogen has only one electron. As you can see from hydrogen's position in the periodic table, hydrogen is in the first row of the periodic table (which means that its electron is in the lowest-lying energy state, $n = 1$), lies in the s- section of the periodic table (which means that the electron sits in an s-orbital), and has one electron in that orbital. Thus, the electron configuration for hydrogen is $1s^1$, where the first 1 stands for the principal quantum number, the s stands for the type of orbital, and the superscript 1 after the s stands for the number of electrons in that orbital.
- Helium has two electrons. Because helium is also in the first row of the periodic table, n is still equal to 1. Helium is in the s-region of the periodic table, meaning that both its electrons are in the $1s$ orbital. As a result, helium's electron configuration is $1s^2$.



THE MOLE SAYS

You might notice that helium is next to hydrogen in Figure 3.6, whereas it's way to the right of the periodic table in the tear card at the front of the book. This is because the tear card lists helium in terms of its properties, whereas Figure 3.6 lists it in terms of where the electrons are.

Example: The electron configuration of phosphorus (P) can be written in shorthand as $\text{[Ar]}4s^23p^3$. Isn't that nicer than writing out all those other terms? Similarly, you can write something as obnoxious as plutonium (Pu) as $[\text{Rn}]7s^25f^6$, a big improvement over the old method.



YOU'VE GOT PROBLEMS

Problem 3: Write the abbreviated electron configurations for the elements yttrium (Y) and polonium (Po).

Orbital Filling Diagrams and Hund's Rule

When I get on the bus, I don't like to sit next to other people. Though somebody might look okay when I first sit next to him or her, invariably the person starts picking his or her nose or coughing uncontrollably or eating a meatball sandwich as soon as I get comfortable. Over the years, I've learned that it's much easier to head for the empty seats if any are available. I've noticed that I'm not the only person to act this way—most people who aren't creepy do the same thing.

Electrons work in the same way. Let's consider the case of carbon, which has the electron configuration of $1s^22s^22p^2$. Because there's only one s-orbital per energy level, the electrons in the 1s and 2s orbitals are stuck pairing up with each other. However, because there are three p-orbitals per energy level, the two electrons in the 2p orbitals don't have to stick together—they can spread out into their own orbitals. If you were to sketch this, with low-energy orbitals farther down on the diagram than high-energy orbitals, you'd see the following:

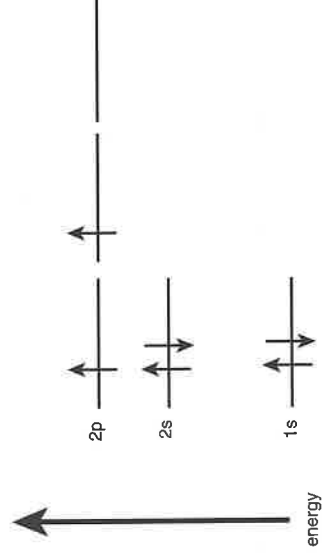


Figure 3.7: The orbital filling diagram for carbon.

A Matter of Organization

Part

2

In this part, I teach you how to cheat your way through chemistry. No, I don't tell you which parts of the body are particularly good at absorbing ink. Instead, I teach you tricks that can help you figure out what chemicals are likely to do.

The best cheat sheet ever invented is the periodic table. Unlike those cheat sheets that are very small and need to be hidden from your teachers, the periodic table can sit on your desk in plain sight and there's nothing your teacher can do to keep you from using it!

After you've mastered the periodic table, you'll learn about how ionic and covalent compounds work. Finally, you'll explore the concept of the mole—don't worry, this one is easy, too!

Elements, Compounds, and Mixtures

Chapter

4

In This Chapter

- Elements, compounds, and mixtures
- The development of the modern periodic table
- The properties of metals, nonmetals, and metalloids
- The main groups of elements in the periodic table
- The octet rule
- Periodic trends

Let's review what you've learned about chemistry. In the previous chapters, you learned what a single atom looks like. This is nothing to sneeze at, but clearly you still have quite a way to go before you can call yourself a master chemist.

In this chapter, you're going to investigate the mysteries of elements, compounds, and mixtures. You'll also learn about the magic and wonders of the periodic table. Though you're probably thinking to yourself that you've already learned all this stuff, you might find that there's more here than meets the eye!

Pure Substances

If you're a salesperson, you know that using the word *pure* is a great way to sell something. When consumers think of the word *pure*, they imagine something that's perfect in every way. Have cockroaches been crawling in the potato salad? Of course not—it's pure!

To a chemist, the word *pure* has a different meaning. When we say that something is pure, we mean only that there's one substance present in the material and that it's completely uniform in composition. Things that have completely uniform compositions are said to be homogeneous.

Shake It Up: Mixtures

Mixtures are materials that contain more than one type of pure element or compound. For example, even if cockroaches haven't been crawling around on it, potato salad is referred to as a mixture because the potatoes are made of different chemical compounds than the mayonnaise. Likewise, salt water is a mixture because it consists of pure sodium chloride (salt) mixed with pure water.

You've already seen that elements and compounds are homogeneous materials because they have a completely uniform composition. Some mixtures can also be said to be homogeneous because they contain two or more pure substances mixed together in a uniform fashion. These mixtures are called, straightforwardly enough, *homogeneous mixtures*.



DEFINITION

A **homogeneous mixture** is a mixture with completely uniform composition, whereas a **heterogeneous mixture** contains unevenly mixed components.

On the other hand, *heterogeneous mixtures* are mixtures in which the components aren't completely uniformly mixed. For example, I had a burrito for lunch today, and when I bit into it, the cheese stuck to my face while a blob of beans fell into my lap. Because I was able to separate the burrito into several distinct parts, the burrito is said to be a heterogeneous mixture.

It's usually easy to tell the difference between a homogeneous and heterogeneous mixture by looking at it. If a mixture appears to the eye to contain several different things, it's probably heterogeneous. That's why it's possible for us to guess that air is a homogeneous mixture and singer Jello Biafra is a heterogeneous mixture.



YOU'VE GOT PROBLEMS

Problem 1: Identify each of the following as being either a homogeneous mixture or a heterogeneous mixture:

- a) Turkey stuffing
- b) Sugar water
- c) Chunky peanut butter
- d) Vinegar

Separating Mixtures

A variety of methods have been devised to separate mixtures back into their components. Let's take a look at some of the most common:

- **Filtration**—If one component in a mixture is a solid and another is a liquid, a filter can separate the two. One filtration process I rely on each day occurs in a coffeemaker, where coffee grounds are separated from the coffee using a paper filter.
- **Distillation**—When one compound is dissolved in another, the most commonly used method to separate them is distillation. In distillation, the mixture is slowly heated until the component with the lower boiling point boils. The vapor from this compound can then be collected, isolating it from the other compounds in the mixture.



CHEMISTRIVIA

Distillation is a process used in manufacturing distilled spirits. Scotch whiskey originally doesn't have a high quantity of alcohol, but distilling it increases the alcohol content by a factor of three. By the way, if you're tempted to try this at home, don't. Distillation requires specialized equipment to prevent fires from taking place!

- **Chromatography**—If you've ever gotten black ink on your shirt, you've probably already performed a crude form of chromatography. You see, black ink is a mixture of several different colors of ink, and when you try to wash it from a shirt, some of the colors are more easily washed out than others. Because some of these colors stick better to shirts than others, a washed black spot on your shirt usually ends up as a lighter blue spot. In a similar way, the relative "stickiness" of other compounds can be used to separate them in a laboratory, although we usually use silica gel and organic solvents, rather than shirts and washing machines, to separate them.

The Modern Periodic Table

You've already seen the periodic table and probably aren't all that amazed by it. However, you're going to see that the periodic table is more than a seriously cool invention—it can also give you the answers to many of the questions that chemistry teachers love to ask. Let's see how.

The main properties of nonmetals, metals, and metalloids are shown in the following table:

Property	Metals	Nonmetals	Metalloids
Luster	Shiny	Not shiny	Varies
Hardness	Hard	Brittle	Hard/brittle
Bendiness	Yes	No	No
Conductivity of heat and electricity	Yes	No	Sometimes
Usual state	Solid	Varies	Solid

We explore more about why each type of substance has these properties in Chapter 8.

Periodic Families

As mentioned, elements in the same family of the periodic table have similar properties. Some of the most important families are the following:

- **Group 18, noble gases**—Noble gases are almost entirely unreactive. Completely filled s- and p-orbitals makes them stable. As a result, only a few noble gas compounds are known. Noble gases are commonly used in advertising signs, toy balloons, and blimps, and as inert atmospheres in locations where chemical reactions are undesirable.
- **Group 1 (except for hydrogen), alkali metals**—Alkali metals are highly reactive, combining readily with air and water. Although they are metallic, their densities are low (only rubidium and cesium are denser than water) and they are soft enough to be cut with a knife. The high reactivity of the alkali metals comes from the fact that they have only one more electron than the stable noble gases. As a result, they react vigorously in an attempt to lose this extra electron. Alkali metals are found in sodium vapor fog lamps and in the psychiatric drug lithium carbonate.
- **Group 2, alkaline earth metals**—The alkaline earth metals have many of the same properties as the alkali metals, but in a less extreme form. For example, most alkaline earth metals react with air and water, but they do so much less violently than the alkali metals. Alkaline earth metals are generally harder than the alkali metals but still softer than many other metals. The diminished reactivity of the alkaline earth metals can be explained by their electron configurations.

The Four Main Periodic Trends

One of the ideas that keeps coming up in this chapter is the octet rule. Because it's really, really important, let's make it fancy.



Figure 4.1: *The octet rule is the driving force for chemical reactions and properties. If you learn only one thing from this chapter, learn the octet rule!*

All elements follow the octet rule because having completely filled s- and p-orbitals in the outermost energy level makes elements stable. As you saw when we were talking about the properties of the alkali metals and halogens, elements that are only one electron away from a noble gas electron configuration are reactive.

By analogy, imagine taking a long car trip with a small, shrill child. If it's six o'clock and you see a sign that says, "Stop at Burger World, only 95 miles ahead!" you can be sure that the child will ask for a burger. When you tell the kid that Burger World is an hour and a half ahead, the kid will whine a moment and settle back down to pulling his sister's hair. Although he's hungry, he knows he'll have to wait. This is similar to how nitrogen feels about electrons—though he wants to gain three electrons to have the same electron configuration as neon, it can wait.



BAD REACTIONS

Chemistry teachers frequently get annoyed when students say, "Nitrogen wants to ..." in reference to chemical properties. However, your teacher might (correctly) tell you that atoms don't have any particular desires and that to say that they want to do something is wrong. During the course of this book, when I say that an atom "wants to do [something]" what I really mean is that "the atom will become more stable when it does [something]."

Now let's imagine that you drive a while and the sign now says "Burger World, next exit!" If you're unfortunate enough to have a halfway-literate child, the screaming from the back seat will rise to a fevered pitch until hamburgers fly into the child's mouth. This is roughly how fluorine feels about electrons—because it's so close to having a noble gas configuration, it reacts violently so that it can get that extra electron *now!*

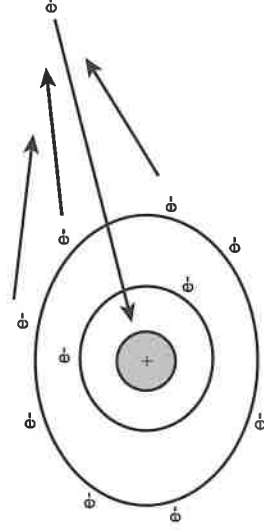


Figure 4.3: Inner electrons tend to push outer electrons away from the nucleus because both have negative charge. This results in a net decrease of attraction between an atomic nucleus and the atom's outer electrons.



DEFINITION

The **shielding effect** is the tendency for electrons close to the nucleus to repel electrons that are farther from the nucleus. This repulsion causes the outer electrons to be less tightly bound to the atom than inner electrons.

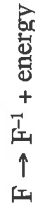
Outer electrons are strongly attracted to the nucleus because electrons and nuclei have opposite charges. Simultaneously, the outer electrons are pushed outward due to repulsion by electrons within the inner energy levels. As a result of this shielding effect, outer electrons are less tightly bound to the nucleus than inner electrons.



CHEMISTRIVIA

With enough energy, it's possible to pull more than one electron from an atom. The amount of energy needed to pull off the first electron is thus more correctly called the *first ionization energy*. Likewise, the amount of energy needed to pull off the second electron is called the *second ionization energy*, and so forth for subsequent electrons.

- **Electron affinity**—The energy change that occurs when a gaseous atom picks up an extra electron, in kilojoules per mole. This gain of an electron occurs in the following way for fluorine:



- **Atomic radius**—The atomic radius of an atom is a measure of how big it is.

Now, you might think that it's easy to figure out how big an atom is. Just measure the distance from the nucleus to the outside of the atom and you're in good shape, right?

Wrong! Quantum mechanics doesn't actually set a limit on the size of an atom, so atoms are, theoretically speaking, infinitely large. As a result, you can't just measure them like little beach balls.

So how do you find the radius of an atom? One of the most common ways is to examine two atoms of the same element that have bonded to one another. The midpoint between the two atoms is where you declare that one of the atoms stops and the other one starts. Our definition for atomic radius, then, is that it is one half the distance between the nuclei of two bonded atoms of the same element.

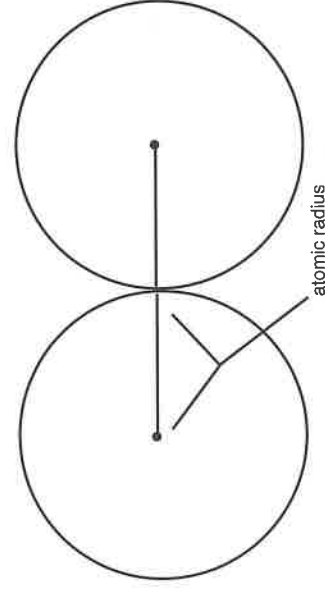


Figure 4.4: Because both of the atoms are of the same element, we say that the halfway point between them is where one atom stops and the other starts.

As you move from left to right across the periodic table, the atomic radii of the elements decrease. This might seem counterintuitive, because every element has one more electron than the one before it. Shouldn't a larger number of electrons cause the atom to be larger?

Not necessarily. Remember that the atoms within a period all have outer electrons with roughly the same energies. However, although we're adding electrons as we go along a row, we're also adding protons, because every element has one more proton than the one before it. This leaves us with a situation in which the electrons of elements on the right side of the periodic table have similar energies to those on the left, but the pull of the additional protons causes the nucleus to pull them more tightly to the nucleus. As a result of this increased nuclear

Ionic Compounds

In This Chapter

- How ionic compounds are formed
- The properties of ionic compounds
- Ionic compound naming and formula writing

It might have crossed your mind that you're four chapters into a book about chemistry without actually talking about chemical compounds. Sure, I defined what a compound was in Chapter 4, but I never really explained it further than that.

As you might have guessed, chemical compounds are far more numerous and common than pure elements. The coffee you drink is a mixture of many different chemical compounds, as is the hamburger you had for lunch and the peanut brittle you had for a snack last night while watching monster movies at 3 A.M. Okay, that last one was what *I* ate yesterday morning at 3 A.M., but you get the idea.

Unlike with elements, there's no periodic table of the compounds to make them easier to work with. However, with a little bit of practice, you'll find that the regular periodic table of elements tells you just about everything you need to know about chemical compounds. In this chapter, we talk about ionic compounds.

What's an Ionic Compound?

In Chapter 4, we talked about the octet rule, which states that all elements gain or lose electrons so that they wind up with the same electron configuration as the nearest noble gas. Basically, this means that a neutral atom of any element other than a noble gas isn't entirely stable. As a result, it will gain or lose electrons until it attains the stable electron configuration of a noble gas. Atoms that gain electrons are called *anions* and have negative

How Ionic Compounds Are Formed

As you might expect from the previous section, the octet rule plays a huge role in ionic compound formation. Let's see what happens when lithium reacts with chlorine to form an ionic compound.

Lithium has a low electronegativity because it wants to lose electrons to become like a noble gas rather than gain them. Chlorine, on the other hand, has a high electronegativity because it wants to gain electrons to become like its nearest noble gas. As a result, lithium gives electrons to chlorine when one atom of each comes into contact with the other.

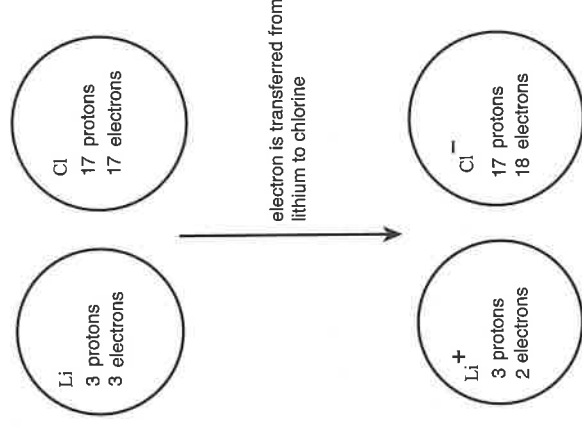


Figure 5.1: Chlorine's high electronegativity causes it to pull electrons from lithium, resulting in the formation of the ionic compound LiCl.

When this transfer of the electron occurs, lithium goes from having no charge to having a +1 charge, whereas the gain of an electron gives chlorine a -1 charge. Because the lithium cation and chlorine anion have opposite charges, they attract each other and form lithium chloride, LiCl.



BAD REACTIONS

This electrostatic attraction between cations and anions is sometimes erroneously referred to as an "ionic bond." Even though ionic attractions do represent a strong attractive force, the term "chemical bond" implies shared electrons, which simply aren't present in this case.

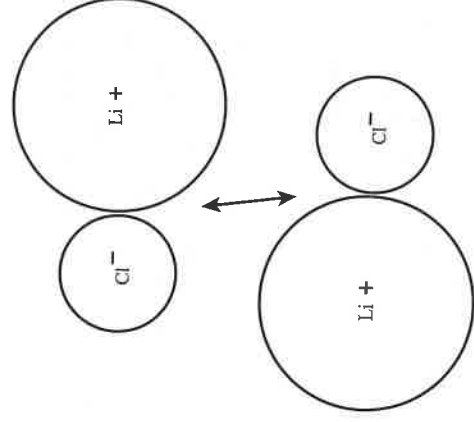


Figure 5.2: The positive charges of the lithium cations in LiCl tend to attract the negative charges of the chloride ions, causing them to stick together in large groups.

These large arrangements of ions are referred to as *crystals*. Though most ionic compounds form crystals, other nonionic compounds, such as the silicon dioxide in sand, can also form crystals. We talk about this in much greater detail in Chapter 9.



DEFINITION

Crystals are large arrangements of ions or atoms that are stacked in regular patterns.

Ionic Compounds Often Have High Melting and Boiling Points

What happens when you heat something in your kitchen? You might have discovered while cooking (or microwaving random things while bored) that most of the foods you eat either melt or burn when heated. Some foods even do both! As you can probably guess, I'm an expert when it comes to putting out house fires.

Ionic compounds frequently melt and boil at much higher temperatures than other materials. For ionic compounds to melt, enough energy must be added to make the cations and anions move away from each other. Because these attractions are so strong, it takes a lot of energy to pull these ions apart. Adding so much energy requires a great deal of heat, which explains the high melting and boiling points.

Let's say, however, that you just want to break an ionic crystal by hitting it really hard with a blunt object. Although this requires a lot of force, adding enough energy to the crystal can cause the ions to shift in relation to one another, misaligning them and causing the crystal to shatter.

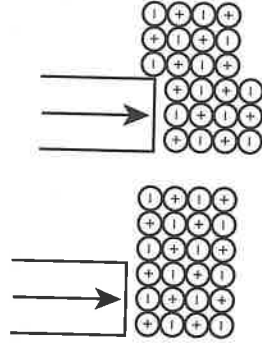


Figure 5.4: When enough blunt force is applied, the ions no longer line up so that positive and negative charges alternate. This misalignment causes ions with the same charge to sit next to each other, destabilizing the crystal and causing it to break.

Ionic Compounds Conduct Electricity Only When Dissolved in Water or Melted

Once upon a time, an inventor came up with a device for drying hair. This hair dryer, as he called it, heated air with electricity and blew it across the hair of the person holding it. Because water evaporates when heated, the hair dried more quickly.

Shortly afterward, a guy decided that he didn't want to wait to get out of the bathtub before drying his hair. His legacy: a hair dryer warning sticker with a picture of a guy getting electrocuted.

When ionic compounds are placed in water, they cause the water to conduct electricity. Normally, water doesn't conduct electricity well. However, when salts dissolve in the presence of water (something that occurs in the reservoirs that feed our water system), they break up into their constituent cations and anions. The presence of these ions enables the water to conduct electricity. Because salts conduct electricity when dissolved in water, they are referred to as *electrolytes*.



DEFINITION

Electrolytes are compounds that conduct electricity when dissolved in water. Many ionic compounds are considered to be electrolytes. However, some ionic compounds don't dissolve in water, so they don't share this property.



THE MOLE SAYS

It's not much fun, but I strongly recommend that you make some flash cards and memorize the names and formulas of the common polyatomic ions. They're used so often in chemistry that you simply have to commit them to memory.

Writing Ionic Names from Formulas

Now that you're familiar with polyatomic ions, let's see how to name ionic compounds when given their chemical formulas.

1. Determine the base name of the compound. Ionic compound base names contain two words:
 - The first word is the name of the cation. Unless the cation is ammonium (in which case you already know its name), the name of the cation is the same as the name of the element. For example, the first word in NaOH is *sodium*.
 - The second word is the name of the anion. If the anion is a polyatomic ion, just search your brain (or the table of polyatomic ions) to find its name. For example, NaOH is sodium hydroxide. If the anion is a single element, replace the ending of the element name with “-ide.” For example, NaBr is sodium bromide.
2. Determine whether the compound requires a Roman numeral.

For many compounds, you can stop with the base name. However, some elements that form cations can have more than one possible charge. For example, iron can form two ionic compounds with chlorine: FeCl₂ and FeCl₃. Because the naming system you just learned calls both of these compounds iron chloride, you need some way to distinguish between them. To do this, you write a Roman numeral after the name of the cation to indicate its positive charge.



BAD REACTIONS

Use Roman numerals only when naming ionic compounds that have more than one possible positive charge. If you place Roman numerals in all compound names, they will be wrong if they're not needed.

Unfortunately, you can't just go around adding Roman numerals to the name of every ionic compound. You do this only for compounds containing cations that can have more than one possible charge. To figure out when you need a Roman numeral, take a look at the following figure, which shows the positive charges of the most common transition metal cations.

FeCl_3 contains three chloride ions with a -1 charge. Because there's only one iron atom, the Roman numeral needed for FeCl_3 is $\frac{-(-1)(3)}{1} = 3$, giving this compound the name iron (III) chloride.



YOU'VE GOT PROBLEMS

Problem 3: Name the following ionic compounds:

- a) Na_2CO_3
- b) Cu_2O
- c) CoCO_3
- d) NH_4Cl
- e) CdSO_4
- f) $\text{Fe}_3(\text{PO}_4)_2$

Writing Ionic Formulas from Names

As you might have guessed, writing formulas from names is pretty much the reverse of writing names from formulas. Let's learn the steps:

1. From the base name, determine the formula and charge of the ions.

Let's say that you were told to write the formula of calcium sulfate. From the name calcium, you know that the cation will be Ca^{+2} . The "Ca" part is simply the atomic symbol for calcium, and the "+2" is derived from the octet rule, because calcium needs to lose two electrons to achieve the same electron configuration as argon.

Because sulfate isn't an element on the periodic table, many people start screaming in panic and confusion. If an unfamiliar ion shows up, take a look at the chart of polyatomic ions and see whether the ion is listed there. As it turns out, the word *sulfate* refers to the SO_4^{-2} ion. Feel better?

2. Devise an ionic formula that gives this compound a neutral charge.

In our example, the charges on the calcium cation and the sulfate ion cancel each other out exactly. As a result, the compound will be electrically neutral when one calcium ion combines with one sulfate ion, forming CaSO_4 .

Getting to Know Covalent Compounds

Chapter

6

In This Chapter

- What are covalent compounds?
- How are covalent compounds formed?
- Properties of covalent compounds
- Naming covalent compounds

Now that we know everything (well, almost everything) about ionic compounds, we're ready to broaden our horizons to other parts of chemistry. It's time to spread our metaphorical wings and soar to the realm of the covalent compound!

Okay, maybe that last paragraph was a little overdone. However, the idea is pretty much true. In this chapter, we learn almost everything about covalent compounds—what they are, what their properties are, and how to name them.

What Are Covalent Compounds?

In Chapter 5, you learned that ionic compounds are formed when a strongly electronegative atom grabs an electron from an atom with low electronegativity. The reason for this is the octet rule, which states that elements want to gain or lose electrons so they have the same electron configuration as the closest noble gas.

One thing we didn't discuss, however, is what happens when two electronegative atoms react with one another. For example, both nitrogen and hydrogen want to gain electrons to be like their nearest noble gas, suggesting that neither will want to give electrons to the other. Despite this, hydrogen and nitrogen actually form a large number of chemical compounds with one another, including everybody's favorite household cleaner, ammonia (NH_3). How does this work, anyway?

From the preceding figure, we can see that both atoms have the correct number of valence electrons. The left side of this figure shows that iodine has only seven valence electrons. However, after it has bonded with hydrogen, it has eight valence electrons around it. Of course, the total number of valence electrons hasn't changed, but any shared electrons count toward the valence electrons for both atoms. Likewise, hydrogen has two valence electrons around it, making it stable.

Another example: Let's see what happens when hydrogen combines with oxygen to form water, H_2O :

Each hydrogen atom has only one valence electron. In order to get the desired two valence electrons, it needs to gain another electron. Oxygen, however, has only six valence electrons. To be like neon with its eight valence electrons, it needs to gain an additional two. The following figure illustrates how one oxygen atom and two hydrogen atoms bond to form water.

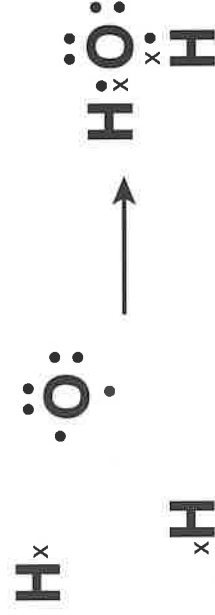


Figure 6.2: When two hydrogen atoms each share one electron with an oxygen atom, the three atoms form a chemical compound with two covalent bonds.



THE MOLE SAYS

You might have noticed that this figure initially shows the electrons on oxygen as two pairs and two single electrons, rather than as three pairs. We show electrons unpaired whenever possible because Hund's rule states that electrons prefer to remain unpaired (for more information, refer to Chapter 3).

As you can see, oxygen has two unpaired electrons that need to be paired up in order to be like neon. As in our previous example, each hydrogen atom needs one more electron to be like helium. The problem is solved when both hydrogen atoms form covalent bonds with oxygen, forming H_2O .

You might also have noticed in our two examples of covalent bonding that both iodine and oxygen have paired electrons that didn't seem to do anything at all. These electrons are referred to as "lone pairs" or "unshared pairs" because they're not involved in chemical bonding.

**BAD REACTIONS**

It's common to assume that if an atom forms a triple bond in one compound, it forms triple bonds in *every* compound. This isn't the case, so make sure you treat each example individually.

**YOU'VE GOT PROBLEMS**

Problem 2: Sketch what will happen when two atoms of oxygen combine with one atom of carbon to form carbon dioxide (CO_2).

Properties of Covalent Compounds

When we talked about solid ionic compounds in Chapter 5, we found that their properties are often derived from the strong attractions between opposite charges. It should come as no surprise to find that the properties of covalent compounds are largely due to the nature of covalent bonds.

One of the most important things to remember about covalent compounds is that they're not ionic. This seems obvious, but the difference is actually subtler than you might imagine. To illustrate this concept, take a look at the following figure.

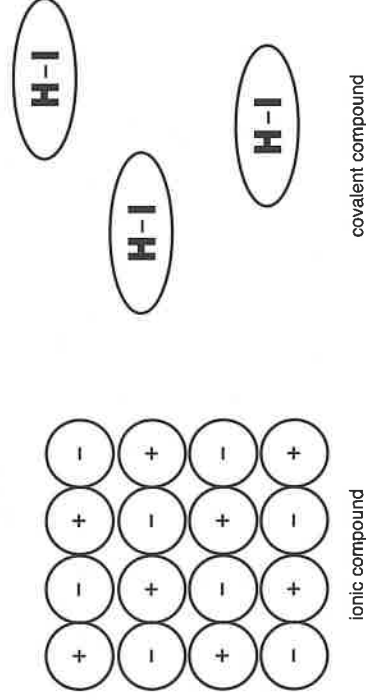


Figure 6.5: *Unlike ionic compounds, the properties of covalent compounds are based on the fact that the molecules don't strongly attract one another.*

Unlike ionic compounds, in which all the ions in a large crystal help to hold one another together, the molecules in a covalent compound are held together by forces called intermolecular forces. Intermolecular forces are much weaker than covalent chemical bonds

**DEFINITION**

Organic compounds are covalent compounds that contain carbon and hydrogen. They might also contain smaller amounts of other elements, such as sulfur, phosphorus, or any of the halogens.

It's important to keep in mind that not all covalent compounds burn. For example, water is a covalent compound, and we use it to put fires out. However, many more covalent than ionic compounds are flammable.

What's in a Name? Covalent Nomenclature

Fortunately, it's much easier to name covalent compounds than it is to name ionic compounds (see Chapter 5). Organic compounds have a separate naming system that we discuss in Chapter 23, so if you've already seen names such as benzene and 3-methylhexane, don't worry about them just yet.

Naming Covalent Compounds from Formulas

Covalent compounds have two-word names. The following rules will enable you to name nonorganic covalent compounds with the greatest of ease.

- The first word is the name of the first element in the chemical formula. For example, in CF_4 , the first word is *carbon*.
- The second word is the name of the other element in the compound, with “-ide” replacing the end of the element name. At this point, we refer to CF_4 as carbon fluoride.
- Prefixes are sometimes added to the beginning of the names of elements to indicate that more than one atom of the element is present. The most commonly used prefixes are shown in the following table:

Number of Atoms	Prefix
1	mono- (used only for <i>monoxide</i>)
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-

continues

**THE MOLE SAYS**

If you examine the pattern of the seven diatomic elements on the periodic table, you'll see that one (hydrogen) is all by itself, and the others form a big 7 at the right side of the periodic table. As a result, it's easy to remember the seven diatomic elements as "the big seven and hydrogen."

The only thing that might give you trouble are some of the elements. When most elements are named, you simply write the atomic symbol of the element. For example, carbon is just C. However, some of the elements are diatomic, meaning that they naturally occur in molecules containing two bonded atoms. These elements include the halogens (F_2 , Cl_2 , Br_2 , I_2), oxygen (O_2), nitrogen (N_2), and hydrogen (H_2). As a result, if anybody tells you that they're doing a reaction with any of these seven elements, you'll need to remember the previous formulas.

**YOU'VE GOT PROBLEMS**

Problem 4: Write the formulas of the following covalent compounds:

- a) Hydrogen bromide
- b) Oxygen dichloride
- c) Carbon tetraiodide
- d) Diphosphorus pentoxide
- e) Fluorine
- f) Diboron tetrafluoride

The Least You Need to Know

- Covalent compounds are formed when two electronegative elements are forced to share one or more pairs of electrons.
- Valence electrons are the s- and p-electrons added as the atomic number increases after the previous noble gas. They are responsible for much of the reactivity of the elements in the main block of the periodic table.
- Single covalent bonds are created when two atoms share one pair of electrons, whereas multiple covalent bonds are formed when they share more than one pair.
- The properties of covalent compounds depend strongly on the fact that covalent molecules are not chemically bonded to one another.
- Naming covalent compounds isn't difficult.

Bonding and Structure in Covalent Compounds

Chapter

7

In This Chapter

- What hybrid orbitals are and how they're formed
- How to draw Lewis structures
- Valence shell electron pair repulsion (VSEPR) theory

I've got good news and bad news for you. The good news is that you now know that covalent compounds are formed when two atoms with similar electronegativities react with one another. The bad news is that we don't yet know much about how that happens, except that electrons are somehow shared.

As it turns out, electronegativity isn't enough to explain the bewildering variety of covalent compounds that exists. It explains why we see covalent bonds, but not the shapes of the molecules that are formed nor the number of bonds they want to form.

In this chapter, we talk about the mysteries of hybrid orbitals, the valence shell electron pair repulsion (VSEPR) theory, and Lewis structures. These are not usually topics for the faint of heart, but I have full confidence that you'll get through it without much trouble.

Covalent Compounds Get Mysterious

You know from Chapter 6 that covalent compounds involve the sharing of electron pairs between electronegative atoms. But you haven't yet learned where these electrons are located. As you might imagine, they're located within orbitals, but what sort of orbitals exist between two atoms?

I'm glad you asked! Covalent bonds are formed when two orbitals from different atoms, each of which has one electron, overlap so that these two electrons are shared. Because these orbitals need to overlap for a bond to be formed, it's important to understand the shapes of orbitals that are formed in covalent compounds.

no bonding with the s-orbital (it's already full), two bonds from the half-filled p-orbitals, and no bonds from the empty p-orbital. As a result, carbon can bond only twice, a conclusion that doesn't match reality.

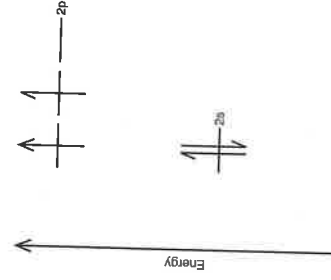


Figure 7.2: This is an orbital filling diagram for the valence electrons on carbon.

What really happens when carbon covalently bonds with other elements is that these four dissimilar s- and p-orbitals mix with one another to form four identical hybridized orbitals. The names of these new hybridized orbitals are a combination of the names of the original four atomic orbitals. In this example, one s-orbital combines with three p-orbitals to form four sp^3 orbitals:

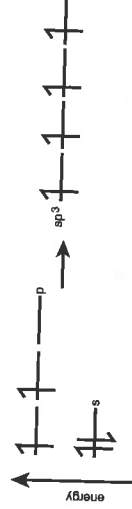


Figure 7.3: When orbitals combine to form hybrid orbitals, both their shapes and energies are averaged.

As you can see from this diagram, the hybridized orbital configuration of carbon allows room for four covalent bonds, which matches the actual structure of methane.

This mixing of orbitals has another nice effect—it allows the bond angles to be averaged as well. Because s-orbitals are spherical and p-orbitals are offset from each other by 90° angles, the hybrid sp^3 orbitals have intermediate bond angles of 109.5° . This value is exactly what's needed to keep the electrons in the bonds as far from each other as possible.

The number of hybrid orbitals that are formed when a covalent molecule bonds depends on the number of single bonds and pairs of unbonded electrons (also known as lone pairs or unshared electron pairs) that are present in the molecule. The electrons in both single bonds and unbonded pairs exist within hybrid orbitals.

hybrid orbitals. Each multiple bond is referred to as a pi bond (denoted by the Greek letter π) and is created by the overlap of unhybridized p-orbitals.

Drawing Lewis Structures

The time has come for you to learn a powerful secret. When you're done with this section, you'll not only know the secret of the covalent bond, but you'll be able to draw covalent molecules yourself. These drawings of covalent molecules, called *Lewis structures*, show all the valence electrons and atoms in a covalently bonded molecule.



DEFINITION

Lewis structures are pictures that show all the valence electrons and atoms in a covalently bonded molecule.

If you've been exposed to Lewis structures, you might have the erroneous idea that they're difficult to draw. The reason for this is simple: It's a difficult concept for teachers to explain, and books don't usually do much better. Fortunately, I have a foolproof method that can make anybody a Lewis structure mastermind.

Step 1

Count the total number of valence electrons in the molecule.

As an example, let's use carbon tetrachloride (CCl_4). The single carbon atom contains four valence electrons, and each of the four chlorine atoms contains seven valence electrons. This brings the total number of valence electrons for this molecule to $4 + (4 \times 7) = 32$.

Occasionally, you have to find the Lewis structure for a polyatomic ion. This is done by adjusting the number of valence electrons with the charge shown on the ion. For example, NH_4^+ has a positive charge, which indicates that it has one fewer electron than it should. To compensate for this, subtract one electron from the valence electron count. Likewise, CO_3^{2-} has a -2 charge that indicates two extra electrons, so you add two to your valence electron count.

Step 2

Count the total number of octet electrons in the molecule.

The number of octet electrons in a molecule is equal to the number of valence electrons that each atom will have when it has the same electron configuration as the nearest noble gas (our old friend the octet rule). The number of octet electrons that atoms want can be determined by the following rules.

- Nitrogen's family and boron bond three times in neutral molecules and can bond two, three, or four times in polyatomic ions.
- Carbon's family nearly always bonds four times.

In our example, there's only one carbon atom, so you put that in the middle of the molecule, with four chlorine atoms arranged around it (as in Figure 7.5). Between the carbon and each chlorine atom is a single chemical bond, totaling four. In this structure, both carbon and chlorine follow the rules for the number of bonds each wants. Even better, you used the number of bonds you thought you'd need from step 4.

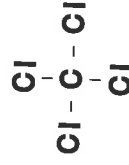


Figure 7.5: *Almost there!*



THE MOLE SAYS

If you place single bonds between all the atoms and some bonds are left over, you might need to start adding double or triple bonds. There's nothing wrong with this—just make sure that all the atoms have the correct number of bonds when you're done!

Step 6

Add lone pairs of electrons to each atom until each atom is surrounded by the number of electrons we said they wanted in step 2.

Take a look at the carbon atom in the diagram. The four bonds around it contain eight electrons. Because carbon wants eight electrons, it doesn't require lone pairs.

Each chlorine atom, on the other hand, has only one bond, for a total of two electrons. Because chlorine wants eight electrons, three pairs of electrons (totaling six electrons) need to be added to each. This gives us the Lewis structure shown here.

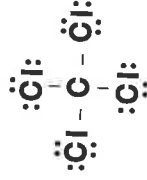


Figure 7.6: *We're done! I told you it wasn't that hard!*



Figure 7.7: These represent the three equivalent resonance structures for the nitrate ion.

These three figures are the resonance structures of the nitrate ion. When more than one valid Lewis structure can be drawn for a given formula, and all of these structures have the atoms in the same positions, they are referred to as the *resonance structures* that represent that compound. In resonance structures, although the atoms have to be in the same place, charges, lone pairs, and bonds might move around.



DEFINITION

Resonance structures occur when more than one valid Lewis structure can be drawn for a given arrangement of atoms in a covalent compound. In resonance structures, the atoms are all in the same positions, but the number and locations of the bonds and lone pairs may differ. The actual form of the molecule is an average of all possible resonance structures that can be drawn for it.

You might wonder which of the three resonance structures is the true structure of the nitrate ion. As it turns out, the actual structure of this ion is a combination of equal portions of the three. Instead of one double bond and two single bonds between the nitrogen and three oxygen atoms, imagine a situation in which there are really $1\frac{1}{3}$ bonds between each of the atoms. Likewise, instead of each oxygen atom occasionally holding a -1 charge, they all really hold a $-\frac{2}{3}$ charge. Because the concept of odd numbers of bonds and uneven charges gives most sane people a headache, we usually just draw all the possible resonance structures for a molecule and let it go at that.



THE MOLE SAYS

Resonance structures are found in compounds with double or triple bonds. If you don't have these in a compound, don't worry about resonance structures.



YOU'VE GOT PROBLEMS

Problem 2: Draw all the resonance structures for the following molecules. Remember to show the charges on each atom!

- a) CHO_2^{-1}
- b) CO_3^{-2}
- c) NO_2^{-1}

than predicted solely by the hybridization of the central atom. The following figure shows this phenomenon.

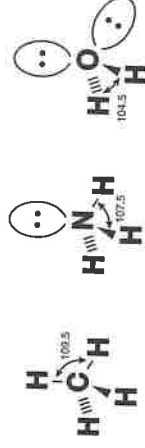


Figure 7.9: Although all three of these molecules have an sp^3 hybridized central atom, the bond angles get smaller because the increasing number of lone pairs pushes the atoms closer together.

Bond Angles and Molecular Shapes the Easy Way

So how can ordinary people like you and me remember the bond angles and shapes of all the atoms in all the covalent compounds known to man? We could memorize them, but that would cut into our valuable television time. Instead, we use the Lewis structures we covered earlier to give us a hint about how covalent molecules are put together.

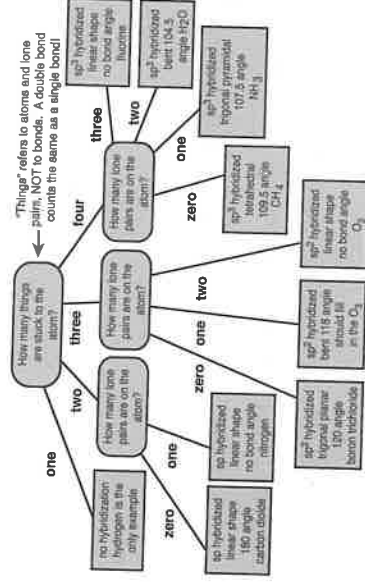


Figure 7.10: When you have a valid Lewis structure, you can use this flow chart to find the hybridization of each atom, as well as the bond angle and shape of covalent compounds.

Here's how you use this flow chart.

- At the top, you're asked, "How many things are stuck to the atom?" The word *things* refers to the number of atoms plus the number of lone pairs that are connected to the atom that you're trying to learn about (which is usually the central atom in the molecule). For PBr_3 (see the following figure), the answer is 4 because three atoms and one lone pair are stuck to phosphorus.

The Mole

Chapter

8

In This Chapter

- What's a mole?
- Molar mass
- Moles, molecules, and mass calculations
- Mass percent problems

Depending on your educational background and life experiences, different things might come to mind when you hear the word *mole* mentioned.

If you've studied the life sciences, you probably think of members of the species *Talpa europaea*, a small burrowing animal. If you work for the CIA, you probably think a mole is somebody who has infiltrated your ranks to steal state secrets. If you've spent a lot of time outside without sunscreen, you might think of a dark, raised skin blemish that sometimes grows big, weird hairs.

All these definitions are true and work well for various purposes. However, moles are none of these things in a chemistry class. In this chapter, we examine the mysterious mole and find out what it's good for.

What's a Mole?

I'm wearing a pair of shoes right now. My question for you is, "How many shoes am I wearing?" If you're thinking two, you obviously know your footwear.

Another pop quiz: I'm feeling ill because I just ate a dozen eggs. How many eggs did I eat? If you said 12, you're clearly in tune with the poultry industry.

Finding the molar masses of compounds isn't difficult if you know their formulas. To do this, multiply the numbers of atoms of each element in a compound by their atomic masses from the periodic table. When you add these numbers together, you get the molar mass of the compound.

But why does this work? Didn't I say earlier that the unit of the atomic masses on the periodic table was atomic mass units? Well, to make life easier for all of us, Avogadro's number directly converts to grams on the periodic table. For example, the average atomic weight of boron is 10.811 amu, and the molar mass of boron is 10.811 grams.



CHEMISTRIVIA

Other common terms that mean the same as *molar mass* include *molecular mass*, *molecular weight*, and *gram formula mass*.

For example, let's find the molar mass of sulfuric acid (H_2SO_4):

Element	Number of Atoms	Atomic Mass (g)	Mass \times Atoms
H	2	1.01	2.02
S	1	32.07	32.07
O	4	16.00	64.00
Total:			98.09 g

As a result, the molar mass of sulfuric acid is 98.09 g/mol.



YOU'VE GOT PROBLEMS

Problem 1: Find the molar masses of the following compounds:

- Na_2SO_4
- Nitrogen trichloride
- Fluorine
- Iron (II) phosphate

Converting Among Moles, Molecules, and Grams

When you know how to find molar masses, you can learn to convert between moles, grams, and molecules of a substance. To do so, use the following figure.

6. Do the math, making sure to cancel out units where necessary.

$$63 \text{ grams NH}_3 \times \frac{1 \text{ moles NH}_3}{17.03 \text{ grams NH}_3} = 3.7 \text{ moles NH}_3$$

In this case, you have 3.7 moles of NH_3 . However, you're trying to find the number of molecules of ammonia, not the number of moles. As a result, we need to go through the preceding six steps to find molecules. When this calculation is set up, it should look like the following:

$$3.7 \text{ mol NH}_3 \times 6.02 \times 10^{23} \text{ molecules NH}_3 = 2.2 \times 10^{24} \text{ molecules}$$

$$\frac{1 \text{ mol NH}_3}{}$$

Your answer, then, is 2.2×10^{24} molecules NH_3 .



YOU'VE GOT PROBLEMS

Problem 2:

- How many grams are in 4.3×10^{22} molecules of PF_3 ?
- How many moles are in 23 grams of CaCO_3 ?
- How many molecules are in 7.59 moles of NO_2 ?

Percent Composition

Sometimes it's handy to figure out how much of an element is present in a chemical compound. For example, let's say that you want to add extra calcium to your diet by taking a calcium supplement containing 1.00 grams of CaCO_3 . How much calcium are you actually getting?

To solve this problem, you need to figure out the percentage of calcium present in this compound. The percentage is referred to as either the mass percent or the weight percent and is found using the following formula:

$$\frac{\text{percent composition of element in a compound}}{\text{mass of the element you're interested in}} = \frac{\text{mass of the element you're interested in}}{\text{molar mass of the compound}} \times 100\%$$

Let's use this equation to find the amount of calcium in 1.00 grams of a CaCO_3 supplement.

The mass of calcium in this compound is 40.08 grams per mole because one atom of calcium is present in calcium carbonate, and calcium has an atomic mass of 40.08 grams.

Solids, Liquids, and Gases

Part 3

As you've probably already noticed, atoms and molecules are really, really small. As a result, it's hard to figure out what they're doing. Are they sitting patiently on your desk or quietly plotting to overthrow the government? You can't tell, because they're too small to see.

In Part 3, you learn about what atoms and molecules are really thinking. Whether they're uptight and rigid solids, squishy and sloshy liquids, or hyperactive and speedy gases, you figure out not only what they look like, but also what sorts of things they're likely to do.

Solids

In This Chapter

- Descriptions and definitions of solids
- Crystals and crystal structures explained
- The six main types of solids and their properties

You might have noticed that this book assumes you know what solids, liquids, and gases are. Though most chemistry books define these terms early in the first chapter, I think you already know enough to realize that if I throw something solid at your head, it will hurt. Likewise, if I throw a liquid on your head, you'll get wet, and you'll feel a mild breeze if I blow a gas at you.

However, if you learn nothing else, you should be aware that one of the big goals of chemistry is to explain why things happen on a microscopic level. Therefore, it's not enough to know that hitting you in the head with a solid rock will hurt—we want to find out why the rock has a hard surface and how the atoms in the rock will feel about their impact with your head. In this chapter, you learn about why rocks are hard, among other things.

What Are Solids?

Solids are the state of matter in which atoms or molecules are locked into place by either chemical bonds or forces between molecules called intermolecular forces. Solids are usually hard, have a shape that doesn't change, and possess a fixed volume.

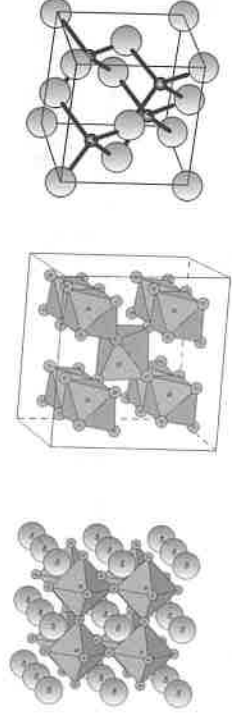


Figure 9.1: From left to right are the crystal structures of CaTiO_3 , TiO_2 , and ZnS .



DEFINITION

The term **close-packed** refers to a crystal structure in which all the atoms are as close together as possible. The two types of close-packed structures are the hexagonal close-packed (hcp) structure and the cubic close-packed (ccp) structure.

Atoms can stack together in a close-packed arrangement in two different ways. The first way is called the hexagonal close-packed (hcp) arrangement, in which the layers of atoms that make up the crystal structure alternate in an ABAB pattern, with the atoms in the third layer directly above those in the first layer. The second type of arrangement is the cubic close-packed (ccp) arrangement. In a ccp arrangement, the layers of atoms are arranged in an ABCABC pattern, enabling the atoms in the fourth layer to be located directly above the first layer. The following figure shows both types of packing arrangement.

The atoms in metals usually arrange themselves in either an hcp or a ccp arrangement. Generally, the more valence electrons a metal has, the more likely it is to have the ccp structure. Some metals take on other crystal structures, particularly at high temperatures.

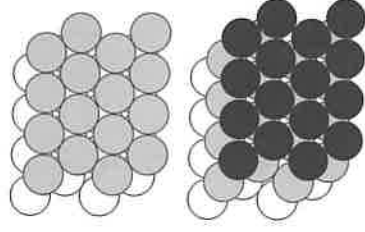


Figure 9.2: The top diagram represents the hexagonal close-packed (hcp) arrangement, in which the third row of atoms is directly above the first. The bottom diagram represents the cubic close-packed (ccp) arrangement, in which the fourth row of atoms is directly above the first.

The electron sea theory does a good job of explaining the properties of metals. Because electrons are able to move freely throughout the entire solid, metals are excellent conductors of electricity. The high mobility of electrons also causes metals to conduct heat because they do a good job of dispersing energy. Because metal nuclei can move from place to place without breaking these delocalized bonds, metals are both malleable and ductile. That's quite a theory!

Sometimes other elements are added to pure metals to give them desired properties, such as enhanced hardness, durability, strength, or corrosion resistance. The resulting mixture is called an *alloy*.



DEFINITION

An **alloy** is a metallic material in which several elements are present. The elements added to a pure metal to form an alloy are selected to maximize some desired property.

Two types of alloy exist:

- Substitutional alloys form when one of the atoms in a metal is replaced with a different element. For example, in sterling silver, some of the silver atoms are replaced with copper.
- Interstitial alloys form when smaller atoms fill some of the spaces between the atoms in a metal. One of the most important interstitial alloys is carbon steel, in which carbon atoms are located between iron atoms.

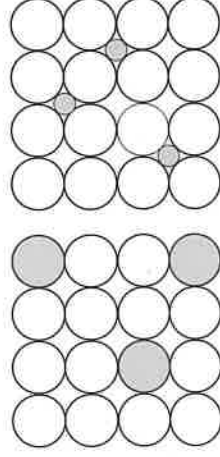


Figure 9.4: The picture on the left represents a substitutional alloy; the picture on the right represents an interstitial alloy.

Network Atomic Solids

Network atomic solids are formed when many atoms are covalently bonded to form one gigantic molecule. Unlike regular covalent molecules that are generally small, network atomic solids might grow quite large. Common examples of network atomic solids include diamonds and quartz.

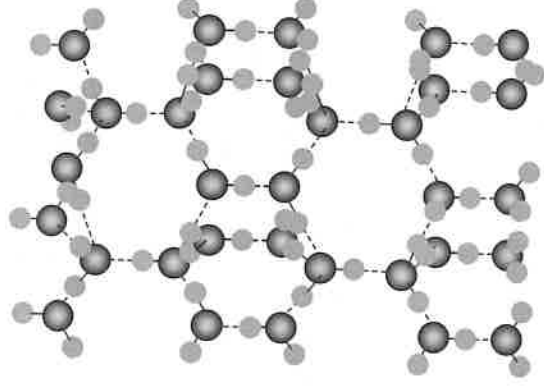


Figure 9.6: *The dotted lines in this structure correspond to the intermolecular forces holding the water molecules together in an ice crystal.*

Intermolecular forces are weaker than chemical bonds, causing the molecules in molecular solids to be less tightly bound to one another than other forms of crystal. As a result, molecular solids frequently have low melting points and are easily broken apart. Molecular solids are also extremely poor conductors of electricity. Other examples of molecular solids include sugar and dry ice.

Atomic Solids

Atomic solids form when the noble gases become cold enough to freeze. As with molecular solids, weak intermolecular forces hold these atoms together. Because their attractions are extremely weak, frozen noble gases are soft and have low melting points. For more information about the specific forces that hold atomic solids together, stay tuned for the “London Dispersion Forces” section in Chapter 10.

Amorphous Solids

Some solids don’t have a particular structure. Instead of being arranged into regular crystal lattices, the atoms bond in irregular and nonrepeating patterns. These materials are referred to as amorphous solids.

Liquids and Intermolecular Forces

Chapter

10

In This Chapter

- Properties of liquids
- The three types of intermolecular forces
- How intermolecular forces affect the properties of liquids

In Chapter 9, we discussed the properties and types of solids. It seems only fitting, then, that this chapter is about liquids. For those of you who are good at spotting patterns, you can look forward to learning about gases sometime in the near future.

If you've ever taken a chemistry class, you've probably noticed that many chemical reactions take place in the liquid phase. The reason for this is simple: liquids are easy to work with. Liquids are easy to measure, easy to pour, and easy to put in squirt guns to attack members of rival research groups. In short, liquids are the perfect medium for chemistry!

Of course, because liquids are used for a lot of things, it's important to know how they behave. That's where this chapter comes in!

What's a Liquid?

This seems like a silly question, because we all have a good feeling for what a liquid is. If I throw something on you and it pokes you in the eye, it's probably not a liquid. On the other hand, if it makes your shirt wet and sticky, it probably is.

Liquids are the state of matter in which molecules can move around freely but still experience forces that keep them near one another. For example, if you pour a glass of water on the floor, the water will tend to pool in a few spots rather than spread into a gigantic, thin puddle.

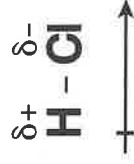


Figure 10.1: Because chlorine is more electronegative than hydrogen, the electrons are attracted to it more strongly, giving it a partial negative charge.

Let's see what the symbols in this figure mean. Chlorine is more electronegative than hydrogen, so it pulls harder on the electrons in the covalent bond than hydrogen does. As a result, the electrons in this bond spend more time hanging around the chlorine atom than the hydrogen atom. This gives chlorine a partial negative charge, denoted by the symbol (δ^-). Likewise, because electrons spend less time around the hydrogen atom, it has a partial positive charge, denoted by (δ^+). The arrow in the diagram is called a dipole arrow and points toward the side of the bond with the partial negative charge (Cl).

Because the electrons in this bond are distributed unevenly, it's referred to as being a polar covalent bond. Polar covalent bonds form whenever two elements with dissimilar electronegativities form covalent bonds.

A good way to tell whether a molecule is going to be polar is to take a look at the central atom. If any of the things (for example, lone pairs or atoms) on the central atom differ from the others, the molecule will be polar. Take a look at the Lewis structure of OF_2 to see what I mean.



Figure 10.2: Because oxygen has lone pairs and fluorine atoms stuck to it, and because a lone pair is definitely not the same thing as an oxygen atom, this molecule is polar.



THE MOLE SAYS

If a Lewis structure shows a molecule to be polar but you still can't tell where the partial positive and negative sides of the molecule are, switch atoms with lone pairs in the Lewis structure until the molecule looks asymmetrical. After you do this, the overall polarity of the molecule should be clearer.

**DEFINITION**

The term **hydrogen bond** isn't technically accurate. Though the interaction between hydrogen and highly electronegative atoms is strong for an intermolecular force, it's nowhere near as strong as a true chemical bond. However, because everybody uses this term, we're stuck using it to describe these interactions.

To understand how this happens, let's consider the example of hydrogen fluoride. In HF, we have a very polar H-F bond due to fluorine's extremely high electronegativity. As a result, most of the electrons in this bond are pulled toward fluorine, leaving little electron density around hydrogen.



Figure 10.4: In HF, fluorine pulls most of the electron density from hydrogen. Because hydrogen has no inner electrons, the partial positive charge on it is strong, leading to strong interactions with the lone pairs on the fluorine atoms from other HF molecules.

Because hydrogen has little electron density, it has a partial positive charge. However, unlike other elements with partial positive charges, hydrogen has no inner electrons to shield the nucleus from other atoms. This lack of inner electrons enables atoms with partial negative charges to have extremely strong electrostatic interactions with hydrogen. These hydrogen “bonds,” although still not as strong as covalent bonds or the attractive forces between anions and cations, are much stronger than other intermolecular forces.

London Dispersion Forces

So far, we've seen the forces that bind polar molecules in a liquid. But what forces cause the molecules in a nonpolar liquid to be attracted to one another? You might be surprised to find that nonpolar molecules also depend on the attraction of opposite charges to stay together in a liquid.

How does this process work? After all, nonpolar molecules, by definition, don't have any positive or negative charges! The following figure shows how this works when helium is liquefied.

**THE MOLE SAYS**

If you want to figure out what kind of intermolecular force is experienced by a covalent molecule, draw its Lewis structure. If it has an O-H, N-H, or F-H bond, the intermolecular force is hydrogen bonding. If it doesn't have any of these bonds but is polar, the primary intermolecular force is the dipole-dipole force. If it's completely nonpolar, look to London dispersion forces as the main intermolecular force.

**YOU'VE GOT PROBLEMS**

Problem 2: Determine the intermolecular force that's strongest in each of the following compounds:

- a) NBr_3
- b) CO_2
- c) NH_3
- d) N_2
- e) CH_2O

The Effects of Intermolecular Forces

The intermolecular forces present in a compound play a role in that compound's properties. This isn't surprising when you think about it. After all, if the molecules in one liquid are held tightly together by a strong intermolecular force, you'd expect this liquid to behave differently than a second liquid in which the molecules are attracted to one another weakly. The following are two ways in which intermolecular forces affect the properties of a liquid:

**THE MOLE SAYS**

To recap, hydrogen bonds are the strongest intermolecular force, dipole-dipole forces are of intermediate strength, and London dispersion forces are the weakest. None of these three forces is anywhere near as strong as covalent bonds or the attractions between cations and anions in ionic compounds.

- **Melting and boiling point**—Generally, compounds that undergo hydrogen bonding melt and boil at higher temperatures than compounds that experience dipole-dipole forces. Likewise, compounds that experience dipole-dipole forces have higher melting and boiling temperatures than those that experience London dispersion forces. For example, consider the following four molecules.

Solutions

Chapter

11

In This Chapter

- What's a solution?
- Why some things dissolve and other things don't
- Concentration
- Factors affecting solubility
- Dilutions

In Chapter 10, you learned a lot about liquids and their properties. However, we didn't address what happens when you dissolve something in a liquid. After all, many of the chemicals you work with are solids, and even though they're easy to measure and manipulate, they tend to react at glacial rates. By dissolving solids in liquids, you can manipulate their concentrations to ensure quick reaction rates. The resulting mixtures are referred to as solutions.

However, not all solids can dissolve in all liquids. If they did, your drinking glass would dissolve every time you poured yourself a refreshing glass of milk. In this chapter, you learn about the formation, behavior, care, and feeding of solutions.

What Are Solutions?

The word *solution* is just another fancy term for a homogeneous mixture (see Chapter 4). In solutions, one material (called the solute) is completely dissolved in another (called the solvent). Examples of solutions that I use around my house every day are fruit punch and contact lens solution, both of which contain solid solutes dissolved in water.

Ionic solids always contain cations and anions. As a result, when an ionic solid such as sodium chloride is placed in water, the following takes place:

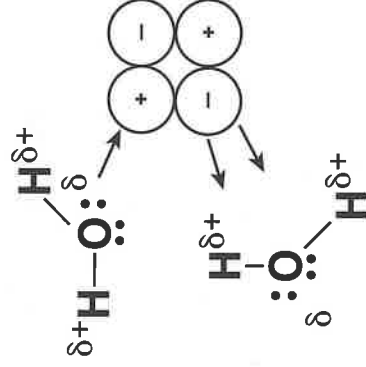


Figure 11.2: *The partial charges on water interact strongly with the ions in sodium chloride, making sodium chloride highly soluble in water.*

When sodium chloride is placed in water, the partial positive charges on the hydrogen atoms in water are attracted to the negatively charged chloride ions. Likewise, the partial negative charges on the oxygen atoms in water are attracted to the positively charged sodium ions. Because the attractions of the water molecules for the sodium and chloride ions are greater than the forces holding the crystal together, the salt dissolves.

Similarly, polar solutes such as methanol, ethanol, and isopropanol are highly soluble in water because they are also polar.



CHEMISTRIVIA

In some cases, the attraction of water molecules for the polar solute isn't enough to pull the solute molecules apart. As a result, some polar solutes (such as the calcium carbonate found in limestone) don't dissolve well in water.

Why Polar Solvents Don't Dissolve Nonpolar Solutes

The “like dissolves like” rule indicates that polar solvents do a poor job of dissolving nonpolar solutes. You can understand this by looking at the following figure.

Why Nonpolar Solvents Dissolve Nonpolar Solutes

If you place a nonpolar solid in a nonpolar solvent, “like dissolves like” implies that the solid will dissolve. However, the only forces that cause the liquid to be attracted to the solid are weak London dispersion forces. Why should the solid dissolve?

Imagine that you place a chunk of carbon tetrabromide in a beaker containing carbon tetrachloride. The carbon tetrabromide molecules in the solid are attracted to one another by weak London dispersion forces, as are the carbon tetrachloride molecules in the solvent. You might expect, then, that the solute dissolves for no particular reason.

As it turns out, another force involved pushes the solute to dissolve in the solvent. Processes that increase the randomness of a system usually tend to occur spontaneously

(you learn more about this concept, known as entropy, in Chapter 27). Because the molecules in carbon tetrabromide are made more random if they’re mixed with another compound and are floating around in solution, the carbon tetrabromide will dissolve in the carbon tetrachloride.



YOU’VE GOT PROBLEMS

Problem 1: Based on the polarity of each solvent and solute, determine whether the solvent in each of the pairs will likely dissolve the solute listed.

- Solvent: water. Solute: lithium chloride.
- Solvent: methanol (CH_3OH). Solute: HBr .
- Solvent: carbon tetrachloride (CCl_4). Solute: NH_3 .

Determining the Concentration of a Solution

You can measure the amount of solute present in a solution in many ways. Each method is useful for a different purpose in chemistry, so unfortunately, you’re stuck learning all of them. Without further ado, here they are.

Qualitative Concentrations

The amount of solute present in a solution can be described without numbers using one of the following terms:

- **Unsaturated**—A solution that is unsaturated has not yet dissolved the maximum possible amount of solute. For example, if you dissolve a teaspoon of salt into a swimming pool of water, the water in the pool is said to be unsaturated in salt because more salt can still be dissolved. The problem with this method of describing concentration comes from its lack of specificity. For example, an entire

A solution with a molarity of 1.4 is said to be a 1.4 molar solution.



YOU'VE GOT PROBLEMS

Problem 2: What is the molarity of a solution if 120 grams of acetic acid ($C_2H_3O_2H$) have been diluted to a final volume of 3,100 mL?

Molality (m)

Molality is defined as the number of moles of solute per kilogram of solvent. For example, if you add 2.5 kilograms of water to 4.5 mol of sugar, the molality equals this:

$$m = \frac{\text{moles of solute}}{\text{kg of solvent}} = \frac{4.5 \text{ mol}}{2.5 \text{ kg}} = 1.8 \text{ m}$$

A solution with a molality of 1.8 is said to have a concentration of 1.8 molal.

When doing calculations with water, keep in mind that the density of water is 1.0 g/mL under standard conditions, so the number of kilograms of water is equal to the number of liters of water.



YOU'VE GOT PROBLEMS

Problem 3: Determine the molality of a solution in which 45 grams of calcium acetate are added to 560 mL of water.

Normality (N)

The normality of a solution is defined as the number of moles of a reactive species, usually referred to as equivalents per liter of solution. The use of “equivalents” depends on the reaction performed, so some knowledge of the specific chemical process taking place is necessary before computing normality. At least, that’s the *normal* way of solving this problem (I couldn’t resist).

Mole Fraction (χ)

The mole fraction of a component in a solution is defined as the number of moles of this component divided by the total number of moles of all components in the mixture (including the solvent). In equation form, you can express the mole fraction of one

Surface Area of the Solute

Imagine that you're trying to dissolve 1.0 g of NaCl in a glass of water. Which would dissolve more quickly, a large crystal or the same mass ground into powder?

If you guessed that the powder would dissolve more quickly, you're right! Because the powder has a larger surface area than the crystal, more of the ions in the salt are exposed to the solvent at a given time, causing them to dissolve more quickly. Note that breaking a solute into smaller pieces doesn't change how much of it will dissolve—only how *quickly* it will dissolve.

Pressure

When dissolving a gas within a liquid, the pressure of the gas has a huge effect on its solubility. When the pressure of a gas is low, the number of gas molecules that hit the surface of the liquid at any given time is low. As a result, the gas has fewer chances to dissolve. However, if the pressure of the gas increases, the number of collisions between the gas molecules and the solvent increases, causing more of the gas molecules to dissolve.

This relationship between pressure and solubility is called Henry's law, which states the following:

$$P = kC$$

In this equation, P represents the pressure of the gas above the solvent, k is a mathematical constant that depends on the particular solution, and C represents the concentration of the gaseous solute in the solution. As you can see from the equation, the higher the pressure of the gas, the more concentrated the solution will be. Though pressure is an important factor in the solubility of a gas, it has little effect on the solubilities of liquids or solids.

Temperature

The temperature of a liquid affects the solubility of both solids and gases. For solids, the warmer the solvent is, the more soluble it is and the faster it dissolves (though there are exceptions). On the other hand, gases become less soluble as the temperature of the solution increases, which is why carbonated beverages (which contain CO_2) go flat more quickly on hot days than on cold ones.

The Kinetic Molecular Theory of Gases

Chapter

12

In This Chapter

- Properties of gases
- The kinetic molecular theory of gases
- Ideal gases
- Root mean square velocity
- Effusion and diffusion

Gases are more difficult to visualize than other forms of matter. If you're anything like me, you learned at an early age that solids are hard when you bashed your face on the coffee table, spraying blood all over your grandparents' house while the 14-year-old babysitter cried. You learned that liquids are wet when you were 2 and your parents let you sleep through the night in your big-boy pants instead of diapers and you responded by wetting the bed.

Gases, on the other hand, are probably something you didn't really learn about until you headed off to school. After all, you can't bash your head on air as a kid, and gases are typically both transparent and odorless. In this chapter, we investigate both the basic properties of gases and explore why gases have these properties.

What Are Gases?

Gases are the phase of matter in which particles are usually far apart from one another, move quickly, and aren't particularly attracted to one another. Because the molecules in a gas are so far apart from one another, gases are much less dense than liquids or solids. That's why it's easier to pick up a balloon full of air than it is to pick up a big water balloon.

The Kinetic Molecular Theory of Gases: Why Gases Do What They Do

As mentioned previously, gases are harder to visualize than other phases of matter. This is true not only because it's difficult to see and study them, but also because the molecules in a gas are flying around at high speed all over the place. This makes the structure of gases harder to pin down than the atoms in a nice, boring crystal.

Because it's tough to study all the particles in a gas, scientists have come up with a variety of theories to simplify gases' behavior so they can be more easily understood. Probably the most important of these is referred to as the kinetic molecular theory (KMT).



BAD REACTIONS

All theoretical models (including the KMT) only approximate the behavior of what's being modeled. The approximations that define each model are designed to make the real phenomenon easier to understand and predict. However, no model is perfect, which explains why weather-forecasting models usually get the five-day forecast wrong.

The KMT makes the following assumptions about the behavior of the particles in a gas. These assumptions are not always completely true, but they are good enough that they enable us to understand gases more easily.

The Particles in a Gas Are Infinitely Small

Atoms and molecules are really, *really* small ($\sim 10^{-10}$ m). The kinetic molecular theory not only says that atoms and molecules are really small, but goes even further and says that they have no volume at all.

We make this assumption because, like many models, the KMT is a mathematical model and this assumption makes the math easier to work out. Plus, it's *almost* true—in a sample of steam at 100°C , the water molecules make up only $1/1700$ of the total volume of the gas. Improving the model to account for this volume makes the math behind the model a lot more difficult but doesn't add much accuracy. As a result, we just assume that gas molecules are infinitely tiny.

The Particles in a Gas Are in Constant Random Motion

The KMT correctly assumes that the particles in a gas, like small children, constantly move from place to place in an unpredictable fashion. Furthermore, the KMT assumes

shorten to 273° to make our lives easier). You can easily convert degrees Celsius to Kelvin using the following equation:

$$K = ^\circ C + 273$$

For example, if the temperature outside is $20^\circ C$, the temperature in Kelvin is $293 K$. Note that Kelvin temperatures are written simply as “Kelvin,” not as “Kelvins” or “degrees Kelvin.”



THE MOLE SAYS

We use Kelvin instead of degrees Celsius when working with gases because gases often exist at temperatures less than $0^\circ C$. As a result, if we said that the kinetic energy of a gas was proportional to the temperature in degrees Celsius, the kinetic energy of gases cooled below the freezing point of water would be negative—which would obviously be nonsense because the molecules are still moving rapidly in the gas.

Gas Molecules Undergo Perfectly Elastic Collisions

Elastic collisions are collisions in which kinetic energy is transferred from one thing to another without any loss. If you’ve ever played pool, you know that the balls slow down just a little bit when they hit the sides of the table. If these collisions were perfectly elastic, the balls would bounce off the walls moving exactly as quickly as they hit them in the first place, never stopping until either they hit a pocket or you get tired of watching them bounce around the table.

Why This Is Important: Ideal Gases

From the very definition of a model, we know that the kinetic molecular theory of gases isn’t true. Instead of telling us how gases actually behave in the real world, it gives us an idealized version of how gases *should* behave under perfect conditions. Gases that follow all the assumptions of the KMT are referred to as ideal gases.



THE MOLE SAYS

If you haven’t yet picked up on this idea, here’s a clarification: There is no such thing as an ideal gas! Ideal gases are imaginary! There are as many ideal gases in the world as there are tooth fairies! Don’t tell anybody that you believe ideal gases are real, because they’ll lock you up for being delusional!

- **Standard temperature and pressure (STP)**—STP is the most common reference condition for expressing the pressure of gases. Standard temperature is defined as 0°C (273 K), and standard pressure is 1 atm .
- **The ideal gas constant (R)**—Also known as the universal gas constant, this value is handy when working with gases. Depending on what units you're working with, the values for R are either $0.08206\text{ L atm/mol K}$ or 8.314 kPa/mol K . The value you should use for a particular problem depends on the unit of pressure (kPa or atm) you're given.
- **Room temperature**—Because we like to do scientific discoveries in heated rooms, many scientists refer to “room temperature” in a calculation. Although it's not a formally accepted term, room temperature is usually understood to be 25°C (298 K).

How Fast Do Gas Molecules Move?

How fast do the molecules in a gas move? We've said several times that they go “really fast,” but how fast is “really fast”?

To answer this question, you first need to look at some of the factors that determine the speed of gas molecules.

The Mass of the Gas Molecules

The KMT says that the kinetic energy of the molecules in a gas is proportional only to the temperature in Kelvin. As a result, heavy objects and light ones have the same kinetic energies at the same temperature.

Let's say that I'm a very bad driver (which is true). During an ice storm several years ago, I drove my car into the fence surrounding my workplace at approximately 5 miles an hour.

When I destroyed the fence, my car was like a heavy gas molecule. Here's an interesting question: how fast would a bicyclist need to go to destroy the fence with the same amount of energy that my car used? If you guessed “really, really fast,” you're right! Because bicycles are much lighter than my car, they need a lot more speed to build up the same amount of kinetic energy.

Likewise, if two molecules have the same amount of kinetic energy, the lighter one will move more quickly than the heavy one. In other words, the velocity of the molecules in a gas depends on their masses!

The Random Walk

Our calculation found that ammonia molecules move 661 m/sec at room temperature. If ammonia moves this quickly, why don't you immediately smell it whenever your neighbor across the street mops his floor?

If ammonia molecules traveled straight from your neighbor's floor to your nose, you *would* smell it almost immediately. However, molecules don't travel in straight paths. Instead, they bump into each other in random fashion.

To see what I mean, imagine that the Olympic committee has decided to make the marathon more interesting by blindfolding all the runners. Though the runners will eventually finish the marathon even if blindfolded, it will probably take them days to finish the race because they won't be going in a straight path. Instead, they'll bump into trees, spectators, each other, and so on. The path that these runners will take is called a random walk because they'll travel quickly in random directions.

Molecules do the same thing. Though ammonia molecules travel 661 m/sec at room temperature, they take a long time to cross a room because they keep bumping into things and bouncing off in a random direction. The length of time it takes for molecules to travel from one place to another depends not only on their rms velocity, but also on the average distance between collisions, called the mean free path.

Effusion and Diffusion

Other important behaviors of gases explained by the kinetic molecular theory are effusion and diffusion. Effusion is the rate at which gas escapes through a small hole in a container. Diffusion is the rate at which a gas travels across a room. The following figure illustrates both of these phenomena.

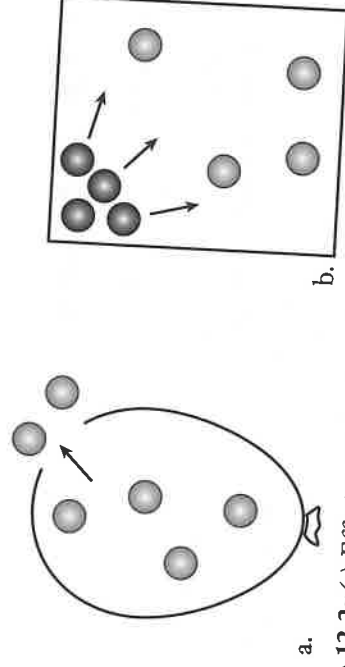


Figure 12.2: (a) Effusion occurs when gas escapes through a small opening in a container. (b) Diffusion is the rate at which a gas travels across a room, mixing with the other gases already present.

Gas Laws

In This Chapter

- Gas laws named after famous people (Boyle, Charles, Gay-Lussac, Avogadro, and Dalton)
- The combined gas law
- The ideal gas law

In Chapter 12, we spent a lot of time discussing how gases behave on a molecular level with the kinetic molecular theory (KMT). After we developed the KMT, we were able to explain the easily observed properties of gases in terms of this theory.

Unfortunately, the calculations you did in Chapter 12 don't help with most of the common problems you need to solve. For example, what happens to the pressure of a gas in a closed container when you raise the temperature from 25° to 500°? At first glance, this might not seem like an interesting problem. However, if you throw a can of spray paint into a campfire, you'll see a spectacular demonstration of why this is interesting. (By the way, don't do this—if you want to see it, check it out on YouTube.)

Of course, scientists in the fields spend relatively little time throwing compressed gases into campfires. Here you not only learn about how gases behave, but also examine some ways in which sane chemists might use these behaviors to make the world a better place.

Boyle's Law: Why Compressed Gas Is Small

Time for a demonstration. Inflate a balloon and put it on your chair. Now flop down on the chair as hard as you can, squishing the balloon.

Charles's Law: The Incredible Imploding Can

Let's do another demonstration. You need a brand-new, never-used metal can with a screw-on cap—and don't even think about doing this with a glass jar! Remove the cap, place the can on the stove, and turn the stove knob to high. After the can heats for about two minutes, *carefully* screw the top onto the can with some tongs and turn off the heat.

Because I know that none of you actually did this demonstration (shame on you!), I'll just tell you what you would have seen—over a period of several minutes, the can would shrink until the sides caved in.

Way back in 1787, French scientist Jacques Charles did the same experiment while sitting around the house on a rainy day. (*Editor's note:* Our research shows that only the year in the preceding statement is correct.) When he observed the implosion of the can, he devised the following law to explain his findings. Being a huge egomaniac, he demanded that everybody call it Charles's law:

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



BAD REACTIONS

When working with gases, remember to always convert temperatures from degrees Celsius to Kelvin ($K = ^\circ\text{C} + 273$). If you don't, your answer will be wrong!

In this law, V_1 represents the initial volume of the can, T_1 is the initial temperature of the can (in Kelvin), V_2 is the final volume of the can, and T_2 is the final temperature (in Kelvin). You can assume that the pressure and number of moles of air are constant. If the can has an initial volume of 5.00 L, the temperature of the can before you turned off the heat was 250°C (523 K), and the temperature of the can after it cooled was 25°C (298 K), you can use this equation to find the final volume of the can.

$$\frac{5.00\text{ L}}{523\text{ K}} = \frac{V_2}{298\text{ K}}$$

$$V_2 = 2.85\text{ L}$$

Why does the volume of the can decrease? At high temperatures, the gas molecules in the can are bouncing on the inside wall, exerting pressure on the wall. However, when you close the can and cool the gas, these gas molecules don't exert as much force when they hit the side of the can, so the pressure inside the can drops. This decrease in pressure causes the can to implode, leading to the new (and much smaller) volume.

**YOU'VE GOT PROBLEMS**

Problem 4: You decide to go deep sea exploring in a giant transparent balloon. If you start your voyage in a balloon with a pressure of 1.00 atm, a temperature of 20° C, and a volume of 5.00×10^3 L, what will the pressure inside the balloon be when you reach a depth at which the temperature of the balloon is 2° C and the volume of the balloon is 1,170 L?

Avogadro's Law and the Ideal Gas Law

Amadeo Avogadro was an Italian scientist way back in the first half of the nineteenth century. Though he was an odd-looking guy with a big forehead, he was a strikingly good researcher in the field of molecular theory. Among his discoveries was the concept that any two gases with identical volume, temperature, and pressure will contain the same number of molecules. This idea is now known as *Avogadro's law*.

**DEFINITION**

Avogadro's law states that any two gases with identical volume, temperature, and pressures will contain the same number of molecules.

Of course, there is one catch: this statement is true only for ideal gases. Fortunately, under the conditions of pressure, temperature, and volume that you're used to working with, real gases behave in a nearly ideal fashion.

Better yet, because all ideal gases have the same number of molecules per unit of volume under some set of conditions, you can devise a single equation to express the relationship between the number of moles of a gas and these other variables. This relationship is called the ideal gas law:

$$PV = nRT$$

P denotes pressure (in either atm or kPa), V denotes the volume of the gas in liters, n is the number of moles of gas present, R is the ideal gas constant, and T is the temperature of the gas in Kelvin.

**THE MOLE SAYS**

The ideal gas constant, R, has two values that you need to remember. For problems in which pressure is given to you in atm, R should be given as 0.08206 L atm/mol K. For problems in which pressure is given in kPa, the correct value for R is 8.314 L kPa/mol K.

Consider an example of how this works:

Example: Your refrigerator has a volume of 1,100 L. If the temperature inside the refrigerator is 3.0° C and the air pressure is 1.0 atm, how many moles of air are in your refrigerator?

Solution: To solve this problem, simply insert the appropriate values of P, V, and T into the ideal gas law. As for R, use the value 0.08206 L atm/mol K because the pressure was given to you in units of atm.

$$(1.0 \text{ atm})(1,100 \text{ L}) = n(0.08206 \text{ L atm/mol K})(276 \text{ K})$$

$$n = 49 \text{ mol}$$



YOU'VE GOT PROBLEMS

Problem 5: If your oven has a volume of 1,100 L, a temperature of 250° C, and a pressure of 1.0 atm, how many moles of gas does it hold?



CHEMISTRIVIA

The ideal gas law explains how hot air balloons work. The number of moles of air inside a hot balloon is less than the moles that are displaced from the cooler air. Because there are fewer moles of warm air than are present in the equivalent volume of cool air from outside the balloon, the mass of the air in the balloon is also less, causing the balloon to “float” above the surrounding cold air.

Dalton's Law of Partial Pressures

Let's say that, for one reason or another, you're not happy with the regular air you've been breathing your entire life. Instead of breathing that same old boring air that's floating around outside, you're interested in making custom air that fits your youthful, “extreme” personality.

To improve your air, you decide to fill your house with a supercharged mixture of 40 percent oxygen by volume, 40 percent nitrogen by volume, and 20 percent helium by volume (because squeaky voices are fun).

As it turns out, John Dalton was also interested in making his own special blend of custom air. (*Editor's note:* No, he wasn't.) He reasoned that the total pressure of the custom air in his house would be equal to the sum of the individual pressure of each gas inside

- The partial pressure of N_2 :
 $(P_{\text{nitrogen}})(48,000 \text{ L}) = (250 \text{ mol})(0.08206 \text{ L atm/mol K})(298 \text{ K})$
 $P_{\text{nitrogen}} = 0.13 \text{ atm}$
- The partial pressure of He:
 $(P_{\text{helium}})(48,000 \text{ L}) = (75 \text{ mol})(0.08206 \text{ L atm/mol K})(298 \text{ K})$
 $P_{\text{helium}} = 0.038 \text{ atm}$

Using Dalton's law, the total pressure of all the gases in this mixture is:

$$P_{\text{tot}} = P_{\text{oxygen}} + P_{\text{nitrogen}} + P_{\text{helium}}$$

$$P_{\text{tot}} = 0.076 \text{ atm} + 0.13 \text{ atm} + 0.038 \text{ atm}$$

$$P_{\text{tot}} = 0.24 \text{ atm}$$

This is roughly the same air pressure that exists at the top of Mt. Everest, which makes breathing difficult.



YOU'VE GOT PROBLEMS

Problem 6: You found a pressurized cylinder of nitrogen and oxygen in your basement. The temperature of the cylinder is 15°C , and its volume is 55 L. The partial pressure of oxygen in the cylinder is 5.0 atm, and the partial pressure of nitrogen is 8.0 atm. Given this information:

- How many moles of oxygen and nitrogen are present in the cylinder?
- What is the total pressure of gas in the cylinder?

The Least You Need to Know

- Boyle's law, Charles's law, Gay-Lussac's law, and the combined gas law all express the relationships among the volume, temperature, and pressure of a gas.
- The ideal gas law enables you to determine the relationship among the pressure, volume, temperature, and number of moles of a gas.
- Dalton's law of partial pressures is good for calculating the pressures within a mixture of gases.

Phase Diagrams and Changes of State

Chapter

14

In This Chapter

- How the vapor pressure of a liquid affects phase changes
- Colligative properties
- What happens when compounds undergo phase changes
- Phase diagrams

In the past few chapters, we discussed the properties, structures, and behaviors of solids, liquids, and gases. By now, you should be a real pro when it comes to the three states of matter.

What we haven't yet mentioned is that you can change materials from one phase to another. The most familiar example is that of water—if you start with ice, you can heat it until it melts to make liquid water. If you heat it further (past 100° C), you can make steam.

In this chapter, you wrap up your voyage into the world of the states of matter by describing how to convert between the phases of matter. By the end of this chapter, you'll not only know about each state of matter, but you'll understand how to control their every action. It's a big responsibility, but I'm sure you're up to it!

Vapor Pressure: Why Phase Changes Occur

When I was a kid, I had a pet fish. I'm sorry to say that I didn't do a good job taking care of it. The fish eventually died because the bowl went dry when I forgot to refill the water. As a budding young scientist, I wondered why the water in the fish bowl evaporated. Unfortunately, at 6 years old, I wasn't very bright—I concluded that my brother had drained the water as a mean trick.

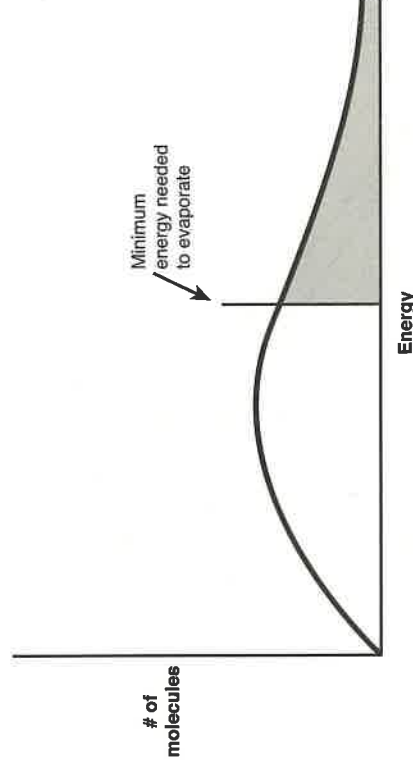


Figure 14.2: At elevated temperatures, a larger proportion of the molecules have enough energy to evaporate. As a result, evaporation takes place more quickly in warm liquids than cool ones.

Vapor Pressure and Boiling

As the temperature of a liquid increases, the vapor pressure due to the evaporation of the molecules also increases. Eventually, when the vapor pressure of a liquid becomes equal to the vapor pressure of the surrounding gas, it begins to boil. The *normal boiling point* of a liquid is defined as the temperature at which its vapor pressure equals one atmosphere (atm).



DEFINITION

The **vapor pressure** of a liquid is the gas pressure in a closed container due to the molecules that have evaporated from the liquid. The **normal boiling point** of a liquid is the temperature at which its vapor pressure is 1.00 atm.

From the figure on the next page, you can see the dependence of the vapor pressure of **two** liquids on temperature.

As you can see, the vapor pressure of acetone is more than that of water at any given temperature. The reason for this is that the dipole-dipole forces attracting acetone molecules to one another are weaker than the hydrogen bonds attracting water molecules to one another. Consequently, at any given temperature, more acetone molecules than water molecules have enough energy to overcome the intermolecular forces present and enter the vapor phase through evaporation.

In pure water, any molecule at the surface of the liquid that has enough energy to evaporate can do so. However, in saltwater, the sodium and chloride ions (which don't evaporate) occupy some of the surface area of the liquid, decreasing the area over which the water molecules can evaporate. As a result, the vapor pressure of saltwater solutions is less than that of pure water at any given temperature.

Because saltwater has a lower vapor pressure than pure water, the boiling point of saltwater is higher than that of pure water. You can see this in the following figure, which shows the dependence of the vapor pressure of both pure water and saltwater on temperature.

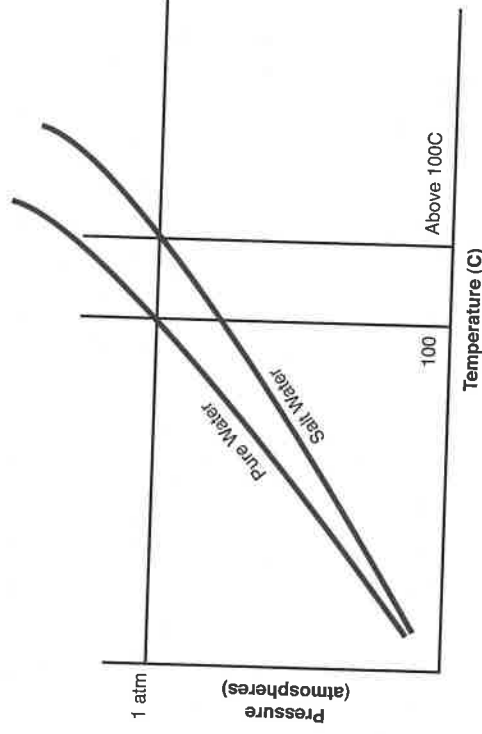


Figure 14.5: *The dependence of the vapor pressure of pure water and saltwater on temperature.*

From the diagram, you can see that the temperature required for saltwater to boil (which is, again, the temperature at which the vapor pressure of the liquid is 1 atm) is higher than that required for pure water. You can calculate this temperature from the concentration of the solution using the following equation:

$$\Delta T = K_b m_{\text{solute}}$$

ΔT represents the change in boiling point from the pure solvent, K_b is the boiling point elevation constant (which is different for every liquid), and m_{solute} is the molality of the solution (molality, in case you forgot the lessons of Chapter 11, is moles of solute per kilogram of solvent). The boiling point elevation constant is different for every liquid—for water, it is 0.51°C/m .

ΔT is the amount that the freezing point decreases, K_f is the freezing point depression constant (which is different for each substance), and m_{solute} is the molality of the solute. For water, K_f is 1.86°C/m .



CHEMISTRIVIA

The reason people salt roads in the winter is that the salt forms a concentrated solution with the water from the snow and ice on the road. Because salt solutions have lower melting points than pure water, the ice melts, making it less likely that an unlucky motorist will drive his or her car into a snowbank.

Boiling point elevation and freezing point depression are both examples of *colligative properties*. Colligative properties are any properties of a solution that depend on the concentration of solute in the solution. Because the melting and boiling points of solutions both depend on how much solute is present, both are colligative properties.



DEFINITION

Colligative properties are any properties of a solution that depend on the concentration of the solute.



YOU'VE GOT PROBLEMS

Problem 4: What does the concentration of an aqueous NaOH solution have to be to give it a melting point of -1.50°C ?

Melting and Freezing

Melting is the process by which a solid becomes a liquid. Freezing occurs when a liquid is converted to a solid. The freezing point of a solution is the same as its melting point, because both of these processes happen at the same temperature.

What Happens When Something Melts?

You're familiar with what happens when ice melts. As the ice warms up, it starts forming a big puddle until the ice has completely vanished. This is what we observe with our naked eyes, but what *really* happens at the molecular level?

The answer: It depends on what kind of material is melting. As it turns out, ionic and covalent compounds melt in different ways.

What Happens When Something Freezes?

When something freezes, the same thing happens as when it melts, except in reverse. As the temperature of the compound decreases, the molecules or ions (depending on the type of compound) have less and less energy. Eventually, the particles have so little energy that the forces between them lock the particles in place, reforming the crystal.

Boiling and Condensing

As we mentioned, when a compound boils, its vapor pressure has increased to the ambient atmospheric pressure. Condensing—the opposite process—occurs when a gas is converted to a liquid because its vapor pressure has decreased below the ambient atmospheric pressure. As with melting and freezing, these processes are the opposite of one another.

What Happens When a Liquid Boils?

We all know that if you heat a pot of water on the stove, eventually the water turns to steam and vanishes. However, you might not be familiar with why this happens.

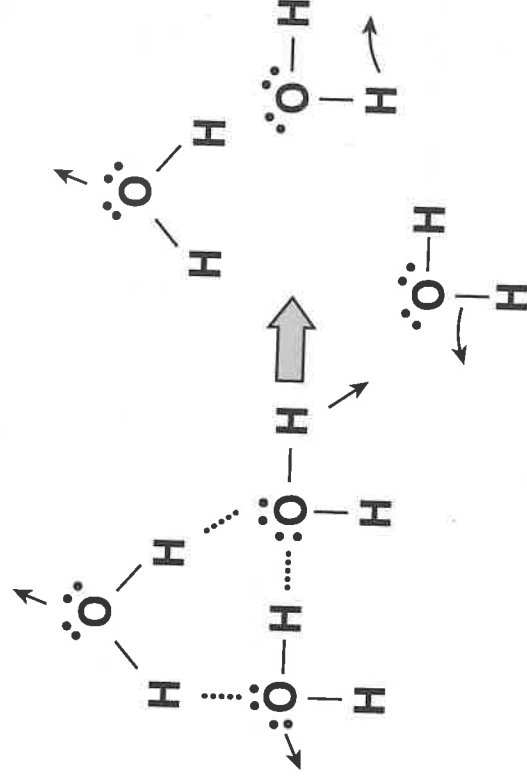


Figure 14.7: Water molecules boil when the amount of energy added to them becomes greater than the strength of the hydrogen bonds that attract them to one another.

diagram, find the conditions of temperature and pressure that you're interested in investigating on the chart. The region where this point can be found on the graph indicates the stable phase of matter for the substance. The following figure shows the phase diagram of water.

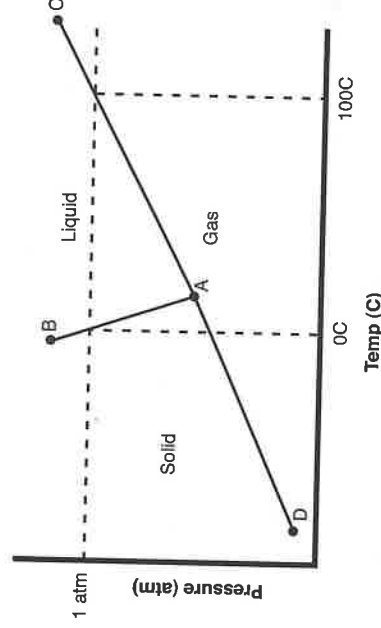


Figure 14.8: The phase diagram of water.

You can use this chart to learn about the important parts of phase diagrams.

- Point A is called the triple point. The triple point in a phase diagram is the temperature at which the solid, liquid, and gas phases are all stable. As you can see from this diagram, the reason you haven't seen water under these conditions is that it corresponds to a very low temperature and pressure.
- The line between points A and B corresponds to a series of pressure and temperature values at which both solid and liquid water can exist in equilibrium at the same time. If you follow the line at a pressure of 1 atm across the table, the temperature at which you cross this line is 0°C. By heating ice above this temperature, you cause it to melt; by cooling liquid water below this temperature, you cause it to freeze.



CHEMISTRIVIA

If you have a very cold refrigerator, you can re-create the conditions at which solid ice and liquid water are in equilibrium. Make a glass of ice water and place it in the refrigerator before you go to bed. When you wake up in the morning, open the refrigerator—if your refrigerator is at exactly the freezing point of water, you'll have exactly the same ratio of ice to water in the glass that you started with!

Solubility: The Hidden Phase Change

You might have noticed that, while talking about phase changes, I didn't mention anything about what happens when a solid dissolves into a solvent. Why would I hide this from you? After all, if you start with a solid and end up with a liquid, isn't this a phase change?

Actually, no. Although a dissolved solid is said to be “in a different phase,” it has not undergone a “phase change” because if you removed the solvent, it would go back to being a boring old solid.

Now, this isn't to say that nothing interesting happens when a solid dissolves in a liquid. Sometimes this process gives off a lot of heat (called the heat of solvation), which can be seen when dissolving sodium hydroxide into a beaker of water. However, being interesting doesn't make something a phase change.

The Least You Need to Know

- Solutions of nonvolatile solutes have higher boiling points and lower melting points than pure solvents because their vapor pressures are lowered.
- When compounds melt, the molecules have enough energy that they can wiggle away from one another and move freely as a liquid. When compounds boil, the molecules in the liquid have enough energy to completely break free from each other.
- Phase diagrams enable you to keep track of a material's state under all conditions of temperature and pressure.

How Reactions Occur

Part

4

Chemicals react with each other! That probably doesn't come as much of a surprise, but up to this point, we really haven't talked much about chemical reactions.

As it turns out, chemical reactions work a lot like cooking. You mix a bunch of things together, you do some stuff to them, and you end up with something entirely different. We start this part by learning about chemical equations and stoichiometry so you can start cooking in a chemical way.

When you know how to cook, it's time to learn about how, as a good cook, you can determine how long it will take for recipes to be complete using the magic of "kinetics." Finally, we talk about equilibria—unlike regular cooking, sometimes the finished product turns itself right back into the ingredients!

Chemical Equations

Chapter

15

In This Chapter

- My famous chili recipe
- Balancing chemical equations
- Interpreting the symbols used in chemical equations
- Identifying the six types of chemical reactions
- Predicting the products of a chemical reaction

Though you've learned a lot of chemistry stuff, we haven't yet talked about chemical reactions! And because chemistry teachers like to ask questions about reactions, it's only fair that you learn something about them. However, before we can talk about the many ways chemicals combine, we must cover the notation and terminology that chemists use to discuss them. I'm speaking, of course, about chemical equations.

Many students think that chemical equations are hard to understand and write because they contain a lot of strange and unfamiliar symbols. If you're one of these people, don't worry about it—as you'll soon see, chemical equations are a piece of cake! Or chili

Let's Make Chili!

My family has a long-standing tradition that we follow every New Year's Day. While most people are watching the Rose Bowl parade on TV or nursing terrible hangovers, my family is firing up the stove, dicing onions, and yelling at the cat to get off the kitchen counter. You see, we take our New Year's Day chili very seriously.

Every member of my family makes chili differently. Some favor a heavy, stew-like chili, while others prefer a hot, less textured chili. However, in the entire history of my family,

This statement tells you a great deal of information. The chemicals to the left of the arrow represent *reactants* and are the ingredients for making the *products* represented at the right side of the equation.



DEFINITION

The ingredients needed to perform a chemical reaction are called **reactants** or **reagents**. The chemicals made in the reaction are called the **products** of the reaction. Together, the entire statement that includes both the products and reactants is called a chemical equation.

All chemical changes are expressed using equations in this general form. The numbers of reactants and products may change from one equation to another, but the general format is always the same.



THE MOLE SAYS

Most chemistry textbooks write the arrows for chemical reactions like this: \rightarrow . However, in this chapter and beyond, I write the arrows like this: \rightleftharpoons . I do this because all chemical reactions are reversible, which means that, in addition to reactants forming products, products can react in reverse to form reactants. Although in many cases the backward reaction doesn't occur at a significant rate, it does occur, making it important to write the \rightleftharpoons arrow instead of the \rightarrow arrow. We discuss this in greater detail in Chapter 19.

How Much of Each Ingredient Do We Need?

All good recipes tell the chef how much of each ingredient is required. The process by which you figure out how much of each reactant is required in a chemical reaction is called balancing the equation.

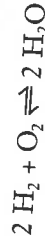
Equation Balancing Made Easy

Let's learn equation balancing by studying the reaction for making water from hydrogen and oxygen: $\text{H}_2 + \text{O}_2 \rightleftharpoons \text{H}_2\text{O}$. This equation, as well as any others you may come across, can be balanced by following these steps.

Step 1: Before balancing equations, tell yourself *never* to change any of the chemical formulas of either the reactants or the products. If you change the formulas by adding subscripts or altering them in any way, your equation is guaranteed to be *wrong!*

The good news is that the oxygen columns now match perfectly, which is exactly what you were trying to do by adding the 2 in front of H_2O . Unfortunately, the hydrogen column, which previously had matched, is now unbalanced. This may make it seem as if we've screwed something up, but this is a common event when balancing equations.

After examining the revised table, it becomes clear that you should add a 2 in front of H_2 before the arrow so that you have four hydrogen atoms to the left of the arrow. The revised equation now looks like this:



Because you've changed another number, you need to redo your table again. Fortunately, this time, the numbers of hydrogen and oxygen atoms are the same on both sides of the equation.

Element	Before the Arrow	After the Arrow
H	4	4
O	2	2

Equation Balancing Tips, Tricks, and Suggestions

As with recipes, some equations are more difficult than others. This can be frustrating, but you can try some things when the going gets tough:

- If you've been working on an equation for a few minutes and you aren't getting any closer to solving it, start over from scratch. This doesn't guarantee that you'll get the right answer, but it sometimes helps to get a fresh perspective on what you're doing.
- If you still can't solve the problem, start over and put a 2 in front of the most complicated-looking compound in the equation. If you still can't solve the problem, start over with a 3 in front of the most complicated-looking compound.
- If something seems like it might intuitively work, give it a shot. Just make sure that you keep track of how many atoms of every element are on the left and right sides of the arrow!

Symbol	What It Means	Example
(s)	The chemical is a solid.	Fe _(s)
(l)	The chemical is a liquid.	H ₂ O _(l)
(g)	The chemical is a gas.	CO _{2(g)}
(aq)	The chemical is aqueous (dissolved in water).	AgNO _{3(aq)}

Sometimes it's easy to tell which symbols of state should be used, and sometimes it's not. For example, water is frequently in a liquid form. However, if you do a chemical reaction that requires a large amount of heat, it may be a gas (for example, steam).



THE MOLE SAYS

If you're doing a chemical reaction in water, check out the solubility chart on the back of the periodic table in the front of the book, to help you figure out which ionic compounds are aqueous.

Reaction Conditions

Frequently, symbols are written around the arrow in a chemical reaction to indicate to the reader what procedures need to be followed to make a chemical reaction occur. Here are some of the most common symbols:

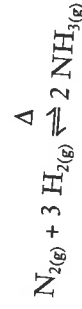
Symbol	What It Means
Δ	Add energy/heat to the reactants.
100° C	Heat the reactants to the specified temperature.
2 atm	The reactants should be combined at the specified pressure.
chemical	The specified chemical is needed for the formula reaction to proceed or is the solvent.
3 hrs	The reaction should proceed for the specified period of time.



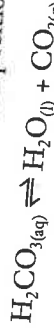
THE MOLE SAYS

You may sometimes see an arrow written after one of the products of a chemical equation. An arrow pointing up (as in CO₂ ↑) indicates that the product will form a gas. An arrow pointing down (as in PbI₂ ↓) indicates that the product will spontaneously precipitate (solidify) out of the solution.

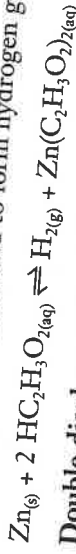
- **Synthesis reaction**—Synthesis reactions occur when small molecules combine to form larger ones. A commercially important example of a synthesis reaction is the Haber process, which results in the formation of ammonia from nitrogen and hydrogen.



- **Decomposition reaction**—Decomposition reactions are the opposite of synthesis reactions: they occur when larger molecules break apart to form smaller molecules. An example of a decomposition reaction is carbon dioxide bubbles formed by the decomposition of carbonic acid in a bottle of soda.



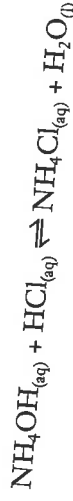
- **Single displacement reactions**—Also called single replacement reactions, these reactions occur when a pure element switches places with one of the elements in another chemical compound. An example of this type of reaction occurs when zinc reacts with acetic acid to form hydrogen gas and zinc acetate.



- **Double displacement reactions**—Also called double replacement reactions, these reactions occur when the positively charged cations of two ionic compounds switch places. A double displacement reaction takes place when dissolved magnesium sulfate is added to sodium hydroxide.



- **Acid-base reactions**—Acid-base reactions occur when an OH^- and H^+ ion combine to form water. An acid-base reaction occurs when household ammonium hydroxide combines with hydrochloric acid to form ammonium chloride and water.



We talk a lot more about acid-base reactions in Chapter 20, so stay tuned for more!

Li is higher on the activity series than copper, but $\text{Ag} + \text{NaNO}_3$ will not form $\text{AgNO}_3 + \text{Na}$ because silver is lower on the activity series than sodium.



THE MOLE SAYS

The activity series is essentially a ranking of how reactive various elements are. The higher an element is on the activity series, the more reactive it is and the more likely it is to displace another element in a single displacement reaction.

- For double displacement reactions, the product will be formed according to the general formula $\text{AB} + \text{CD} \rightleftharpoons \text{AD} + \text{CB}$. However, this reaction will produce a usable product only if both reactants are soluble in water and if only one of the products is soluble (again, check the solubility chart on the back of the periodic table at the front of the book). You see, if the reactants aren't soluble in water, they won't be able to combine with one another, and if both products are soluble, there will be no way of separating them.



BAD REACTIONS

Students make two common mistakes when predicting the products of a chemical reaction. The first is predicting the formation of a theoretically impossible product, such as NaCO_3 or Ag_2Cl . The second is failing to balance the equation after accurately predicting the products.

- For acid-base reactions, the product will be formed according to the general formula $\text{HA} + \text{BOH} \rightleftharpoons \text{BA} + \text{H}_2\text{O}$, where BA is an ionic compound.



YOU'VE GOT PROBLEMS

Problem 4: Write balanced chemical equations for the reactions that might occur when the following reactants are combined:

- $\text{NaOH} + \text{H}_2\text{SO}_4 \rightleftharpoons ?$
- $\text{NH}_3 + \text{I}_2 \rightleftharpoons ?$
- $\text{C}_3\text{H}_8\text{O} + \text{O}_2 \rightleftharpoons ?$
- $\text{Na} + \text{FeSO}_4 \rightleftharpoons ?$
- $\text{NaBr} + \text{NH}_4\text{OH} \rightleftharpoons ?$

The Least You Need to Know

- Chemical equations are recipes that tell you how to perform chemical reactions.
- It's important to balance equations so that they obey the law of conservation of mass.
- The states of the reactants and products of a reaction, as well as the actions we need to perform to make a reaction occur, can be expressed in symbols added to the chemical equation.
- It's handy to know the six types of chemical reaction if you're interested in predicting the products of a chemical reaction.
- Complete ionic equations and net ionic equations are useful ways of expressing what happens in a double displacement reaction.

Stoichiometry

Chapter

16

In This Chapter

- Simple stoichiometry calculations
- Ways to limit reactant problems
- Gas stoichiometry
- Percent yield calculations

As you may have gathered from the previous chapter, I like to cook. My chili recipe is perfect for serving a group of 10 (or a very hungry group of 4). However, what would happen if instead of 1 pound of hamburger, I had only half a pound? What would happen to the recipe then?

You're probably thinking to yourself, "What a knucklehead! Just adjust the ingredients so that the quantities of the other ingredients are also halved!" If that statement makes sense, then you already understand everything in this chapter. Unfortunately, chemists like to use fancy words when describing chemical reactions, so you're not out of the woods yet. While you read the following pages, remember that you already understand this material—it's just a matter of translating what you already know into chemical terms.

Stoichiometry: Fun to Say, Fun to Do!

Before I write another paragraph, let's all pronounce the word *stoichiometry* together. Ready, set, "stoy-key-ah-meh-tree." Say it again! Now say it five times as quickly as you can. I told you it was fun to pronounce!

Answer: Yikes! This doesn't look much like the ice water example, does it? Before you even attempt to solve this problem, let me give you a handy diagram to serve all your stoichiometric needs.

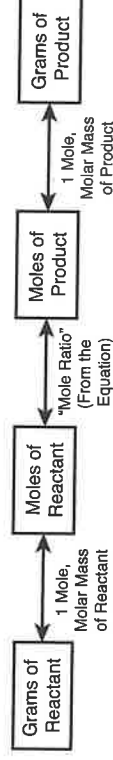


Figure 16.1: This handy chart can help you figure out how to do stoichiometry calculations.

To use the diagram, follow these steps:

1. Find the box that corresponds to the information you were given in the problem. In the example, you were given 48.0 g H₂, so the “grams of reactant” box is where you start.
2. Find the box that corresponds to the value that you’re ultimately trying to find. Because the problem wants you to determine how many grams of water can be formed, you want to end up at the box that corresponds to “grams of product.”
3. Write down the number and unit that you were given in the problem:
48.0 g H₂
4. Write a multiplication sign after the number and unit you were given, followed by a straight, horizontal line:
48.0 g H₂ × _____



THE MOLE SAYS

Stoichiometric calculations are set up and performed in the same manner as the unit conversions from Chapter 1 and the mole calculations from Chapter 8. If you have problems with stoichiometry, go back and do some unit conversion problems as a warm-up.

5. Below the line, write the same unit that’s in the top left. Include any chemical formulas:

48.0 g H₂ × _____
g H₂



THE MOLE SAYS

Although you found how much product can be made in the previous example, you can just as easily figure out how much of a reactant is required to form a given amount of product. In such an example, you simply go from right to left in the diagram instead of left to right. The math is done in *exactly* the same way!



YOU'VE GOT PROBLEMS

Problem 1: For the reaction that follows the equation $\text{CaCl}_2 + 2 \text{NaOH} \rightleftharpoons \text{Ca(OH)}_2 + 2 \text{NaCl}$, determine how many grams of sodium hydroxide need to react with an excess of calcium chloride to form 110.0 grams of NaCl.

Limiting Reactant Problems

Now that you're a pro at simple stoichiometry problems, let's try a more complex one. Using our recipe for ice water (1 glass of water + 4 ice cubes = 1 glass of ice water), determine how much ice water you can make if you have 10 glasses of water and 20 ice cubes.

Hopefully you didn't have much trouble figuring out that you can make only five glasses of ice water. In doing so, you probably did the following mental math:

- Using the recipe, 10 glasses of water can make 10 glasses of ice water.
- Likewise, 20 ice cubes can make five glasses of ice water.
- Because you run out of ice cubes before we run out of water, you can make only five glasses of ice water.

In this example, we say that ice is the *limiting reactant* because it's what you run out of first—it puts a limit on how much ice water you can make. Likewise, water is the *excess reactant* because you have more of it than you need.



DEFINITION

The **limiting reactant** in a stoichiometry problem is the reactant that runs out first in a reaction, thus limiting the amount of product that can be formed. The other reactant is called the **excess reactant**.

You can use the same method to solve stoichiometry calculations. Again, if you're given a problem in which you know the quantities of both reactants, all you need to do is figure

In the limiting reactant example for the formation of water, you found that you can form 2.75 moles of water by combining 1.75 moles of oxygen with 2.75 moles of hydrogen. Hydrogen was the limiting reactant, so let's figure out how much water would be left over:

$$\begin{aligned}\text{Leftover } \text{O}_2 &= 1.75 \text{ mol } \text{O}_2 - (1.75 \text{ mol } \text{O}_2)(2.75 \text{ mol } \text{H}_2/3.50 \text{ mol } \text{H}_2\text{O}) \\ &= 0.375 \text{ mol } \text{O}_2 \text{ remaining}\end{aligned}$$



YOU'VE GOT PROBLEMS

Problem 3: Using your results from Problem 2, determine the amount of excess reactant that will be left over.

Gas Stoichiometry

Thus far, we've limited our discussion of stoichiometry to grams and moles, but you can also do stoichiometric calculations using volumes of gases. To do so, we need to modify our diagram slightly.

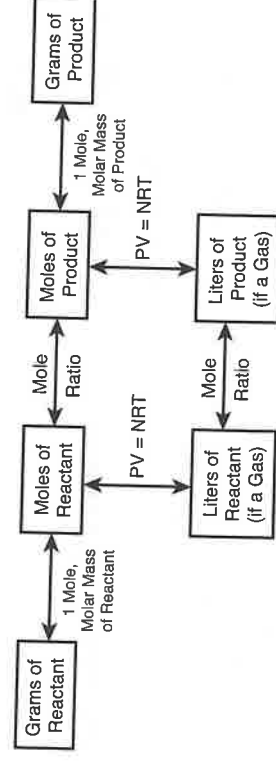


Figure 16.3: Our new and improved stoichiometric diagram.

To use this diagram, you need to be able to convert from liters of a gas to moles. Fortunately, you learned how to do this in Chapter 13 with the ideal gas law. If you've forgotten how to use the ideal gas law, it might be a good idea to brush up on it before continuing with this section.

Aside from this change, stoichiometric calculations for gases are done in exactly the same way. Start at the box that includes the information you've been given, and move through the diagram, box by box, until you arrive at your desired destination. Let's do an example:

Example: For the reaction $2 \text{H}_{2(g)} + \text{O}_{2(g)} \rightleftharpoons 2 \text{H}_2\text{O}_{(g)}$, determine how many liters of hydrogen gas are required to produce 175 g of water vapor (steam). Assume that you have an excess of oxygen gas, a partial pressure of hydrogen of 1.00 atm, and a temperature of 20.0°C .

However, we usually like to keep track of how badly we've screwed up, so we have a handy equation that gives us some idea of how we've done:

$$\text{Percent yield} = \frac{\text{Actual yield from experiment}}{\text{Theoretical yield from stoichiometry}} \times 100\%$$

Let's imagine that you've done a chemical reaction and your stoichiometry calculations said that you should make 30.0 grams of product. If you actually made only 15.0 grams of product, your percent yield would be $(15.0 \text{ g}/30.0 \text{ g}) \times 100\% = 50.0\%$. A yield of this sort indicates that you can certainly do better the next time you do this experiment!



THE MOLE SAYS

When doing a percent yield calculation, an answer greater than 100% is a Very Bad Thing. A yield of over 100% implies that you made more of your product than is theoretically possible, which is a fancy way of saying that you violated the law of conservation of mass. Because this is one of those inviolable laws in chemistry, a result like this requires a thorough search for the errors that caused it.



YOU'VE GOT PROBLEMS

Problem 5: If you actually produced 125 liters of gas when performing the experiment from Problem 4, what is your percent yield? What does this yield say about the quality of your laboratory skills?

The Least You Need to Know

- Stoichiometry is the mathematical method by which you relate the quantities of products and reactants of a chemical reaction to each other.
- The reactant that runs out first in a chemical process is called the limiting reactant, and the one that is left over after the reaction is the excess reactant.
- By using the ideal gas law, you can perform stoichiometry calculations using gases.
- Percent yield calculations are useful in telling you how well you performed a reaction.