Part III
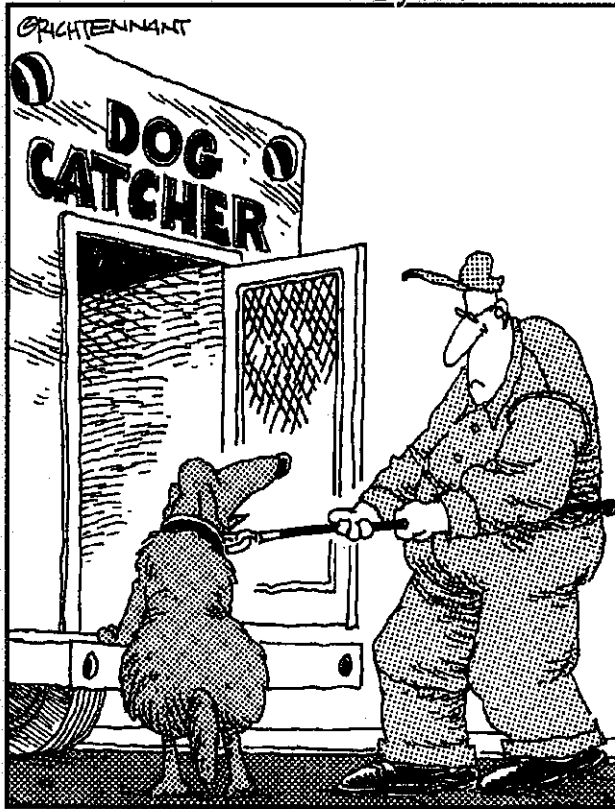# The Mole: The Chemist's Best Friend

"So what if you have a Ph.D. in chemistry?
I used to have my own circus act."

In this part . . .

*C*hemists operate in the world of quantities that they can see and touch — the macroscopic world of grams, liters, and meters. They carry out chemical reactions by weighing grams of reactants, and they use grams to measure the amount of product formed. They use liters to measure the amount of gas produced. They test a solution with litmus paper to see if it's an acid or a base.

But chemists also operate in the microscopic world of atoms and molecules. Because atoms and molecules are so very small, chemists have only just recently been able to see them, thanks to advanced technology in the most powerful of microscopes. Chemists think of acids and bases in terms of the donating and accepting of protons, not just as color changes in indicators. Models help chemists understand and predict the processes that take place in the microscopic world. These models also translate into real-life applications.

These chapters show you the bridge between the macroscopic and microscopic worlds — the mole. I explain reaction stoichiometry — how much reactant it takes to produce a particular amount of product. I introduce you to solutions and colloids, acids and bases, and gases. I show you the multitude of relationships between the properties of gases and relate them back to stoichiometry. In chemistry, everything is connected.

Chapter 10

The Mole: Can You Dig It?

● ●

● ●

*C*hemists do a lot of things. One is to make new substances, a process called *synthesis*. And a logical question they ask is "How much?"

"How much of this reactant do I need to make this much product?" "How much product can I make with this much reactant?" In order to answer these questions, chemists must be able to take a balanced chemical equation, expressed in terms of atoms and molecules, and convert it to grams or pounds or tons — some type of unit that they can actually weigh out in the lab. The mole concept enables chemists to move from the microscopic world of atoms and molecules to the real word of grams and kilograms and is one of the most important central concepts in chemistry. In this chapter, I introduce you to Mr. Mole.

Counting by Weighing

Suppose that you have a job packing 1,000 nuts and 1,000 bolts in big bags, and you get paid for each bag you fill. So what's the most efficient and quickest way of counting out the nuts and bolts? Weigh out a hundred, or even ten, of each and then figure out how much a thousand of each will weigh. Fill up the bag with nuts until it weighs the amount you figured for 1,000 nuts. After you have the correct amount of nuts, use the same process to fill the bag with bolts. In other words, count by weighing; that's one of the most efficient ways of counting large numbers of objects.

In chemistry, you count very large numbers of particles, such as atoms and molecules. To count them efficiently and quickly, you need to use the count-by-weighing method, which means you need to know how much individual atoms and molecules weigh. You can get the weights of the individual atoms on the

periodic table, but what about the weights of the compounds? Well, you can simply add together the weights of the individual atoms in the compound to figure the molecular weight or formula weight. (*Molecular weights* refer to covalently bonded compounds, and *formula weights* refer to both ionic and covalent compounds. Check out Chapters 6 and 7 for details on ionic and covalent bonds.)

Here's a simple example that shows how to calculate the molecular weight of a compound: Water, H_2O, is composed of two hydrogen atoms and one oxygen atom. By looking on the periodic table, you can find that one hydrogen atom equals 1.0079 amu and one oxygen atom weighs 15.999 amu (amu stands for *atomic mass units* — see Chapter 3 for details). To calculate the molecular weight of water, you simply add together the atomic weights of two hydrogen atoms and one oxygen atom:

2×1.0079 amu = 2.016 amu	(two hydrogen atoms)
1×15.999 amu = 15.999 amu	(one oxygen atom)
2.016 amu + 15.999 amu = 18.015 amu	(the weight of the water molecule)

Now try a little harder one. Calculate the formula weight of aluminum sulfate, $Al_2(SO_4)_3$. In this salt, you have 2 aluminum atoms, 3 sulfur atoms, and 12 oxygen atoms. After you find the individual weights of the atoms on the periodic table, you can calculate the formula weight like this:

[(2 x 26.982 amu)+(3 x 32.066 amu)+(12 x 15.999 amu)] = 332.150 amu
for the aluminum   for the sulfur   for the oxygen   for $Al_2(SO_4)_3$

Pairs, Dozens, Reams, and Moles

When we humans deal with objects, we often think in terms of a convenient amount. For example, when a woman buys earrings, she normally buys a pair of them. When a man goes to the grocery store, he buys eggs by the dozen. And when I go to the office supply store, I buy copy paper by the ream.

We use words to represent numbers all the time — a pair is 2, a dozen is 12, and a ream is 500. All these words are units of measure, and they're convenient for the objects they're used to measure. Rarely would you want to buy a ream of earrings or a pair of paper.

Likewise, when chemists deal with atoms and molecules, they need a convenient unit that takes into consideration the very small size of atoms and molecules. There is such a unit. It's called a mole.

Avogadro's number: Not in the phone book

The word *mole* stands for a number — 6.022×10^{23}. It's commonly called *Avogadro's number*, named after Amedeo Avogadro, the scientist who laid the groundwork for the mole principle.

Now a mole — 6.022×10^{23} — is a really *big* number. When written in long-hand notation, it's

602,200,000,000,000,000,000,000

And *that* is why I like scientific notation.

If you had a mole of marshmallows, it would cover the United States to a depth of about 600 miles. A mole of rice grains would cover the land area of the world to a depth of 75 meters. And a mole of moles . . . no, I don't even want to think about that!

Avogadro's number stands for a certain number of *things*. Normally, those things are atoms and molecules. So the mole relates to the microscopic world of atoms and molecules. But how does it relate to the macroscopic world where I work?

The answer is that a *mole* (abbreviated as mol) is also the number of particles in exactly 12 grams of a particular isotope of carbon (C-12). So if you have exactly 12 grams of ^{12}C, you have 6.022×10^{23} carbon atoms, which is also a mole of ^{12}C atoms. For any other element, a mole is the atomic weight expressed in grams. And for a compound, a mole is the formula (or molecular) weight in grams.

Using moles in the real world

The weight of a water molecule is 18.015 amu (see the section "Counting by Weighing" for how to calculate the weight of compounds). Because a mole is the formula (or molecular) weight in grams of a compound, you can now say that the weight of a mole of water is 18.015 grams. You can also say that 18.015 grams of water contains 6.022×10^{23} H_2O molecules, or a mole of water. And the mole of water is composed of two moles of hydrogen and one mole of oxygen.

The mole is the bridge between the microscopic and the macroscopic world:

6.022×10^{23} particles $\leftrightarrow$ mole $\leftrightarrow$ atomic/formula weight in grams

If you have any one of the three things — particles, moles, or grams — then you can calculate the other two.

For example, suppose you want to know how many water molecules there are in 5.50 moles of water. You can set up the problem like this:

5.50 moles $\times$ 6.022 x 10^{23}molecules/mol = 3.31 $\times$ 10^{24} molecules

Or suppose that you want to know how many moles are in 25.0 grams of water. You can set up the problem like this (and see Appendix B for more on exponential arithmetic):

$$\frac{25.0g\,H_2O}{1} \times \frac{1mol\,H_2O}{18.015g\,H_2O} = 1.39\,moles\,H_2O$$

You can even go from grams to particles by going through the mole. For example, how many molecules are there in 100.0 grams of carbon dioxide?

The first thing you have to do is determine the molecular weight of CO_2. Look at the periodic table to find that one carbon atom equals 12.011 amu and one oxygen atom weighs 15.999 amu. Now figure the molecular weight, like this:

[(1 $\times$ 12.011 g/mol) + (2 $\times$ 15.999 g/mol)] = 44.01 g/mol for CO_2

Now you can work the problem:

$$\frac{100.0g\,CO_2}{1} \times \frac{1mol\,CO_2}{44.01g} \times \frac{6.022\times 10^{23}\,molecules}{1mol} = 1.368\times 10^{24}\,CO_2\,molecules$$

And it's just as easy to go from particles to moles to grams.

You can also use the mole concept to calculate the empirical formula of a compound using the *percentage composition* data for that compound — the percentage by weight of each element in the compound. (The *empirical formula* indicates the different types of elements in a molecule and the lowest whole-number ratio of each kind of atom in the molecule. See Chapter 7 for details.)

When I try to determine the empirical formula of a compound, I often have percentage data available. The determination of the percentage composition is one of the first analysis that a chemist does in learning about a new compound. For example, suppose I determine that a particular compound has the following weight percentage of elements present: 26.4% Na, 36.8% S, and 36.8% O. Since I'm dealing with percentage data (amount per hundred) I will assume that I have 100 grams of the compound so that my percentages can be used as weights. I then convert each mass to moles, like this:

$$\frac{26.4g\,Na}{1} \times \frac{1\,mol\,Na}{22.99g} = 1.15\,mol\,Na$$

$$\frac{36.8g\,S}{1} \times \frac{1\,mol\,S}{32.07} = 1.15\,mol\,S$$

$$\frac{36.8g\,O}{1} \times \frac{1\,mol\,O}{16.00g} = 2.30\,mol\,O$$

Now I can write an empirical formula of $Na_{1.15}S_{1.15}O_{2.30}$. I know that my subscripts have to be whole numbers, so I divide each of these by the smallest, 1.15, to get $NaSO_2$. (If a subscript is 1, it's not shown.) I can then calculate a weight for the empirical formula, by adding together the atomic masses on the periodic table of 1 sodium (Na), 1 sulfur (S) and 2 oxygen (O). This gives me an empirical formula weight of 87.056 grams. Suppose, however, in another experiment I determined that the actual molecular weight of this compound was 174.112 grams. By dividing 174.112 grams by 87.056 grams (actual molecular weight by the empirical formula weight) I get 2. This means that the molecular formula is twice the empirical formula, so that the compound is actually $Na_2S_2O_4$.

Chemical Reactions and Moles

I think one of the reasons I enjoy being a chemist is that I like to cook. I see a lot of similarities between cooking and chemistry. A chemist takes certain things called reactants and makes something new from them. A cook does the same thing. He or she takes certain things called ingredients and makes something new from them.

For example, I like to fix Fantastic Apple Tarts (FATs). My recipe looks something like this:

apples + sugar + flour + spices = FATs

No, wait. My recipe has amounts in it. It looks more like this:

4 cups of apples + 3 cups of sugar + 2 cups of flour + 1/10 cup of spices = 12 FATs

My recipe tells me how much of each ingredient I need and how many FATs I can make. I can even use my recipe to calculate how much of each ingredient I need for a particular number of FATs. For example, suppose that I'm giving a big dinner party, and I need 250 FATs. I can use my recipe to calculate the amount of apples, sugar, flour, and spices I need. Here, for example, is how I calculate how much sugar I need:
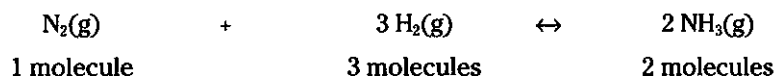
$$\frac{250\,FATs}{1} \times \frac{3\,cups\,sugar}{12\,FATs} = 62.5\,cups\,sugar$$

And I can do the same for the apples, flour, and spices by simply changing the ratio of each *ingredient* (as a multiple of 12 FATs).
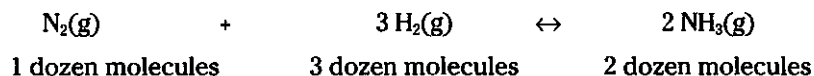
The balanced chemical equation allows you to do the same thing. For example, look at my favorite reaction, the Haber process, which is a method for preparing ammonia (NH_3) by reacting nitrogen gas with hydrogen gas:
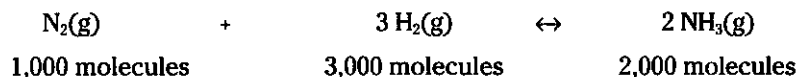
$$N_2(g) + 3\,H_2(g) \leftrightarrow 2\,NH_3(g)$$
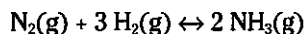
In Chapter 8, I use this reaction over and over again for various examples (like I said, it's my *favorite* reaction) and explain that you can read the reaction like this: 1 molecule of nitrogen gas reacts with 3 molecules of hydrogen gas to yield 2 molecules of ammonia.
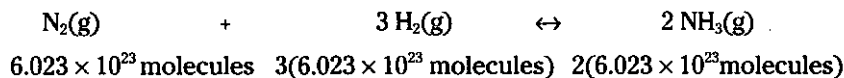
$N_2(g)$	+	$3\,H_2(g)$	$\leftrightarrow$	$2\,NH_3(g)$
1 molecule		3 molecules		2 molecules

Now, you can scale everything up by a factor of 12:

$N_2(g)$	+	$3\,H_2(g)$	$\leftrightarrow$	$2\,NH_3(g)$
1 dozen molecules		3 dozen molecules		2 dozen molecules

You can even scale it up by 1,000:

$N_2(g)$	+	$3\,H_2(g)$	$\leftrightarrow$	$2\,NH_3(g)$
1,000 molecules		3,000 molecules		2,000 molecules

Or how about a factor of 6.023×10^{23}:

$N_2(g)$	+	$3\,H_2(g)$	$\leftrightarrow$	$2\,NH_3(g)$
6.023×10^{23} molecules	$3(6.023 \times 10^{23}$ molecules)	$2(6.023 \times 10^{23}$ molecules)

Wait a minute! Isn't 6.023×10^{23} a mole? So you can write the equation like this:

$N_2(g)$	+	$3\,H_2(g)$	$\leftrightarrow$	$2\,NH_3(g)$
1 mole		3 moles		2 moles
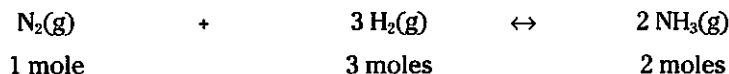
That's right — not only can those coefficients in the balanced chemical equation represent atoms and molecules, but they can also represent the number of moles.

Now take another look at my recipe for FATs:

4 cups of apples + 3 cups of sugar + 2 cups of flour + 1/10 cup of spices = 12 FATs

I have a problem. When I go to the grocery store, I don't buy fresh apples by the cup. Nor do I buy sugar or flour by the cup. I buy all these things by the pound. Now I buy a large excess, but because I'm frugal (translate as cheap), I want to figure as closely as I can the amount I really need. If I can determine the weight per cup for each ingredient, I'll be okay. So I weigh the ingredients and get

> 1 cup of apples = 0.5 lbs; 1 cup of sugar = 0.7 lbs; 1 cup of flour = 0.3 lbs; and 1 cup of spices = 0.2 lbs

Now I can substitute the measurements into my recipe:

> 4 cups of apples + 3 cups of sugar + 2 cups of flour + 1/10 cup of spices = 12 FATs
>
> 4(0.5lbs) 3(0.7 lbs) 2(0.3 lbs) 1/10(0.2 lbs)

Now if I want to know how many pounds of apples I need to make 250 FATs, I can set up the equation this way:

$$\frac{250\ FATs}{1} \times \frac{4\ cups\ apples}{12\ FATs} \times \frac{0.5\ lbs}{1\ cup\ apple} = 41.7\ lbs\ apples$$

I can figure out how much of each ingredient I need (based on weight), just by using the correct ingredient weight per cup.

The exact same thing is true with chemical equations. If you know the formula weight of the reactants and product, you can calculate how much you need and how much you'll get. For example, check out that Haber reaction again:

> $N_2(g)$ + $3 H_2(g)$ ↔ $2 NH_3(g)$
>
> 1 mole   3 moles   2 moles
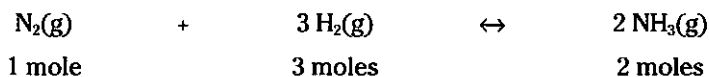
All you need to do is figure the molecular weights of each reactant and product, and then incorporate the weights into the equation. Use the periodic table to find the weights of the atoms and the compound (see the section "Counting by Weighing," earlier in this chapter, for the directions) and multiply those numbers by the number of moles, like this:
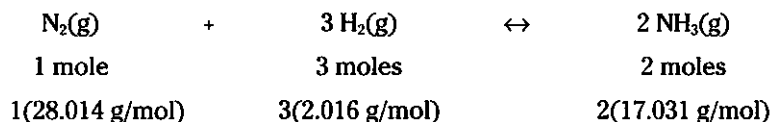
> 1(28.014 g/mol)   3(2.016 g/mol)   2(17.031 g/mol)
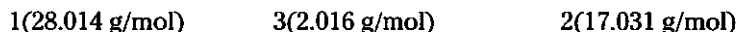
How much needed, how much made: Reaction stoichiometry

Once you have the weight relationships in place, you can do some stoichiometry problems. *Stoichiometry* refers to the mass relationship in chemical equations.

Look at my favorite reaction — you guessed it — the Haber process:

$$N_2(g) + 3\,H_2(g) \leftrightarrow 2\,NH_3(g)$$
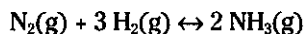
Suppose that you want to know how many grams of ammonia can be produced from the reaction of 75.00 grams of nitrogen with excess hydrogen. The mole concept is the key. The coefficients in the balanced equation are not only the number of individual atoms or molecules but also the number of moles.

| $N_2(g)$	+	$3\,H_2(g)$	$\leftrightarrow$	$2\,NH_3(g)$
1 mole		3 moles		2 moles
1(28.014 g/mol)		3(2.016 g/mol)		2(17.031 g/mol)

First, you can convert the 75.00 grams of nitrogen to moles of nitrogen. Then you use the ratio of the moles of ammonia to the moles of nitrogen from the balanced equation to convert to moles of ammonia. Finally you take the moles of ammonia and convert it to grams. The equation looks like this:

$$\frac{75.00g\ N_2}{1} \times \frac{1}{28.014g\ N_2} \times \frac{2\ mol\ NH_3}{1\ mol\ N_2}$$

$$\frac{17.031\ g\ NH_3}{1\ mol\ NH_3} \times = 91.19g\ NH_3$$

The ratio of the mol NH_3/mol N_2 is called a *stoichiometric ratio*. This ratio enables you to convert from the moles of one substance in a balanced chemical equation to the moles of another substance.

Getting tired of the Haber process? (Me? *Never.*) Take a look at another reaction — the reduction of rust (Fe_2O_3) to iron metal by treatment with carbon (coke). The balanced chemical reaction looks like this:

$$2\,Fe_2O_3(s) + 3\,C \rightarrow 4\,Fe(s) + 3\,CO_2(g)$$
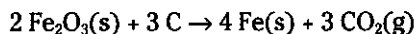
When you get ready to work stoichiometry types of problems, you *must* start with a balanced chemical equation. If you don't have it to start with, you've got to go ahead and balance the equation.

In this example, the formula weights you need are

- ✔ **Fe_2O_3:** 159.69 g/mol
- ✔ **C:** 12.01 g/mol
- ✔ **Fe:** 55.85 g/mol
- ✔ **CO_2:** 44.01 g/mol

Suppose that you want to know how many grams of carbon it takes to react with 1.000 kilogram of rust. You need to convert the kilogram of rust to grams and convert the grams to moles of rust. Then you can use a stoichiometric ratio to convert from moles of rust to moles of carbon and finally to grams. The equation looks like this:

$$\frac{1.000 Kg Fe_2O_3}{1} \times \frac{1000 g}{1 Kg} \times \frac{1 mol Fe_2O_3}{159.69 g Fe_2O_3} \times \frac{3 mol C}{2 mol Fe_2O_3} \times \frac{12.01 g C}{1 mol C} = 112.8 g C$$

You can even calculate the number of carbon atoms it takes to react with that 1.000 kilogram of rust. Basically, you use the same conversions, but instead of converting from moles of carbon to grams, you convert from moles of carbon to carbon atoms using Avogadro's number:

$$\frac{1.000 Kg Fe_2O_3}{1} \times \frac{1000 g}{1 Kg} \times \frac{3 mol Fe_2O_3}{2 mol Fe_2O_3} \times \frac{1 mol Fe_2O_3}{159.69 g Fe_2O_3} \times \frac{6.022 \times 10^{23} atoms}{1 mol C}$$

$2.839 \times 10^{24} C \ atoms$

Now I want to show you how to calculate the number of grams of iron produced from reacting 1.000 kilogram of rust with excess carbon. It's the same basic process as before — kilograms of rust to grams of rust to moles of rust to moles of iron to grams of iron:

$$\frac{1.000 Kg Fe_2O_3}{1} \times \frac{1000 g}{1 Kg} \times \frac{1 mol Fe_2O_3}{159.69 g Fe_2O_3} \times \frac{4 mol Fe_2O_3}{2 mol Fe_2O_3} \times \frac{55.85 g}{1 mol Fe} \ 699.5 g Fe$$

So you predict that you'll get 699.5 grams of iron metal formed. What if, however, you carry out this reaction and only get 525.0 grams of iron metal formed? There may be several reasons that you produce less than you expect, such as sloppy technique or impure reactants. It may also be quite likely that the reaction is an equilibrium reaction, and you'll never get 100 percent conversion from reactants to products. (Turn to Chapter 8 for details on equilibrium reactions.) Wouldn't it be nice if there was a way to label the efficiency of a particular reaction? There is. It's called the percent yield.

Where did it go? Percent yield

In almost any reaction, you're going to produce less than expected. You may produce less because most reactions are equilibrium reactions (see Chapter 8) or because some other condition comes into play. Chemists can get an idea of the efficiency of a reaction by calculating the *percent yield* for the reaction using this equation:

$$\% yield = \frac{Actual yield}{Theoretical yield} \times 100$$

The *actual yield* is how much of the product you get when you carry out the reaction. The *theoretical yield* is how much of the product you *calculate* you'll get. The ratio of these two yields gives you an idea about how efficient the reaction is. For the reaction of rust to iron (see the preceding section), your theoretical yield is 699.5 grams of iron; your actual yield is 525.0 grams. Therefore, the percent yield is

$$\%\text{yield} = \frac{525.0\,g}{699.5\,g} \times 100 = 75.05\%$$

A percent yield of about 75 percent isn't too bad, but chemists and chemical engineers would rather see 90+ percent. One plant using the Haber reaction has a percent yield of better than 99 percent. Now that's efficiency!

Running out of something and leaving something behind: Limiting reactants

I love to cook, and I'm always hungry. So I want to talk about making some ham sandwiches. Because I'm a chemist, I can write an equation for a ham sandwich lunch:

2 pieces of bread + 1 ham + 1 cheese → 1 ham sandwich
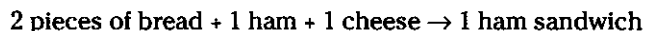
Suppose I check my supplies and find that I have 12 pieces of bread, 5 pieces of ham, and 10 slices of cheese. How many sandwiches can I make? I can make five, of course. I have enough bread for six sandwiches, enough ham for five, and enough cheese for ten. But I'm going to run out of ham first — I'll have bread and cheese left over. And the ingredient I run out of first really limits the amount of product (sandwiches) I'll be able to make; it can be called the *limiting ingredient.*

The same is true of chemical reactions. Normally, you run out of one of the reactants and have some others left over. (In some of the problems sprinkled throughout this chapter, I tell you which reactant is the limiting one by saying you have *an excess* of the other reactant(s).)

In this section, I show you how you can calculate which reactant is the limiting reactant.

Here is a reaction between ammonia and oxygen:

$4\,NH_3(g) + 5\,O_2(g) \rightarrow 4\,NO(g) + 6\,H_2O(l)$
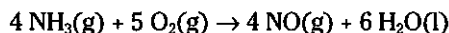
Suppose that you start out with 100.0 grams of both ammonia and oxygen, and you want to know how many grams of NO (nitrogen monoxide, sometimes called nitric oxide) you can produce. You must determine the limiting reactant and then base your stoichiometric calculations on it.

In order to figure out which reactant is the limiting reactant, you can calculate the mole-to-coefficient ratio: You calculate the number of moles of both ammonia and oxygen, and then you divide each by their coefficient in the balanced chemical equation. The one with the smallest mole-to-coefficient ratio is the limiting reactant. For the reaction of ammonia to nitric oxide, you can calculate the mole-to-coefficient ratio for the ammonia and oxygen like this:

$$\frac{100.0 gNH_3}{1} \times \frac{1 molNH_3}{17.03g} = 5.87\, mol \div 4 = 1.47$$

$$\frac{100.0 gNH_3}{1} \times \frac{1 molO_2}{32.00g} = 3.13\, mol \div 5 = 0.625$$

Ammonia has a mole-to-coefficient ratio of 1.47, and oxygen has a ratio of 0.625. Because oxygen has the lowest ratio, oxygen is the limiting reactant, and you need to base your calculations on it.

$$\frac{100.0 gO_2}{1} \times \frac{1 molO_2}{32.00g} \times \frac{4 molNO}{5 molO_2} \times \frac{30.01 gNO}{1 molNO} = 75.02 gNO$$

That 75.02 grams NO is your theoretical yield. But you can even calculate the amount of ammonia left over. You can figure the amount of ammonia consumed with this equation:

$$\frac{100.0 gO_2}{1} \times \frac{1 molO_2}{32.00g} \times \frac{4 molNH_3}{5 molO_2} \times \frac{17.03 gNH_3}{molNH_3} = 42.58 gNH_3$$

You started with 100.0 grams of ammonia, and you used 42.58 grams of it. The difference (100 grams – 42.58 grams = 57.42 grams) is the amount of ammonia left over.

Chapter 11

Mixing Matter Up: Solutions

⊕ ⊕

In This Chapter

▷ Finding out about solutes, solvents, and solutions

▷ Working with the different kinds of solution concentration units

▷ Checking out the colligative properties of solutions

▷ Figuring out colloids

⊕ ⊕

*Y*ou encounter solutions all the time in everyday life. The air you breathe is a solution. That sports drink you use to replenish your electrolytes is a solution. That soft drink *and* that hard drink are both solutions. Your tap water is most likely a solution, too. In this chapter, I show you some of the properties of solutions. I introduce you to the different ways chemists represent a solution's concentration, and I tell you about the colligative properties of solutions and relate them to ice cream making and antifreeze. So sit back, sip on your solution of choice, and read all about solutions.

Solutes, Solvents, and Solutions

A *solution* is a homogeneous mixture, meaning that it is the same throughout. If you dissolve sugar in water and mix it really well, for example, your mixture is basically the same no matter where you sample it.

A solution is composed of a solvent and one or more solutes. The *solvent* is the substance that's present in the largest amount, and the *solute* is the substance that's present in the lesser amount. These definitions work most of the time, but there are a few cases of extremely soluble salts, such as lithium chloride, in which more than 5 grams of salt can be dissolved in 5 milliliters of water. However, water is still considered the solvent, because it's the species that has not changed state. In addition, there can be more than one solute in a solution. You can dissolve salt in water to make a brine solution, and then you can dissolve some sugar in the same solution. You then have two solutes, salt and sugar, but you still have only one solvent — water.

When I talk about solutions, most people think of liquids. But there can also be solutions of gases. Our atmosphere, for example, is a solution. Because air is almost 79 percent nitrogen, it's considered the solvent, and the oxygen, carbon dioxide, and other gases are considered the solutes. There are also solid solutions. Alloys, for example, are solutions of one metal in another metal. Brass is a solution of zinc in copper.

A discussion of dissolving

Why do some things dissolve in one solvent and not another? For example, oil and water will not mix to form a solution, but oil will dissolve in gasoline. There's a general rule of solubility that says *like-dissolves-like* in regards to polarity of both the solvent and solutes. Water, for example, is a polar material; it's composed of polar covalent bonds with a positive and negative end of the molecule. (For a rousing discussion of water and its polar covalent bonds, see Chapter 7.) Water will dissolve polar solutes, such as salts and alcohols. Oil, however, is composed of largely nonpolar bonds. So water will not act as a suitable solvent for oil.

You know from your own experiences, I'm sure, that there's a limit to how much solute can be dissolved in a given amount of solvent. Most of us have been guilty of putting far too much sugar in iced tea. No matter how much you stir, there's some undissolved sugar at the bottom of the glass. The reason is that the sugar has reached its maximum solubility in water at that temperature. *Solubility* is the maximum amount of solute that will dissolve in a given amount of a solvent at a specified temperature. Solubility normally has the units of grams solute per 100 milliliters of solvent (g/100 mL).

If you heat that iced tea, the sugar at the bottom will readily dissolve. The solubility is related to the temperature of the solvent. For solids dissolving in liquids, solubility normally increases with increasing temperature. However, for gases dissolving in liquids, such as oxygen dissolving in lake water, the solubility goes down as the temperature increases. This is the basis of *thermal pollution*, the addition of heat to water that decreases the solubility of the oxygen and affects the aquatic life.

Saturated facts

A *saturated* solution contains the maximum amount of dissolved solute possible at a given temperature. If it has less than this amount, it's called an *unsaturated* solution. Sometimes, under unusual circumstances, the solvent may actually dissolve more than its maximum amount and become *supersaturated.* This supersaturated solution is unstable, though, and sooner or later solute will precipitate (form a solid) until the saturation point has been reached.

If a solution is unsaturated, then the amount of solute that is dissolved may vary over a wide range. A couple of rather nebulous terms describe the relative amount of solute and solvent that you can use:

✔ You can say that the solution is *dilute,* meaning that, relatively speaking, there's very little solute per given amount of solvent. If you dissolve 0.01 grams of sodium chloride in a liter of water, for example, the solution is dilute. I once asked some students to give me an example of a dilute solution, and one replied "A $1 margarita." She was right — a lot of solvent (water) and a very little solute (tequila) are used in her example.

✔ A solution may be *concentrated,* containing a large amount of solute per the given amount of solvent. If you dissolve 200 grams of sodium chloride in a liter of water, for example, the solution is concentrated.

But suppose you dissolve 25 grams or 50 grams of sodium chloride in a liter of water? Is the solution dilute or concentrated? These terms don't hold up very well for most cases. And consider the case of IV solutions — they *must* have a very precise amount of solute in them, or the patient will be in danger. So you must have a quantitative method to describe the relative amount of solute and solvent in a solution. Such a method exists — solution concentration units.

Solution Concentration Units

You can use a variety of solution concentration units to quantitatively describe the relative amounts of the solute(s) and the solvent. In everyday life, percentage is commonly used. In chemistry, *molarity* (the moles of solute per liter of solution) is the solution concentration unit of choice. In certain circumstances, though, another unit, *molality* (the moles of solute per kilogram of solvent), is used. And I use parts-per-million or parts-per-billion when I discuss pollution control. The following sections cover some of these concentration units.

Percent composition

Most of us have looked at a bottle of vinegar and seen "5% acetic acid," a bottle of hydrogen peroxide and seen "3% hydrogen peroxide," or a bottle of bleach and seen "5% sodium hypochlorite." Those percentages are expressing the concentration of that particular solute in each solution. *Percentage* is the amount per one hundred. Depending on the way you choose to express the percentage, the units of amount per one hundred vary. Three different percentages are commonly used:

- ✔ Weight/weight (w/w) percentage
- ✔ Weight/volume (w/v) percentage
- ✔ Volume/volume (v/v) percentage

Unfortunately, although the percentage of solute is often listed, the method (w/w, w/v, v/v) is not. In this case, I normally assume that the method is weight/weight, but I'm sure you know about assumptions.

Most of the solutions I talk about in the following examples of these percentages are *aqueous* solutions, solutions in which water is the solvent.

Weight/weight percentage

In *weight/weight percentage*, or *weight percentage*, the weight of the solute is divided by the weight of the solution and then multiplied by 100 to get the percentage. Normally the weight unit is grams. Mathematically, it looks like this:

$$w/w\% = \frac{\text{grams solute}}{\text{grams solution}} \times 100$$

If, for example, you dissolve 5.0 grams of sodium chloride in 45 grams of water, the weight percent is

$$w/w\% = \frac{5.0\,\text{g NaCl}}{50\text{g solution}} \times 100 = 10\%$$

Therefore, the solution is a 10 percent (w/w) solution.

Suppose that you want to make 350.0 grams of a 5 percent (w/w) sucrose, or table sugar, solution. You know that 5 percent of the weight of the solution is sugar, so you can multiply the 350.0 grams by 0.05 to get the weight of the sugar:

350.0 grams × 0.05 = 17.5 grams of sugar

The rest of the solution (350.0 grams − 17.5 grams = 332.5 grams) is water. You can simply weigh out 17.5 grams of sugar and add it to 332.5 grams of water to get your 5 percent (w/w) solution.

Weight percentage is the easiest percentage solution to make, but sometimes you may need to know the volume of the solution. In this case, you can use the weight/volume percentage.

Weight/volume percentage

Weight/volume percentage is very similar to weight/weight percentage, but instead of using grams of solution in the denominator, it uses milliliters of solution:

$$w/v\% = \frac{\text{grams solute}}{\text{mL solution}} \times 100$$

Proof reading

When it comes to ethyl alcohol solutions, another concentration unit, called *proof*, is commonly used to measure the relative amount of alcohol and water. The *proof* is simply twice the percentage. A 50 percent ethyl alcohol solution is 100 proof. Pure ethyl alcohol (100 percent) is 200 proof. This term dates back to earlier times, when the production of ethyl alcohol for human consumption was a cottage industry. (In the part of North Carolina where I grew up, it still is a cottage industry.) There was no quality control back then, so the buyer had to be sure that the alcohol he was buying was concentrated enough (or "strong" enough) for the desired purpose. Some of the alcohol solution was poured over gunpowder and then lit. If there was enough alcohol present, the gunpowder would ignite, giving "proof" that the solution was strong enough.

Suppose that you want to make 100 milliliters of a 15 percent (w/v) potassium nitrate solution. Because you're making 100 milliliters, you already know that you're going to weigh out 15 grams of potassium nitrate (commonly called saltpeter — KNO_3). Now, here comes something that's a little different: You dissolve the 15 grams of KNO_3 in a little bit of water and dilute it to exactly 100 milliliters in a volumetric flask. In other words, you dissolve and dilute 15 grams of KNO_3 to 100 milliliters. (I tend to abbreviate *dissolve and dilute* by writing *d & d*, but sometimes it gets confused with Dungeons & Dragons. Yes, chemists are really, *really* nerds.) You won't know exactly how much water you put in, but it's not important as long as the final volume is 100 milliliters.

You can also use the percentage and volume to calculate the grams of solute present. You may want to know how many grams of sodium hypochlorite are in 500 milliliters of a 5 percent (w/v) solution of household bleach. You can set up the problem like this:

$$\frac{5g\,NaOCl}{100\,mL\,solution} \times \frac{500\,mL\,solution}{1} = 25g\,NaOCl$$

You now know that you have 25 grams of sodium hypochlorite in the 500 milliliters of solution.

Sometimes both the solute and solvent are liquids. In this case, it's convenient to use a volume/volume percentage.

Volume/volume percentage

With *volume/volume percentages*, both the solute and solution are expressed in milliliters:

$$v/v\% = \frac{mL\,solute}{mL\,solution} \times 100$$

Ethyl alcohol (the drinking alcohol) solutions are commonly made using volume/volume percentages. If you want to make 100 milliliters of a 50 percent

ethyl alcohol solution, you take 50 milliliters of ethyl alcohol and dilute it to 100 milliliters with water. Again, it's a case of dissolving and diluting to the required volume. You can't simply add 50 milliliters of alcohol to 50 milliliters of water — you'd get less than 100 milliliters of solution. The polar water molecules will attract the polar alcohol molecules. This tends to fill in the open framework of water molecules and prevents the volumes from simply being added together.

It's number one! Molarity

Molarity is the concentration unit most often used by chemists, because it utilizes moles. The mole concept is central to chemistry, and molarity lets chemists easily work solutions into reaction stoichiometry. (If you're cussing me out right now because you have no idea what burrowing, insect-eating mammals have to do with chemistry, let alone what stoichiometry is, just flip to Chapter 10 for the scoop. Your mother would probably recommend washing your mouth out with soap first.)

Molarity (M) is defined as the moles of solute per liter of solution. Mathematically, it looks like this:

$$M = \frac{mol\,solute}{L\,solution}$$

For example, you can take 1 mole (abbreviated as *mol*) of KCl (formula weight of 74.55 g/mol — you can get the scoop on formula and molecular weights in Chapter 10, too) and dissolve and dilute the 74.55 grams to 1 liter of solution in a volumetric flask. You then have a 1-molar solution of KCl. You can label that solution as 1 M KCl. You don't add the 74.55 grams to 1 liter of water. You want to end up with a final volume of 1 liter. When preparing molar solutions, always dissolve and dilute to the required volume. This process is shown in Figure 11-1.

Here's another example: If 25.0 grams of KCl are dissolved and diluted to 350.0 milliliters, how would you calculate the molarity of the solution? You know that molarity is moles of solute per liter of solution. So you can take the grams, convert them to moles using the formula weight of KCl (74.55 g/mol), and divide them by 0.350 liters (350.0 milliliters). You can set up the equation like this:

$$\frac{25.0g\,KCl}{1} \times \frac{1\,mol\,KCl}{74.55g} \times \frac{1}{0.350L} = 0.958M$$

Now suppose that you want to prepare 2.00 liters of a 0.550 M KCl solution. The first thing you do is calculate how much KCl you need to weigh:

$$\frac{0.550\,mol\,KCl}{L} \times \frac{74.55g\,KCl}{1\,mol} \times \frac{2.00L}{1} = 82.0g\,KCl$$

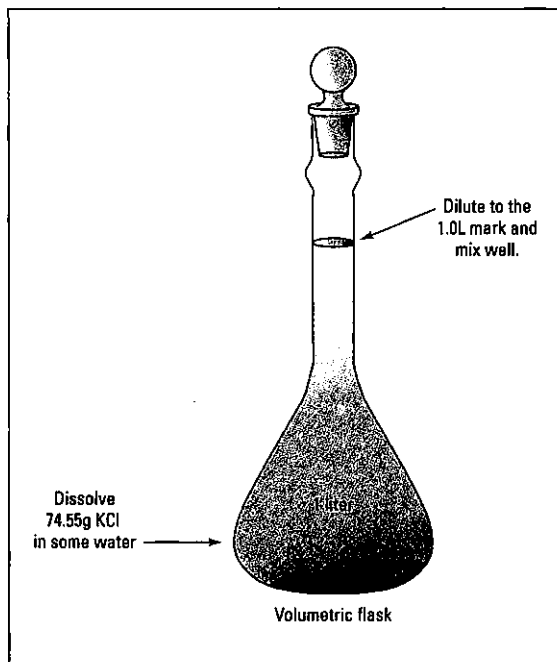You then take that 82.0 grams of KCl and dissolve and dilute it to 2.00 liters.

Figure 11-1:
Making a
1-molar KCl
solution.

Dilute to the
1.0L mark and
mix well.

Dissolve
74.55g KCl
in some water

1 liter

Volumetric flask

There's one more way to prepare solutions — the dilution of a more concentrated solution to a less-concentrated one. For example, you can buy hydrochloric acid from the manufacturer as a concentrated solution of 12.0 M. Suppose that you want to prepare 500 milliliters of 2.0 M HCl. You can dilute some of the 12.0 M to 2.0 M, but how much of the 12.0 M HCl is needed? You can easily figure the volume (V) you need by using the following formula:

$$V_{old} \times M_{old} = V_{new} \times M_{new}$$

In the preceding equation, V_{old} is the old volume, or the volume of the original solution, M_{old} is the molarity of the original solution, V_{new} is the volume of the new solution, and M_{new} is the molarity of the new solution. After substituting the values, you have

$$V_{old} \times 12.0 \text{ M} = 500.0 \text{ milliliters} \times 2.0 \text{ M}$$

$$V_{old} = (500.0 \text{ milliliters} \times 2.0 \text{ M})/12.0 \text{ M} = 83.3 \text{ milliliters}$$

You then take 83.3 milliliters of the 12.0 M HCl solution and dilute it to exactly 500.0 milliliters.

WARNING!

If you're actually doing a dilution of concentrated acids, be sure to *add the acid to the water* instead of the other way around! If the water is added to the concentrated acid, then so much heat will be generated that the solution will quite likely splatter all over you.

So to be safe, you should take about 400 milliliters of water, slowly add the 83.3 milliliters of the concentrated HCl as you stir, and then dilute to the final 500 milliliters with water.

The usefulness of the molarity concentration unit is readily apparent when dealing with reaction stoichiometry. For example, suppose that you want to know how many milliliters of 2.50 M sulfuric acid it takes to neutralize a solution containing 100.0 grams of sodium hydroxide. The first thing you must do is write the balanced chemical equation for the reaction:

$$H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow 2\,H_2O(l) + Na_2SO_4(aq)$$
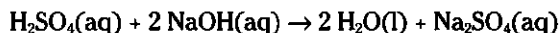
You know that you have to neutralize 100.0 grams of NaOH. You can convert the weight to moles (using the formula weight of NaOH, 40.00 g/mol) and then convert from moles of NaOH to moles of H_2SO_4. Then you can use the molarity of the acid solution to get the volume:

$$\frac{100.0\,g\,NaOH}{1} \times \frac{1\,mol\,NaOH}{40.00\,g} \times \frac{1\,mol\,H2SO4}{2\,mol\,NaOH} \times \frac{L}{2.50\,mol\,H2SO4} \times \frac{1000\,ML}{1L} = 500.0\,ml$$

It takes 500.0 milliliters of the 2.50 M H_2SO_4 solution to completely react with the solution that contains 100. grams of NaOH.

Molality: Another use for the mole

Molality is another concentration term that involves moles of solute. It isn't used very much, but I want to tell you a little about it, just in case you happen to run across it.

Molality (m) is defined as the moles of solute per kilogram of solvent. It's one of the few concentration units that doesn't use the solution's weight or volume. Mathematically, it looks like this:

$$m = \frac{mol\,solute}{Kg\,solvent}$$

Suppose, for example, you want to dissolve 15.0 grams of NaCl in 50.0 grams of water. You can calculate the molality like this (you must convert the 50.0 grams to kilograms before you use it in the equation):

$$\frac{15.0\,g\,NaCl}{1} \times \frac{1\,mol}{58.44\,g\,NaCl} \times \frac{1}{0.0500\,Kg} = 5.13m$$

Parts per million: The pollution unit

Percentage and molarity, and even molality, are convenient units for the solutions that chemists routinely make in the lab or the solutions that are commonly found in nature. However, if you begin to examine the concentrations of certain pollutants in the environment, you'll find that those concentrations

are very, very small. Percentage and molarity will work when you're measuring solutions found in the environment, but they're not very convenient. In order to express the concentrations of very dilute solutions, scientists have developed another concentration unit — parts per million.

Percentage is parts per hundred, or grams solute per 100 grams of solution. *Parts per million (ppm)* is grams solute per one million grams of solution. It's most commonly expressed as milligrams solute per kilogram solution, which is the same ratio. The reason it's expressed this way is that chemists can easily weigh out milligrams or even tenths of milligrams, and, if you're talking about aqueous solutions, a kilogram of solution is the same as a liter of solution. (The density of water is 1 gram per milliliter, or 1 kilogram per liter. The weight of the solute in these solutions is so very small that it's negligible when converting from the mass of the solution to the volume.)

By law, the maximum contamination level of lead in drinking water is 0.05 ppm. This number corresponds to 0.05 milligrams of lead per liter of water. That's pretty dilute. But mercury is regulated at the 0.002 ppm level. Sometimes, even this unit isn't sensitive enough, so environmentalists have resorted to the parts per billion (ppb) or parts per trillion (ppt) concentration units. Some neurotoxins are deadly at the parts per billion level.

Colligative Properties of Solutions

Some properties of solutions depend on the specific nature of the solute. In other words, an effect you can record about the solution depends on the specific nature of the solute. For example, salt solutions taste salty, while sugar solutions taste sweet. Salt solutions conduct electricity (they're electrolytes — see Chapter 6), while sugar solutions don't (they're non-electrolytes). Solutions containing the nickel cation are commonly green, while those containing the copper cation are blue.

There's also a group of solution properties that doesn't depend on the specific type of solute — just the *number* of solute particles. These properties are called *colligative properties* — properties that simply depend on the relative number of solute particles. The effect you can record about the solution depends on the number of solute particles present. These colligative properties — these effects — include

- ✔ Vapor pressure lowering
- ✔ Boiling point elevation
- ✔ Freezing point depression
- ✔ Osmotic pressure

Vapor pressure lowering

If a liquid is contained in a closed container, the liquid eventually evaporates, and the gaseous molecules contribute to the pressure above the liquid. The pressure due to the gaseous molecules of the evaporated liquid is called the liquid's *vapor pressure.*

If you take that same liquid and make it the solvent in a solution, the vapor pressure due to the solvent evaporation will be lower. This is because the solute particles in the liquid take up space at the surface and so the solvent can't evaporate as easily. And many times there may be an attraction between the solute and solvent that also makes it more difficult for the solvent to evaporate. And that lowering is independent of what kind of solute you use. Instead, it depends on the number of solute particles.

In other words, if you add one mole of sucrose to a liter of water and add one mole of dextrose to another liter of water, the amount that the pressure lowers will be the same, because you're adding the same *number* of solute particles. If, however, you add a mole of sodium chloride, the vapor pressure will be lowered by twice the amount of sucrose or glucose. The reasons is that the sodium chloride breaks apart into two ions, so adding a mole of sodium chloride yields two moles of particles (ions).

This lowering of vapor pressure partially explains why the Great Salt Lake has a lower evaporation rate than you may expect. The salt concentration is so high that the vapor pressure (and evaporation) has been significantly lowered.

Why use antifreeze in the summer? Boiling point elevation

Each individual liquid has a specific temperature at which it boils (at a given atmospheric pressure). This temperature is the liquid's *boiling point.* If you use a particular liquid as a solvent in a solution, you find that the boiling point of the solution is always higher than the pure liquid. This is called the *boiling point elevation.*

It explains why you don't replace your antifreeze with pure water in the summer. You want the coolant to boil at a higher temperature so that it will absorb as much engine heat as possible *without* boiling. You also use a pressure cap on your radiator, because the higher the pressure, the higher the boiling point. It also explains why a pinch of salt in the cooking water will cause foods to cook a little faster. The salt raises the boiling point so

that more energy can be transferred to cooking the food during a given amount of time.

As an FYI, you can actually calculate the amount of boiling point elevation by using this formula:

$$\Delta T_b = K_b m$$

ΔT_b is the *increase* in the boiling point, K_b is the boiling point elevation constant (0.512°C kg/mol for water), and m is the molality of particles. (For molecular substances, the molality of particles is the same as the molality of the substance; for ionic compounds, you have to take into consideration the formation of ions and calculate the molality of the ion particles.) Solvents other than water have a different boiling point elevation constant (K_b).

Making ice cream: Freezing point depression

Each individual liquid has a specific temperature at which it freezes. If you use a particular liquid as a solvent in a solution, though, you find that the freezing point of the solution is always lower than the pure liquid. This is called the *freezing point depression,* and it's a colligative property of a solution.

The depression of the freezing point of a solution relative to the pure solvent explains why you put rock salt in the ice/water mix when making homemade ice cream. The rock salt forms a solution with a lower freezing point than water (or the ice cream mix that's to be frozen). The freezing point depression effect also explains why a salt (normally calcium chloride, $CaCl_2$) is spread on ice to melt it. The dissolving of calcium chloride is highly exothermic (it gives off a lot of heat). When the calcium chloride dissolves, it melts the ice. The salt solution that's formed when the ice melts has a lowered freezing point that keeps the solution from refreezing. Freezing point depression also explains the use of antifreeze in your cooling system during the winter. The more you use (up to a concentration of 50/50), the lower the freezing point.

In case you're interested, you can actually calculate the amount the freezing point will be depressed:

$$\Delta T_f = K_f m$$

ΔT_f is the amount the freezing point will be lowered, K_f is the freezing point depression constant (1.86°C kg/mol for water), and m is the molality of the particles.

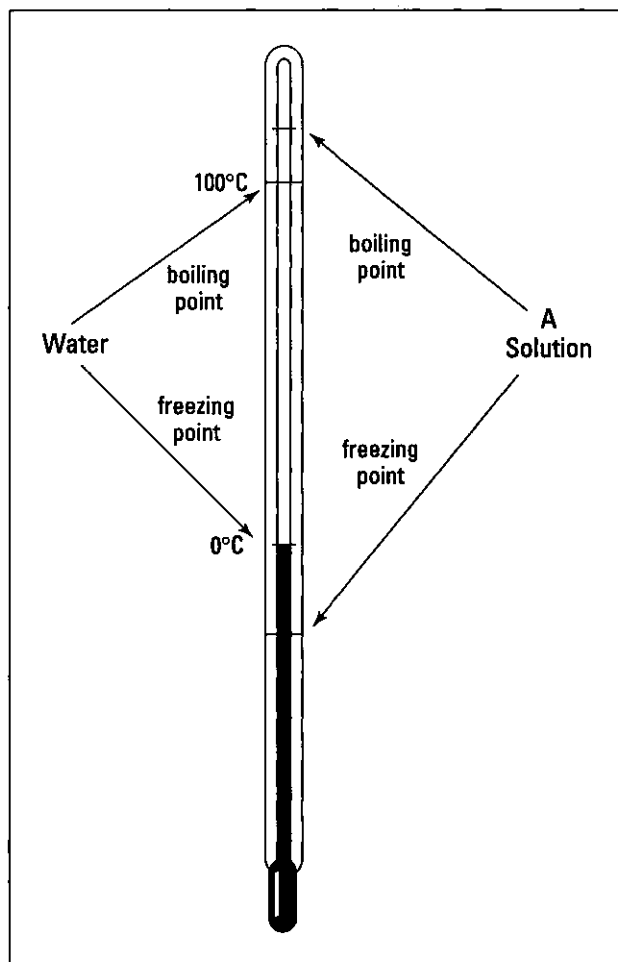Figure 11-2 shows the effect of a solute on both the freezing point and boiling point of a solvent.

Figure 11-2:
Boiling point
elevation
and freez-
ing point
depression
of a solution.

Keeping blood cells alive and well: Osmotic pressure

Suppose that you take a container and divide it into two compartments with a thin membrane containing microscopic pores large enough to allow water molecules but not solute particles to pass through. This membrane type is called a *semipermeable membrane;* it lets some small particles pass through but not other, larger particles.

You then add a concentrated salt solution to one compartment and a more dilute salt solution to the other. Initially, the two solution levels start out the same. But after a while, you notice that the level on the more concentrated side has risen, and the level on the more dilute side has dropped. This

change in levels is due to the passage of water molecules from the more dilute side to the more concentrated side through the semipermeable membrane. This process is called *osmosis*, the passage of a solvent through a semipermeable membrane into a solution of higher solute concentration. The pressure that you have to exert on the more concentrated side in order to stop this process is called *osmotic pressure*. This process is shown in Figure 11-3.
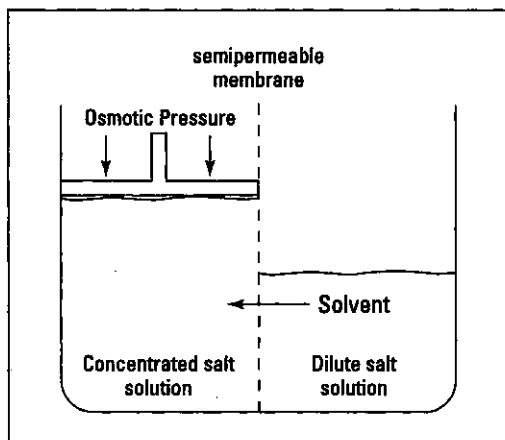

Figure 11-3:
Osmotic pressure.

The solvent always flows through the semipermeable membrane from the more dilute side to the more concentrated side. In fact, you can have pure water on one side and any salt solution on the other, and water always goes from the pure-water side to the salt-solution side. The more concentrated the salt solution, the more pressure it takes to stop the osmosis (the higher the osmotic pressure).

But what if you apply more pressure than is necessary to stop the osmotic process, exceeding the osmotic pressure? Water is forced through the semipermeable membrane from the more concentrated side to the more dilute side, a process called *reverse osmosis*. Reverse osmosis is a good, relatively inexpensive way of purifying water. My local "water store" uses this process to purify drinking water (so-called "RO water"). There are many reverse osmosis plants in the world, extracting drinking water from seawater. Navy pilots even carry small reverse osmosis units with them in case they have to eject at sea.

The process of osmosis is important in biological systems. Cell walls often act as semipermeable membranes. Do you ever eat pickles? Cucumbers are soaked in a brine solution in order to make pickles. The concentration of the solution inside the cucumber is less than the concentration of the brine solution, so water migrates through the cell walls into the brine, causing the cucumber to shrink.

One of the most biologically important consequences of osmotic pressure involves the cells within our own body. You can look at red blood cells as an example. There's an aqueous solution inside the blood cell and another aqueous solution outside the cell (intercellular fluid). When the solution outside the cell has the same osmotic pressure as the solution inside the cell, it's said to be *isotonic*. Water can be exchanged in both directions, helping to keep the cell healthy. However, if the intercellular fluid becomes more concentrated and has a higher osmotic pressure *(hypertonic)*, water flows primarily out of the blood cell, causing it to shrink and become irregular in shape. This is a process called *crenation*. The process may occur if the person becomes seriously dehydrated, and the crenated cells are not as efficient in carrying oxygen. If, on the other hand, the intercellular fluid is more dilute than the solution inside the cells and has a lower osmotic pressure *(hypotonic)*, the water flows mostly into the cell. This process, called *hemolysis*, causes the cell to swell and eventually rupture. Figure 11-4 shows crenation and hemolysis.
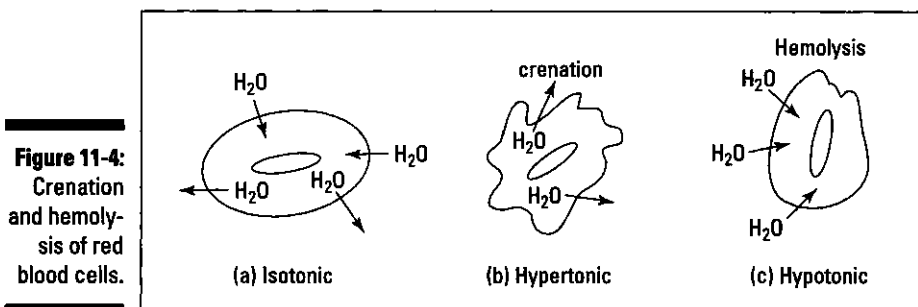

Figure 11-4:
Crenation and hemolysis of red blood cells.

(a) Isotonic   (b) Hypertonic   (c) Hypotonic

The processes of crenation and hemolysis explain why the concentration of IV solutions is so very critical. If they're too dilute, then hemolysis can take place, and if they're too concentrated, crenation is a possibility.

Smoke, Clouds, Whipped Cream, and Marshmallows: Colloids All

If you dissolve table salt in water, you form an aqueous solution. The solute particle size is very small — around 1 nanometer (nm), which is 1×10^9 meters. This solute doesn't settle to the bottom of a glass, and it can't be filtered out of the solution.

If, however, you go down to your local stream and dip out a glass of water, you'll notice that there's a lot of material in it. Many of the solute particles are larger than 1,000 nm. They quickly settle to the bottom of the glass and

can be filtered out. In this case, you have a *suspension* and not a solution. Whether you have one or the other depends on the size of the solute particles.

But there's also something in the intermediate range between solutions and suspensions. When the solute particle size is 1 to 1,000 nanometers, you have a *colloid*. Solutes in colloids don't settle out like they do in suspensions. In fact, it's sometimes difficult to distinguish colloids from true solutions. One of the few ways to distinguish between them is to shine a light through the suspected liquid. If it's a true solution, with very small solute particles, the light beam will be invisible. If you have a colloid, however, you'll be able to see the light beam as it reflects off the relatively large solute particles. This is called the *Tyndall effect*, and it's shown in Figure 11-5.
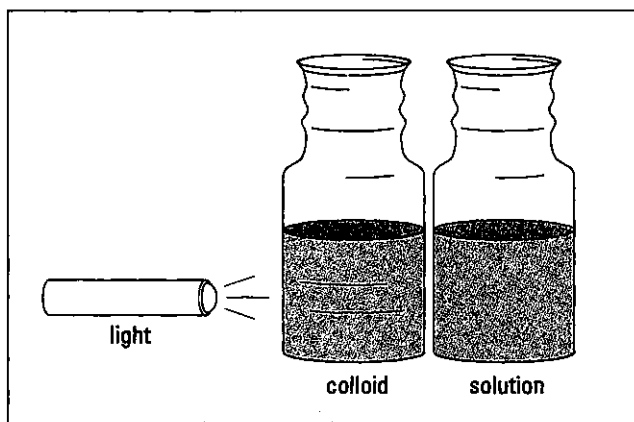

Figure 11-5:
The Tyndall
effect.

light

colloid   solution

There are many types of colloids. Have you ever eaten a marshmallow? It's a colloid of a gas in a solid. Whipped cream is a colloid of a gas in a liquid. Have you ever driven through the fog and seen your headlight beams? You were experiencing the Tyndall effect of a liquid-in-a-gas colloid. Smoke is a colloid of a solid (ash or soot) in a gas (air). Air pollution problems are often caused by the stability of this type of colloid.

Chapter 12

Sour and Bitter: Acids and Bases

*W*alk into any kitchen or bathroom, and you'll find a multitude of acids and bases. Open the refrigerator, and you'll find soft drinks full of carbonic acid. In the pantry, there's vinegar and baking soda, an acid and a base. Peek under the sink, and you'll notice the ammonia and other cleaners, most of which are bases. Check out that can of lye-based drain opener — it's highly basic. In the medicine cabinet, you'll find aspirin, an acid, and antacids of all types. Our everyday world is full of acids and bases. And so is the everyday world of the industrial chemist. In this chapter, I cover acids and bases, indicators and pH, and some good *basic* chemistry.

Properties of Acids and Bases: Macroscopic View

· Look at the properties of acids and bases that can be observed in the world around us.

Acids:

✔ Taste sour (but remember, in the lab, you test, not taste)

✔ Produce a painful sensation on the skin

✔ React with certain metals (magnesium, zinc, and iron) to produce hydrogen gas

✔ React with limestone and baking soda to produce carbon dioxide

✔ React with litmus paper and turn it red

Bases:

✔ Taste bitter (again, in the lab, you test, not taste)

✔ Feel slippery on the skin

✔ React with oils and greases

✔ React with litmus paper and turn it blue

✔ React with acids to produce a salt and water

Quite a number of acids and bases are found in our everyday life. Tables 12-1 and 12-2 show some common acids and bases found around the home.

Table 12-1	Common Acids Found in the Home	
Chemical Name	*Formula*	*Common Name or Use*
hydrochloric acid	HCl	muratic acid
acetic acid	CH_3COOH	vinegar
sulfuric acid	H_2SO_4	auto battery acid
carbonic acid	H_2CO_3	carbonated water
boric acid	H_3BO_3	antiseptic; eye drops
acetylsalicylic acid	$C_{16}H_{12}O_6$	aspirin

| Table 12-2	Common Bases Found in the Home	
Chemical Name	*Formula*	*Common Name or Use*
ammonia	NH_3	cleaner
sodium hydroxide	$NaOH$	lye
sodium bicarbonate	$NaHCO_3$	baking soda
magnesium hydroxide	$Mg(OH)_2$	milk of magnesia
calcium carbonate	$CaCO_3$	antacid
aluminum hydroxide	$Al(OH)_3$	antacid

What Do Acids and Bases Look Like? — Microscopic View

If you look at Tables 12-1 and 12-2 closely, you may recognize the fact that all the acids contain hydrogen, while most of the bases contain the hydroxide ion (OH). Two main theories use these facts in their descriptions of acids and bases and their reactions:

✔ Arrhenius theory

✔ Bronsted-Lowery theory

The Arrhenius theory: Must have water

The Arrhenius theory was the first modern acid-base theory developed. In this theory, an acid is a substance that, when dissolved in water, yields H^+ (hydrogen) ions, and a base is a substance that, when dissolved in water, yields OH (hydroxide) ions. HCl(g) can be considered as a typical Arrhenius acid, because when this gas dissolves in water, it *ionizes* (forms ions) to give the H^+ ion. (Chapter 6 is where you need to go for the riveting details about ions.)

$$HCl(aq) \rightarrow H^+ + Cl^-$$
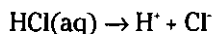
According to the Arrhenius theory, sodium hydroxide is classified as a base, because when it dissolves, it yields the hydroxide ion:

$$NaOH(aq) \rightarrow Na^+ + OH^-$$
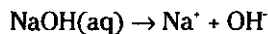
Arrhenius also classified the reaction between an acid and base as a *neutralization* reaction, because if you mix an acidic solution with a basic solution, you end up with a neutral solution composed of water and a salt.

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$
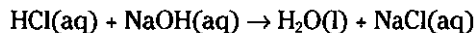
Look at the ionic form of this equation (the form showing the reaction and production of ions) to see where the water comes from:

$$H^+ + Cl^- + Na^+ + OH^- \rightarrow H_2O(l) + Na^+ + Cl^-$$
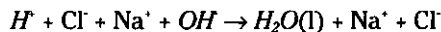
As you can see, the water is formed from combining the hydrogen and hydroxide ions. In fact, the net-ionic equation (the equation showing only those chemical substances that are changed during the reaction) is the same for all Arrhenius acid-base reactions:

$$H^+ + OH^- \rightarrow H_2O(l)$$
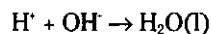
The Arrhenius theory is still used quite a bit. But, like all theories, it has some limitations. For example, look at the gas phase reaction between ammonia and hydrogen chloride gases:

$$NH_3(g) + HCl(g) \rightarrow NH_4^+ + Cl^- \rightarrow NH_4Cl(s)$$

The two clear, colorless gases mix, and a white solid of ammonium chloride forms. I show the intermediate formation of the ions in the equation so that you can better see what's actually happening. The HCl transfers an H^+ to the ammonia. That's basically the same thing that happens in the HCl/NaOH reaction, but the reaction involving the ammonia can't be classified as an acid-base reaction, because it doesn't occur in water, and it doesn't involve the hydroxide ion. But again, the same basic process is taking place in both cases. In order to account for these similarities, a new acid-base theory was developed, the Bronsted-Lowery theory.

The Bronsted-Lowery acid-base theory: Giving and accepting

The Bronsted-Lowery theory attempts to overcome the limitations of the Arrhenius theory by defining an acid as a proton (H^+) donor and a base as a proton (H^+) acceptor. The base accepts the H^+ by furnishing a lone pair of electrons for a *coordinate-covalent bond,* which is a covalent bond (shared pair of electrons) in which one atom furnishes both of the electrons for the bond. Normally, one atom furnishes one electron for the bond and the other atom furnishes the second electron (see Chapter 7). In the coordinate-covalent bond, one atom furnishes both bonding electrons.

Figure 12-1 shows the NH_3/HCl reaction using the electron-dot structures of the reactants and products. (Electron-dot structures are covered in Chapter 7, too.)

Figure 12-1:
Reaction of
NH_3 with
HCl.
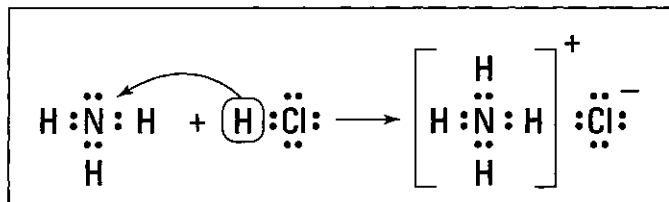
HCl is the proton donor, and the acid and ammonia are the proton acceptor, or the base. Ammonia has a lone pair of nonbonding electrons that it can furnish for the coordinate-covalent bond.

I discuss acid-base reactions under the Bronsted-Lowery theory in the section "Give me that proton: Bronsted-Lowery acid-base reactions," later in this chapter.

Acids to Corrode, Acids to Drink: Strong and Weak Acids and Bases

I want to introduce you to a couple of different categories of acids and bases — strong and weak. However, it's important to remember that acid-base strength is not the same as concentration. *Strength* refers to the amount of ionization or breaking apart that a particular acid or base undergoes. *Concentration* refers to the amount of acid or base that you initially have. You can have a concentrated solution of a weak acid, or a dilute solution of a strong acid, or a concentrated solution of a strong acid or . . . well, I'm sure you get the idea.

Strong acids

If you dissolve hydrogen chloride gas in water, the HCl reacts with the water molecules and donates a proton to them:

$$HCl(g) + H_2O(l) \rightarrow Cl^- + H_3O^+$$

The H_3O^+ ion is called the hydronium ion. This reaction goes essentially to completion, meaning the reactants keep creating the product until they're all used up. In this case, all the HCl ionizes to H_3O^+ and Cl^-; there's no more HCl present. Acids such as HCl, which ionizes essentially 100 percent in water, are called *strong acids*. Note that water, in this case, acts as a base, accepting the proton from the hydrogen chloride.

Because strong acids ionize completely, it's easy to calculate the concentration of the hydronium ion and chloride ion in solution if you know the initial concentration of the strong acid. For example, suppose that you bubble 0.1 moles (see Chapter 10 to get a firm grip on moles) of HCl gas into a liter of water. You can say that the initial concentration of HCl is 0.1 M (0.1 mol/L). *M* stands for molarity, and *mol/L* stands for moles of solute per liter. (For a detailed discussion of molarity and other concentration units, see Chapter 11.)

You can represent this 0.1 M concentration for the HCl in this fashion: [HCl] = 0.1. Here, the brackets around the compound indicate molar concentration, or mol/L. Because the HCl completely ionizes, you see from the balanced

equation that for every HCl that ionizes, you get one hydronium ion and one chloride ion. So the concentration of ions in that 0.1 M HCl solution is

$[H_3O^+] = 0.1$ and $[Cl^-] = 0.1$

This idea is valuable when you calculate the pH of a solution. (And you can do just that in the section "How Acidic Is That Coffee: The pH Scale," later in this chapter.)

Table 12-3 lists the most common strong acids you're likely to encounter.

Table 12-3	Common Strong Acids
Name	*Formula*
Hydrochloric acid	HCl
Hydrobromic acid	HBr
Hydroiodic acid	HI
Nitric acid	HNO_3
Perchloric acid	$HClO_4$
Sulfuric acid (first ionization only)	H_2SO_4

Sulfuric acid is called a *diprotic* acid. It can donate 2 protons, but only the first ionization goes 100 percent. The other acids listed in Table 12-3 are *monoprotic* acids, because they donate only one proton.

Strong bases

You'll normally see only one strong base, and that's the hydroxide ion, OH⁻. Calculating the hydroxide ion concentration is really straightforward. Suppose that you have a 1.5 M (1.5 mol/L) NaOH solution. The sodium hydroxide, a salt, completely *dissociates* (breaks apart) into ions:

$NaOH \rightarrow Na^+ + OH^-$
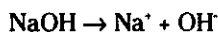
If you start with 1.5 mol/L NaOH, then you have the same concentration of ions:

$[Na^+] = 1.5$ and $[OH^-] = 1.5$

Weak acids

Suppose that you dissolve acetic acid (CH_3COOH) in water. It reacts with the water molecules, donating a proton and forming hydronium ions. It also establishes equilibrium, where you have a significant amount of unionized acetic acid. (In reactions that go to completion, the reactants are completely used up creating the products. But in equilibrium systems, two exactly opposite chemical reactions — one on each side of the reaction arrow — are occurring at the same place, at the same time, with the same speed of reaction. For a discussion of equilibrium systems, see Chapter 8.)

If you want to see whether a person is a chemist, ask him to pronounce *unionized.* A chemist pronounces it *un-ionized,* meaning "not ionized." Everyone else pronounces it *union-ized,* meaning "being part of a union."

The acetic acid reaction with water looks like this:

$$CH_3COOH(l) + H_2O(l) \leftrightarrow CH_3COO^- + H_3O^+$$

The acetic acid that you add to the water is only partially ionized. In the case of acetic acid, about 5 percent ionizes, while 95 percent remains in the molecular form. The amount of hydronium ion that you get in solutions of acids that don't ionize completely is much less than it is with a strong acid. Acids that only partially ionize are called *weak acids.*

Calculating the hydronium ion concentration in weak acid solutions isn't as straightforward as it is in strong solutions, because not all of the weak acid that dissolves initially has ionized. In order to calculate the hydronium ion concentration, you must use the equilibrium constant expression for the weak acid. Chapter 8 covers the K_{eq} expression that represents the equilibrium system. For weak acid solutions, you use a modified equilibrium constant expression called the K_a — the *acid ionization constant.* Take a look at the generalized ionization of some weak acid HA:

$$HA + H_2O \leftrightarrow A^- + H_3O^+$$
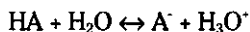
The K_a expression for this weak acid is

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

Note that the *[HA]* represents the molar concentration of HA *at equilibrium,* not initially. Also, note that the concentration of water doesn't appear in the K_a expression, because there's so much that it actually becomes a constant incorporated into the K_a expression.

Now go back to that acetic acid equilibrium. The K_a for acetic acid is 1.8×10^{-5}. The K_a expression for the acetic acid ionization is

$$K_a = 1.8 \times 10^{-5} = \frac{[H_3O^+][CH_3COO^-]}{[CH_3COOH]}$$

You can use this K_a when calculating the hydronium ion concentration in, say, a 2.0 M solution of acetic acid. You know that the initial concentration of acetic acid is 2.0 M. You know that a little bit has ionized, forming a little hydronium ion and acetate ion. You also can see from the balanced reaction that for every hydronium ion that's formed, an acetate ion is also formed — so their concentrations are the same. You can represent the amount of $[H_3O^+]$ and $[CH_3COO^-]$ as x, so

$[H_3O^+] = [CH_3COO^-] = x$

In order to produce the x amount of hydronium and acetate ion, the same amount of ionizing acetic acid is required. So you can represent the amount of acetic acid remaining at equilibrium as the amount you started with, 2.0 M, minus the amount that ionizes, x:

$[CH_3COOH] = 2.0 - x$

For the vast majority of situations, you can say that x is very small in comparison to the initial concentration of the weak acid. So you can say that $2.0 - x$ is approximately equal to 2.0. This means that you can often approximate the equilibrium concentration of the weak acid with its initial concentration. The equilibrium constant expression now looks like this:

$$K_a = 1.8 \times 10^{-5} = \frac{[X][X]}{[2.0]} = \frac{[X]^2}{[2.0]}$$

At this point, you can solve for x, which is the $[H_3O^+]$:

$$\left(1.8 \times 10^{-5}\right)[2.0] = [X]^2$$
$$\sqrt{3.6 \times 10^{-5}} = [X] = [H_3O^+]$$
$$6.0 \times 10^{-3} = [H_3O^+]$$

Table 12-3 shows some common strong acids. Most of the other acids you encounter are weak.

One way to distinguish between strong and weak acids is to look for an acid ionization constant (K_a) value. If the acid has a K_a value, then it's weak.

Weak bases

Weak bases also react with water to establish an equilibrium system. Ammonia is a typical weak base. It reacts with water to form the ammonium ion and the hydroxide ion:

$$NH_3(g) + H_2O(l) \leftrightarrow NH_4^+ + OH^-$$

Like a weak acid, a weak base is only partially ionized. There's a modified equilibrium constant expression for weak bases — the K_b. You use it exactly the same way you use the K_a (see "Weak acids" for the details) except you solve for the $[OH^-]$.

Give me that proton: Bronsted-Lowery acid-base reactions

With the Arrhenius theory, acid-base reactions are neutralization reactions. With the Bronsted-Lowery theory, acid-base reactions are a competition for a proton. For example, take a look at the reaction of ammonia with water:

$$NH_3(g) + H_2O(l) \leftrightarrow NH_4^+ + OH^-$$

Ammonia is a base (it accepts the proton), and water is an acid (it donates the proton) in the forward (left to right) reaction. But in the reverse reaction (right to left), the ammonium ion is an acid and the hydroxide ion is a base. If water is a stronger acid than the ammonium ion, then there is a relatively large concentration of ammonium and hydroxide ions at equilibrium. If, however, the ammonium ion is a stronger acid, much more ammonia than ammonium ion is present at equilibrium.

Bronsted and Lowery said that an acid reacts with a base to form conjugate acid-base pairs. Conjugate acid base pairs differ by a single H+. NH_3 is a base, for example, and NH_4^+ is its conjugate acid. H_2O is an acid in the reaction between ammonia and water, and OH⁻ is its conjugate base. In this reaction, the hydroxide ion is a strong base and ammonia is a weak base, so the equilibrium is shifted to the left — there's not much hydroxide at equilibrium.

Make up your mind: Amphoteric water

When acetic acid reacts with water, water acts as a base, or a proton acceptor. But in the reaction with ammonia (see the preceding section), water acts as an acid, or a proton donor. Water can act as either an acid or a base, depending on what it's combined with. Substances that can act as either an acid or a

base are called *amphoteric.* If you put water with an acid, it acts as a base, and vice versa.

But can it react with itself? Yes, it can. Two water molecules can react with each other, with one donating a proton and the other accepting it:

$$H_2O(l) + H_2O(l) \leftrightarrow H_3O^+ + OH^-$$
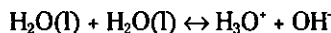
This reaction is an equilibrium reaction. A modified equilibrium constant, called the K_w (which stands for *water dissociation constant*) is associated with this reaction. The K_w has a value of 1.0×10^{-14} and has the following form:

$$1.0 \times 10^{-14} = K_w = [H_3O^+][OH^-]$$

In pure water, the $[H_3O^+]$ equals the $[OH^-]$ from the balanced equation, so $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$. The K_w value is a constant. This value allows you to convert from $[H^+]$ to $[OH^-]$, and vice versa, in *any* aqueous solution, not just pure water. In aqueous solutions, the hydronium ion and hydroxide ion concentrations are rarely going to be equal. But if you know one of them, the K_w allows you to figure out the other one.

Take a look at the 2.0 M acetic acid solution problem in the section "Weak acids," earlier in this chapter. You find that the $[H_3O^+]$ is 6.0×10^3. Now you have a way to calculate the $[OH^-]$ in the solution by using the K_w relationship:

$$K_w = 1.0 \times 10^{-14} = [H_3O^+][OH^-]$$
$$1.0 \times 10^{-14} = [6.0 \times 10^{-3}][OH^-]$$
$$1.0 \times 10^{-14}/6.0 \times 10^{-3} = [OH^-]$$
$$1.7 \times 10^{-12} = [OH^-]$$

An Old Laxative and Red Cabbage: Acid-Base Indicators

Indicators are substances (organic dyes) that change color in the presence of an acid or base. You may be familiar with an acid-base indicator plant — the hydrangea. If it's grown in acidic soil, it turns pink; if it's grown in alkaline soil, it turns blue. Another common substance that acts as a good acid-base indicator is red cabbage. I have my students chop some up and boil it (most of them really *love* this part). They then take the liquid that is left over and use it to test substances. When mixed with an acid, the liquid turns pink; when mixed with a base, it turns green. In fact, if you take some of this liquid, make it slightly basic, and then exhale your breath into it through a straw, the solution eventually turns blue, indicating that the solution has turned slightly

acidic. The carbon dioxide in your breath reacts with the water, forming carbonic acid:

$$CO_2(g) + H_2O(l) \leftrightarrow H_2CO_3(aq)$$

Carbonated beverages are slightly acidic due to this reaction. Carbon dioxide is injected into the liquid to give it fizz. A little of this carbon dioxide reacts with the water to form carbonic acid. This reaction also explains why rainwater is slightly acidic. It absorbs carbon dioxide from the atmosphere as it falls to earth.

In chemistry, indicators are used to indicate the presence of an acid or base. Chemists have many indicators that change at slightly different pHs. (You've probably heard the term *pH* used in various contexts. Me, I even remember it being used to sell deodorant on TV. If you want to know what it actually stands for, check out the section "How Acidic Is That Coffee: The pH Scale.") Two indicators are used most often:

- Litmus paper
- Phenolphthalein

Good old litmus paper

Litmus is a substance that is extracted from a type of lichen and absorbed into porous paper. (In case you're scheduled for a hot game of *Trivial Pursuit* this weekend, a *lichen* is a plant — found in the Netherlands — that's made up of an alga and a fungus that are intimately living together and mutually benefiting from the relationship. Sounds kind of sordid to me.) There are three different types of litmus — red, blue, and neutral. Red litmus is used to test for bases, and blue litmus is used to test for acids, while neutral litmus can be used to test for both. If a solution is acidic, both blue and neutral litmus will turn red. If a solution is basic, both red and neutral litmus will turn blue. Litmus paper is a good, quick test for acids and bases. And you don't have to put up with the smell of boiling cabbage.

Phenolphthalein: Helps keep you regular

Phenolphthalein (pronounced *fe-nul-tha-Leen*) is another commonly used indicator. Until a few years ago, phenolphthalein was used as the active ingredient in a popular laxative. In fact, I used to extract the phenolphthalein from the laxative by soaking it in either rubbing alcohol or gin (being careful not to drink it). I'd then use this solution as an indicator.

Phenolphthalein is clear and colorless in an acid solution and pink in a basic solution. It's commonly used in a procedure called a *titration*, where

the concentration of an acid or base is determined by its reaction with a base or acid of known concentration.

Suppose, for example, that you want to determine the molar concentration of an HCl solution. First, you place a known volume (say, 25.00 milliliters measured accurately with a pipette) in an Erlenmeyer flask (that's just a flat-bottomed, conical-shaped flask) and add a couple drops of phenolphthalein solution. Because you're adding the indicator to an acidic solution, the solution in the flask remains clear and colorless. You then add small amounts of a standardized sodium hydroxide solution of known molarity (for example, 0.100 M) with a buret. (A *buret* is a graduated glass tube with a small opening and a stopcock, which helps you measure precise volumes of solution.) You keep adding base until the solution turns the faintest shade of pink detectable. I call this the *endpoint* of the titration, the point in which the acid has been exactly neutralized by the base. Figure 12-2 shows the titration setup.
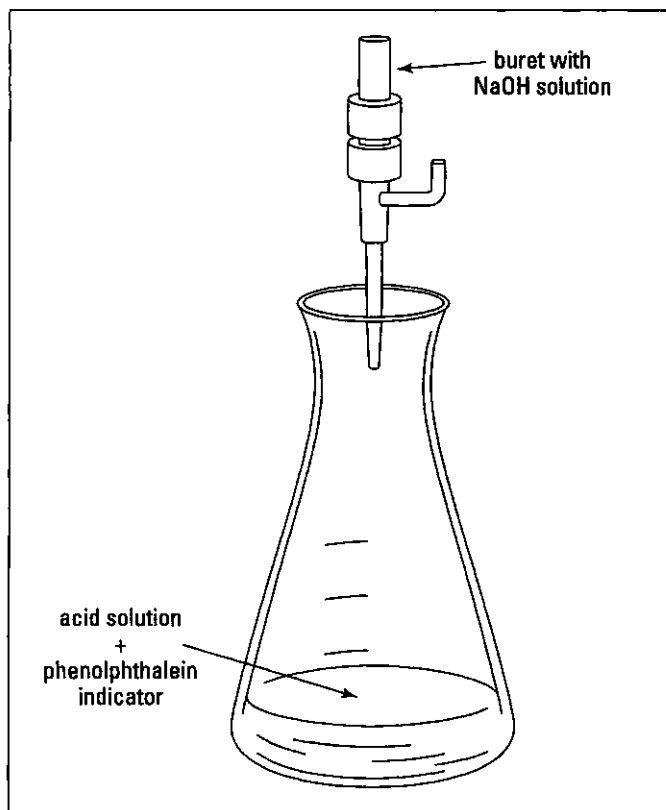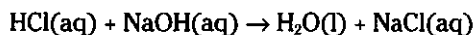

buret with NaOH solution

acid solution + phenolphthalein indicator

Figure 12-2: Titration of an acid with a base.

Suppose that it takes 35.50 milliliters of the 0.100 M NaOH to reach the endpoint of the titration of the 25.00 milliliters of the HCl solution. Here's the reaction:

$$HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$$

From the balanced equation, you can see that the acid and base react in a 1:1 mole ratio. So if you can calculate the moles of bases added, you'll also know the number of moles of HCl present. Knowing the volume of the acid solution then allows you to calculate the molarity (note that you convert the milliliters to liters so that your units cancel nicely):

$$\frac{0.100 \, \cancel{mol \, Na} \, OH}{\cancel{L}} \times \frac{0.03550 \cancel{L}}{1} \times \frac{1 \, mol \, HCl}{1 \, \cancel{mol \, Na} \, OH} \times \frac{1}{0.02500 \cancel{L}} = 0.142 \, M \, HCl$$

The titration of a base with a standard acid solution (one of known concentration) can be calculated in exactly the same way, except the endpoint is the first disappearance of the pink color.

How Acidic Is That Coffee: The pH Scale

The amount of acidity in a solution is related to the concentration of the hydronium ion in the solution. The more acidic the solution is, the larger the concentration of the hydronium ion. In other words, a solution in which the $[H_3O^+]$ equals 1.0×10^{-2} is more acidic than a solution in which the $[H_3O^+]$ equals 1.0×10^{-7}. The *pH scale,* a scale based on the $[H_3O^+]$, was developed to more easily tell, at a glance, the relative acidity of a solution. *pH* is defined as the negative logarithm (abbreviated as *log*) of the $[H_3O^+]$. Mathematically, it looks like this:

$$pH = -\log [H_3O^+]$$

Based on the water dissociation constant, K_w (see "Make up your mind: Amphoteric water," earlier in this chapter), in pure water the $[H_3O^+]$ equals 1.0×10^{-7}. Using this mathematical relationship, you can calculate the pH of pure water:

$$pH = -\log [H_3O^+]$$
$$pH = -\log [1.0 \times 10^{-7}]$$
$$pH = -[-7]$$
$$pH = 7$$

The pH of pure water is 7. Chemists call this point on the pH scale *neutral.* A solution is called *acidic* if it has a larger $[H_3O^+]$ than water and a smaller pH value than 7. A *basic* solution has a smaller $[H_3O^+]$ than water and a larger pH value than 7.

The pH scale really has no end. You can have a solution of pH that registers less than 0. (A 10 M HCl solution, for example, has a pH of –1.) However, the 0 to 14 range is a convenient range to use for weak acids and bases and for dilute solutions of strong acids and bases. Figure 12-3 shows the pH scale.
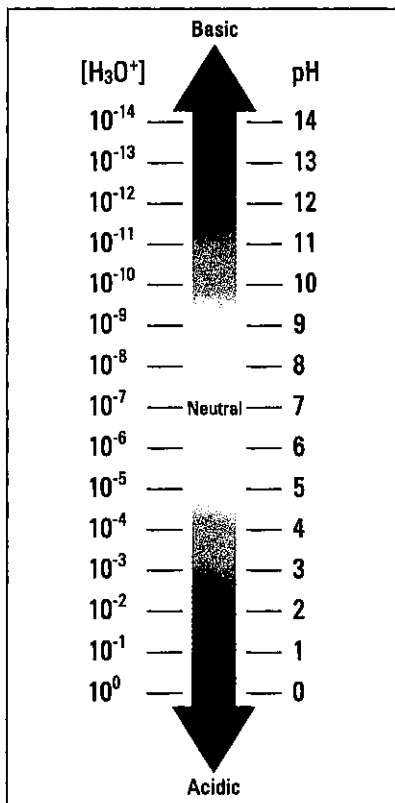
The $[H_3O^+]$ of a 2.0 M acetic acid solution is 6.0×10^{-3}. Looking at the pH scale, you see that this solution is acidic. Now calculate the pH of this solution:

pH = – log $[H_3O^+]$

pH = – log $[6.0 \times 10^{-3}]$

pH = – [–2.22]

pH = 2.22

In the section "Make up your mind: Amphoteric water," I explain that the K_w expression enables you to calculate the $[H_3O^+]$ if you have the $[OH^-]$. Another equation, called the *pOH*, can be useful in calculating the pH of a solution. The pOH is the negative logarithm of the $[OH^-]$. You can calculate the pOH of

a solution just like the pH by taking the negative log of the hydroxide ion concentration. If you use the K_w expression and take the negative log of both sides, you get 14 = pH + pOH. This equation makes it easy to go from pOH to pH.

Just as you can you convert from $[H_3O^+]$ to pH, you can also go from pH to $[H_3O^+]$. To do this, you use what's called the *antilog relationship,* which is

$[H_3O^+] = 10^{-pH}$

Human blood, for example, has a pH of 7.3. Here's how you calculate the $[H_3O^+]$ from the pH of blood:

$[H_3O^+] = 10^{-pH}$

$[H_3O^+] = 10^{-7.3}$

$[H_3O^+] = 5.01 \times 10^{-8}$

The same procedure can be used to calculate the $[OH^-]$ from the pOH.

Substances commonly found in our surroundings cover a wide range of pH values. Table 12-4 lists some common substances and their pH values.

Table 12-4	Average pH Values of Some Common Substances
Substance	*PH*
Oven cleaner	13.8
Hair remover	12.8
Household ammonia	11.0
Milk of magnesia	10.5
Chlorine bleach	9.5
Seawater	8.0
Human blood	7.3
Pure water	7.0
Milk	6.5
Black coffee	5.5
Soft drinks	3.5
Aspirin	2.9
Vinegar	2.8

(continued)

Table 12-4 *(continued)*	
Substance	*PH*
Lemon juice	2.3
Auto battery acid	0.8

Human blood has a pH of around 7.3. There's a narrow range in which blood pH can change and still sustain life, about +/– 0.2 pH units. Many things in our environment, such as foods and hyperventilation, can act to change the pH of our blood. Buffers help to regulate blood pH and keep it in the 7.1 to 7.5 range.

Buffers: Controlling pH

Buffers, or *buffer solutions* as they're sometimes called, resist a change in pH caused by the addition of acids or bases. Obviously, the buffer solution must contain something that reacts with an acid — a base. Something else in the buffer solution reacts with a base — an acid. There are, in general, two types of buffers:

- ✔ Mixtures of weak acids and bases
- ✔ Amphoteric species

The mixtures of weak acids and bases may be conjugate acid-base pairs (such as H_2CO_3/HCO_3^-) or nonconjugate acid-base pairs (such as NH_4+/CH_3COO^-). (For more info about conjugate acid-base pairs, see "Give me that proton: Bronsted-Lowery acid-base reactions," earlier in this chapter.)

In the body, conjugate acid-base pairs are more common. In the blood, for example, the carbonic acid/bicarbonate pair helps to control the pH. This buffer can be overcome, though, and some potentially dangerous situations can arise. If a person exercises strenuously, lactic acid from the muscles is released into the bloodstream. If there's not enough bicarbonate ion to neutralize the lactic acid, the blood pH drops, and the person is said to be in *acidosis.* Diabetes may also cause acidosis. On the other hand, if a person hyperventilates (breathes too fast), she breathes out too much carbon dioxide. The carbonic acid level in the blood is reduced, causing the blood to become too basic. This condition, called *alkalosis,* can be very serious.

Amphoteric species may also act as buffers by reacting with an acid or a base. (For an example of an amphoteric species, see "Make up your mind: Amphoteric water," earlier in this chapter) The bicarbonate ion (HCO_3^-) and the monohydrogen phosphate ion (HPO_4^{-2}) are amphoteric species

that neutralize both acids and bases. Both of these ions are also important in controlling the blood's pH.

Antacids: Good, Basic Chemistry

Go to any drugstore or grocery store and look at the shelves upon shelves of antacids. They represent acid-base chemistry in action!

The stomach secretes hydrochloric acid in order to activate certain enzymes (biological catalysts) in the digestion process. But sometimes the stomach produces too much acid, or the acid makes its way up into the esophagus (leading to heartburn), making it necessary to neutralize the excess acid with — you guessed it — a base. The basic formulations that are sold to neutralize this acid are called *antacids*. Antacids include the following compounds as active ingredients:

- **Bicarbonates** — $NaHCO_3$ and $KHCO_3$
- **Carbonates** — $CaCO_3$ and $MgCO_3$
- **Hydroxides** — $Al(OH)_3$ and $Mg(OH)_2$

Acids with bad press: An introduction to acid rain

Over the past few years, acid rain has emerged as a great environmental problem. Natural rainwater is somewhat acidic (around pH 5.6) due to the absorption of carbon dioxide from the atmosphere and the creation of carbonic acid. However, when acid rain is mentioned in the press, it usually refers to rain in the pH 3 to 3.5 range.

The two major causes of acid rain are automotive and industrial pollution. In the automobile's internal combustion engine, nitrogen in the air is oxidized to various oxides of nitrogen. These nitrogen oxides, when released into the atmosphere, react with water vapor to form nitric acid (HNO_3).

In fossil fuel power plants, oxides of sulfur are formed from the burning of the sulfur impurities commonly found in coal and petroleum. These oxides of sulfur, if released into the atmosphere, combine with water vapor to form both sulfuric and sulfurous acids (H_2SO_4 and H_2SO_3). Oxides of nitrogen are also produced in these power plants.

These acids fall to earth in the rain and cause a multitude of problems. They dissolve the calcium carbonate of marble statues and monuments. They decrease the pH of lake water to such a degree that fish can no longer live in the lakes. They cause whole forests to die or become stunted. They react with the metals in cars and buildings.

Industrial controls have been somewhat effective in reducing the problem, but it's still a major environmental issue. (See Chapter 18 for more info about acid rain.)

Trying to select the "best" antacid for occasional use can be complicated. Certainly price is a factor, but the chemical nature of the bases can also be a factor. For example, individuals with high blood pressure may want to avoid antacids containing sodium bicarbonate because the sodium ion tends to increase blood pressure. Individuals concerned about loss of calcium from the bones, or *osteoporosis,* may want to use an antacid containing calcium carbonate. However, both calcium carbonate and aluminum hydroxide can cause constipation if used in large doses. On the other hand, large doses of both magnesium carbonate and magnesium hydroxide can act as laxatives. Selecting an antacid can really be a balancing act!

Chapter 13

Balloons, Tires, and Scuba Tanks: The Wonderful World of Gases

* *

In This Chapter

▷ Accepting the Kinetic Molecular Theory of Gases

▷ Figuring out pressure

▷ Coming to understand the gas laws

* *

*G*ases are all around you. Because gases are generally invisible, you may not think of them directly, but you're certainly aware of their properties. You breathe a mixture of gases that you call air. You check the pressure of your automobile tires, and you check the atmospheric pressure to see if a storm is coming. You burn gases in your gas grill and lighters. You fill birthday balloons for your loved ones.

The properties of gases and their interrelationships are important to you. Is there enough pressure in my tires? How big is that balloon going to be? Is there enough air in my scuba tanks? The list goes on and on.

In this chapter, I introduce you to gases at both the microscopic and macroscopic levels. I show you one of science's most successful theories — the Kinetic Molecular Theory of Gases. And I explain the macroscopic properties of gases and show you the important interrelationships among them. I also show you how these relationships come into play in reaction stoichiometry. This chapter is a real gas!

Microscopic View of Gases: The Kinetic Molecular Theory

A theory is useful to scientists if it describes the physical system they're examining and allows them to predict what will happen if they change

some variable. The Kinetic Molecular Theory of Gases does just that. It has limitations — all theories do — but it's one of the most useful theories in chemistry. This section describes the theory's basic *postulates* — assumptions, hypotheses, axioms (pick your favorite word) you can accept as being true.

✔ **Postulate 1: Gases are composed of tiny particles, either atoms or molecules.**

Unless you're discussing matter at greatly elevated temperatures, the particles referred to as gases tend to be relatively small. The more massive particles clump together to form liquids or even solids. So gases are normally small with relatively low atomic and molecular weights.

✔ **Postulate 2: The gas particles are so small when compared to the distances between them that the volume the gas particles themselves take up is negligible and is assumed to be zero.**

These gas particles do take up some volume — that's one of the properties of matter. But if the gas particles are small (which they are), and there aren't many of them in a container, you say that their volume is negligible when compared to the volume of the container or the space between the gas particles. This explains why gases are compressible. There's a lot of space between the gas particles, so they can be squeezed together. This isn't true in solids and liquids, where the particles are *much* closer together. (Chapter 2 covers the various states of matter, if you want to have a look-see at the differences between solids, liquids, and gases.)

The concept of a quantity being *negligible* is used a lot in chemistry. An example is in Chapter 12 where I use the acid ionization constant (K_a) of a weak acid, ignoring the amount of weak acid that has ionized compared to the initial concentration of the acid.

In the real world, I like to compare this negligible concept to finding a dollar in the street. If you have no money at all, then that dollar represents a sizable quantity of cash (perhaps your next meal). But if you're a multimillionaire, then that dollar doesn't represent much at all. It may as well be a piece of scrap paper. You may not even pick it up. (I really can't imagine being that rich.) Its value is negligible when compared to the rest of your wealth. This is what I'm saying with respect to the volume of the gas particles — sure, they have a volume, but it's so small that it's insignificant when compared with the distance between the gas particles and the volume of the container.

✔ **Postulate 3: The gas particles are in constant random motion, moving in straight lines and colliding with the inside walls of the container.**

The gas particles are always moving in a straight-line motion. (Gases have a higher kinetic energy — energy of motion — associated with them when compared to solids or liquids; see Chapter 2.) They continue to move in these straight lines until they collide with something — either

with each other or with the inside walls of the container. The particles also all move in different directions, so the collisions with the inside walls of the container tend to be uniform over the entire inside surface. You can observe this uniformity by simply blowing up a balloon. The balloon is spherical because the gas particles are hitting all points of the inside walls the same. The collision of the gas particles with the inside walls of the container is called *pressure*. The idea that the gas particles are in constant, random, straight-line motion explains why gases uniformly mix if put in the same container. It also explains why, when you drop a bottle of cheap perfume at one end of the room, the people at the other end of the room are able to smell it right away.

✔ **Postulate 4: The gas particles are assumed to have negligible attractive or repulsive forces between each other.**

In other words, the gas particles are assumed to be totally independent, neither attracting nor repelling each other. That said, it's hair-splitting time. This statement is actually false; if it were true, chemists would never be able to liquefy a gas, which they can. But the reason you can accept this assumption as true is that the attractive and repulsive forces are generally so small that they can safely be ignored. The assumption is most valid for nonpolar gases, such as hydrogen and nitrogen, because the attractive forces involved are London forces. However, if the gas molecules are polar, as in water and HCl, this assumption can become a problem. (Turn to Chapter 7 for the scoop about London forces and polar things — all related to the attraction between molecules.)

✔ **Postulate 5: The gas particles may collide with each other. These collisions are assumed to be elastic, with the total amount of kinetic energy of the two gas particles remaining the same.**

Not only do the gas particles collide with the inside walls of the container, they also collide with each other. If they hit each other, no kinetic energy is lost, but kinetic energy may be transferred from one gas particle to the other. For example, imagine two gas particles — one moving fast and the other moving slow — colliding. Kinetic energy is transferred from the faster particle to the slower particle. The one that's moving slow bounces off the faster particle and moves away at a greater speed than before, while the one that's moving fast bounces off the slower particle and moves away at a slower speed. The total amount of kinetic energy remains the same, but one gas particle loses energy and the other gains energy. This is the principle behind pool — you transfer kinetic energy from your moving pool stick to the cue ball to the ball you're aiming at.

✔ **Postulate 6: The Kelvin temperature is directly proportional to the *average* kinetic energy of the gas particles.**

The gas particles aren't all moving with the same amount of kinetic energy. A few are moving relatively slow and a few are moving very fast, but most are somewhere in between these two extremes. Temperature, particularly as measured using the Kelvin temperature scale, is directly

related to the *average* kinetic energy of the gas. If you heat the gas so that the Kelvin temperature (K) increases, the average kinetic energy of the gas also increases. (To calculate the Kelvin temperature, add 273 to the Celsius temperature: $K = {}^{\circ}C + 273$. Temperature scales and average kinetic energy are all tucked neatly into Chapter 2.)

A gas that obeys all the postulates of the Kinetic Molecular Theory is called an *ideal gas.* Obviously, no real gas obeys the assumptions made in the second and fourth postulates *exactly.* But a nonpolar gas at high temperatures and low pressure (concentration) approaches ideal gas behavior.

I'm Under Pressure — Atmospheric Pressure, That Is

Although you're not in a container, the gas molecules of the atmosphere are constantly hitting you, your books, your computer, and everything, and exerting a force called *atmospheric pressure.* Atmospheric pressure is measured using an instrument called a *barometer.*

Measuring atmospheric pressure: The barometer

If you get a complete weather report, the atmospheric pressure is normally included. You can get an idea about changes in the weather by observing whether the atmospheric pressure is rising or falling. The atmospheric pressure is measured using a barometer, and Figure 13-1 shows the components of one.

A barometer is composed of a long glass tube that's closed at one end and totally filled with a liquid. You can use water, but the tube would have to be *very* long (about 35 feet long), making for a rather inconvenient barometer. So it makes more sense to use mercury because it's a very dense liquid. The tube filled with mercury is inverted into an open container of mercury so that the open end of the tube is under the surface of the mercury in the container. A couple of things now take place. The force of gravity pulls the mercury in the tube *down,* causing it to drain out into the container. The weight of the gases in the atmosphere exert a force downward on the mercury in the open container and force it *up* into the tube. Sooner or later, these forces balance, and the mercury in the tube comes to rest at a certain height from the top of the pool of mercury in the container. The greater the pressure of the atmosphere, the higher the mercury column that can be measured in the tube; the

lower the pressure of the atmosphere (for example, at the top of a tall mountain), the shorter the column. At sea level, the column is 760 millimeters high, the so-called normal *atmospheric pressure.*
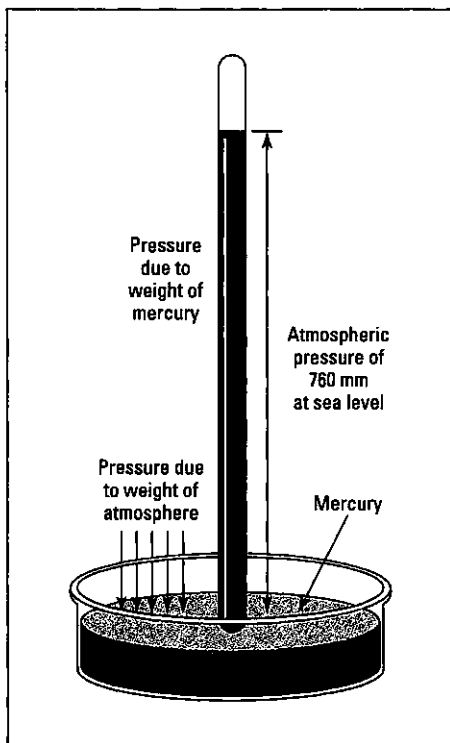

Figure 13-1:
A barometer.

Atmospheric pressure can be expressed a number of different ways. It can be expressed in millimeters of mercury (mm Hg); atmospheres (atm), a unit of pressure where 1 atmosphere is the pressure at sea level; torr, a unit of pressure where 1 torr equals 1 millimeter of mercury; pounds per square inch (psi); pascals (Pa), a unit of pressure where 1 pascal equals 1 newton per square meter (don't worry about what a newton is; just trust me that this stuff is a way to express pressure); or kilopascals (kPa), where 1 kilopascal equals 1,000 pascals.

So you can express the atmospheric pressure at sea level as

760 mm Hg = 1 atm = 760 torr = 14.69 psi = 101,325 Pa = 101.325 kPa

Note that sometimes you also hear atmospheric pressure reported as inches of mercury (1 atm = 29.921 in Hg). In this book, I primarily use atmospheres and torr, with an occasional millimeter of Hg. Variety is the spice of life.

Measuring confined gas pressure: The manometer

You can measure the pressure of a gas confined in a container by using an apparatus called a *manometer*. (It's pronounced *man-AH-muh-ter*, not *man-o-meter*.) Figure 13-2 shows the components of a manometer.
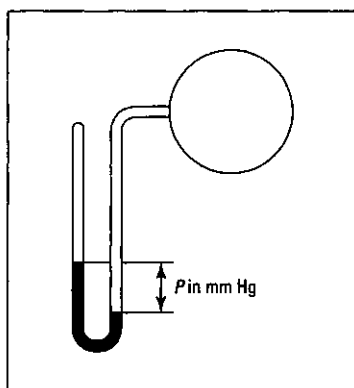
Figure 13-2: The manometer.

P in mm Hg

A manometer is kind of like a barometer. The container of gas is attached to a U-shaped piece of glass tubing that's partially filled with mercury and sealed at the other end. Gravity pulls down the mercury column at the closed end. The mercury is then balanced by the pressure of the gas in the container. The difference in the two mercury levels represents the amount of gas pressure.

Gases Obey Laws, Too — Gas Laws

Various scientific laws describe the relationships between four of the important physical properties of gases:

- ✔ Volume
- ✔ Pressure
- ✔ Temperature
- ✔ Amount

This section covers those various laws. Boyle's, Charles's, and Gay-Lussac's Laws each describe the relationship between two properties while keeping the other two properties constant. (In other words, you take two properties,

change one, and then see its effect on the second — while keeping the remaining properties constant.) Another law — a combo of Boyle's, Charles's, and Gay-Lussac's individual laws — enables you to vary more than one property at a time.

But that combo law doesn't let you vary the physical property of amount. Avogadro's Law, however, does. And there's even an ideal gas law, which lets you take into account variations in all four physical properties.

Yes, this section is so chock-full of laws, it'll probably *give* you gas just trying to digest it.

Boyle's Law: Nothing to do with boiling

Boyle's Law, named after Robert Boyle, a seventeenth-century English scientist, describes the pressure-volume relationship of gases if the temperature and amount are kept constant. Figure 13-3 illustrates the pressure-volume relationship using the Kinetic Molecular Theory.
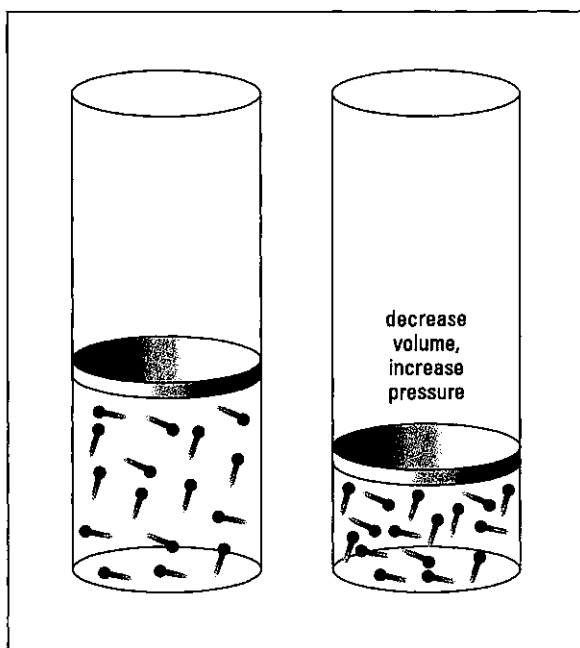

Figure 13-3: Pressure-volume relationship of gases — Boyle's Law.

decrease volume, increase pressure

The left-hand cylinder in the figure contains a certain volume of gas at a certain pressure. (*Pressure* is the collision of the gas particles with the inside walls of the container.) When the volume is decreased, the same number of

gas particles are now contained in a much smaller volume and the number of collisions increases significantly. Therefore, the pressure is greater.

Boyle's Law states that there's an inverse relationship between the volume and pressure. As the volume decreases, the pressure increases, and vice versa.

Boyle determined that the product of the pressure and the volume is a constant (k):

$$PV = k$$

Now consider a case where you have a gas at a certain pressure (P_1) and volume (V_1). If you change the volume to some new value (V_2), the pressure also changes to a new value (P_2). You can use Boyle's Law to describe both sets of conditions:

$$P_1V_1 = k$$
$$P_2V_2 = k$$

The constant, k, is going to be the same in both cases. So you can say

$$P_1V_1 = P_2V_2 \quad \text{(with temperature and amount constant)}$$

This equation is another statement of Boyle's Law — and it's really a more useful one, because you'll normally deal with changes in pressure and volume. If you know three of the preceding quantities, you can calculate the fourth one. For example, suppose that you have 5.00 liters of a gas at 1.00 atm pressure, and then you decrease the volume to 2.00 liters. What's the new pressure?

To find the answer, use the following setup:

$$P_1V_1 = P_2V_2$$

Substituting 1.00 atm for P_1, 5.00 liters for V_1, and 2.00 liters for V_2, you get

$$(1.00 \text{ atm})(5.00 \text{ liters}) = P_2(2.00 \text{ liters})$$

Now solve for P_2:

$$(1.00 \text{ atm})(5.00 \text{ liters})/2.00 \text{ liters} = P_2 = 2.50 \text{ atm}$$

The answer makes sense, because you decreased the volume and the pressure increased, which is exactly what Boyle's law says.

Charles's Law: Don't call me Chuck

Charles's Law, named after Jacques Charles, a nineteenth-century French chemist, has to do with the relationship between volume and temperature, keeping the pressure and amount constant. You run across situations dealing with this relationship in everyday life, especially in terms of the heating and cooling of balloons.

Figure 13-4 shows the temperature-volume relationship.

Figure 13-4:
The
temperature-
volume
relationship
of gases —
Charles's
Law.
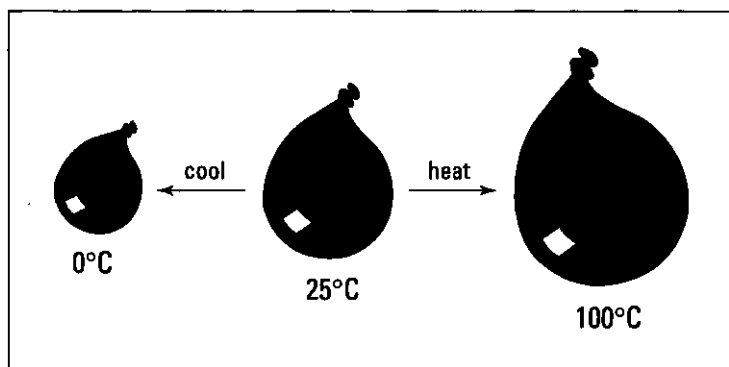
Look at the balloon in the middle of Figure 13-4. What do you think would happen to the balloon if you placed it in the freezer? It'd get smaller. Inside the freezer, the external pressure, or atmospheric pressure, is the same, but the gas particles inside the balloon aren't moving as fast, so the volume shrinks to keep the pressure constant. If you heat the balloon, the balloon expands and the volume increases. This is a *direct relationship* — as the temperature increases, the volume increases, and vice versa.

Jacques Charles developed the mathematical relationship between temperature and volume. He also discovered that you must use the Kelvin (K) temperature when working with gas law expressions.

In gas law calculations, the Kelvin temperature *must* be used.

Charles's Law says that the volume is directly proportional to the Kelvin temperature. Mathematically, the law looks like this:

$$V = bT \text{ or } V/T = b \qquad (\text{where } b \text{ is a constant})$$

If the temperature of a gas with a certain volume (V_1) and Kelvin temperature (T_1) is changed to a new Kelvin temperature (T_2), the volume also changes (V_2).

$$V_1/T_1 = b \qquad V_2/T_2 = b$$

The constant, b, is the same, so

$V_1/T_1 = V_2/T_2$ (with the pressure and amount of gas held constant and temperature expressed in K)

If you have three of the quantities, you can calculate the fourth. For example, suppose you live in Alaska and are outside in the middle of winter, where the temperature is –23 degrees Celsius. You blow up a balloon so that it has a volume of 1.00 liter. You then take it inside your home, where the temperature is a toasty 27 degrees Celsius. What's the new volume of the balloon?

First, convert your temperatures to Kelvin by adding 273 to the Celsius temperature:

–23 °C + 273 = 250 K   (outside)

27 °C + 273 = 300 K   (inside)

Now you can solve for V_2, using the following setup:

$V_1/T_1 = V_2/T_2$

Multiply both sides by T_2 so that V_2 is on one side of the equation by itself:

$[V_1 T_2]/T_1 = V_2$

Then substitute the values to calculate the following answer:

$[(1.00 \text{ liter})(300 \text{ K})]/250 \text{ K} = V_2 = 1.20 \text{ liters}$

It's a reasonable answer, because Charles's Law says that if you increase the Kelvin temperature, the volume increases.

Gay-Lussac's Law

Gay-Lussac's Law, named after the 19th century French scientist Joseph-Louis Gay-Lussac deals with the relationship between the pressure and temperature of a gas if its volume and amount are held constant. Imagine, for example, that you have a metal tank of gas. The tank has a certain volume, and the gas inside has a certain pressure. If you heat the tank, you increase the kinetic energy of the gas particles. So they're now moving much faster, and they're not only hitting the inside walls of the tank more often but also with more force. The pressure has increased.

Gay-Lussac's Law says that the pressure is directly proportional to the Kelvin temperature. Figure 13-5 shows this relationship.

Figure 13-5:
The pressure-temperature relationship of gases — Gay-Lussac's Law.
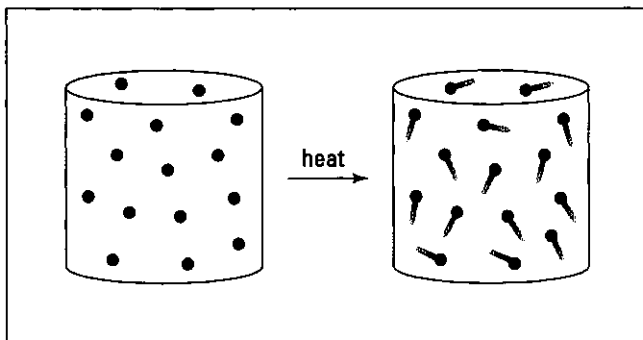
Mathematically, Gay-Lussac's Law looks like this:

$P = kT$ (or $P/T = k$ at constant volume and amount)

Consider a gas at a certain Kelvin temperature and pressure (T_1 and P_1), with the conditions being changed to a new temperature and pressure (T_2 and P_2):

$P_1/T_1 = P_2/T_2$

If you have a tank of gas at 800 torr pressure and a temperature of 250 Kelvin, and it's heated to 400 Kelvin, what's the new pressure?

Starting with $P_1/T_1 = P_2/T_2$, multiply both sides by T_2 so you can solve for P_2:

$[P_1 T_2]/T_1 = P_2$

Now substitute the values to calculate the following answer:

$[(800 \text{ torr})(400 \text{ K})]/250 \text{ K} = P_2 = 1{,}280 \text{ torr}$

This is a reasonable answer because if you heat the tank, the pressure should increase.

The combined gas law

All the preceding examples assume that two properties are held constant and one property is changed to see its effect on a fourth property. But life is rarely that simple. How do you handle situations in which two or even three properties change? You can treat each one separately, but it sure would be nice if you had a way to combine things so that wouldn't be necessary.

Actually, there is a way. You can combine Boyle's Law, Charles's Law, and Gay-Lussac's Law into one equation. Trust me, you don't want me to show

you exactly how it's done, but the end result is called the *combined gas law,* and it looks like this:

$$P_1V_1/T_1 = P_2V_2/T_2$$

Just like in the preceding examples, *P* is the pressure of the gas (in atm, mm Hg, torr, and so on), *V* is the volume of the gas (in appropriate units), and *T* is the temperature (in Kelvin). The *1* and *2* stand for the initial and final conditions, respectively. The amount is still held constant: No gas is added, and no gas escapes. There are six quantities involved in this combined gas law; knowing five allows you to calculate the sixth.

For example, suppose that a weather balloon with a volume of 25.0 liters at 1.00 atm pressure and a temperature of 27 degrees Celsius is allowed to rise to an altitude where the pressure is 0.500 atm and the temperature is –33 degrees Celsius. What's the new volume of the balloon?

Before I show you how to work this problem, do a little reasoning. The temperature is decreasing, so that should cause the volume to decrease (Charles's Law). However, the pressure is also decreasing, which should cause the balloon to expand (Boyle's Law). These two factors are competing, so at this point, you don't know which will win out.

You're looking for the new volume (V_2), so rearrange the combined gas law to obtain the following equation (by multiplying each side by T_2 and dividing each side by P_2, which puts V_2 by itself on one side):

$$[P_1V_1T_2]/[P_2T_1] = V_2$$

Now identify your quantities:

P_1 = 1.00 atm; V_1 = 25.0 liters; T_1 = 27°C + 273 = 300. K

P_2 = 0.500 atm; T_2 = –33°C + 273 = 240. K

Now substitute the values to calculate the following answer:

[(1.00 atm)(25.0 liters)(240. K)]/[(0.500 atm)(300. K)] = V_2 = 40.0 liters

Because the volume increased overall in this case, Boyle's Law had a greater effect than Charles's Law.

Avogadro's Law

The combined gas equation gives you a way to calculate changes involving pressure, volume, and temperature. But you still have the problem of amount to deal with. In order to account for amount, you need to know another law.

Amedeo Avogadro (the same Avogadro that gave us his famous number of particles per mole — see Chapter 10) determined, from his study of gases, that equal volumes of gases at the same temperature and pressure contain equal numbers of gas particles. So Avogadro's Law says that the volume of a gas is directly proportional to the number of moles of gas (number of gas particles) at a constant temperature and pressure. Mathematically, Avogadro's Law looks like this:

$$V = kn \text{ (at constant temperature and pressure)}$$

In this equation, k is a constant and n is the number of moles of gas. If you have a number of moles of gas (n_1) at one volume (V_1), and the moles change due to a reaction (n_2), the volume also changes (V_2), giving you the equation

$$V_1/n_1 = V_2/n_2$$

I'm not going to work any problems with this law, because it's basically the same idea as the other gas laws covered in this chapter.

A very useful consequence of Avogadro's Law is that the volume of a mole of gas can be calculated at any temperature and pressure. An extremely useful form to know when calculating the volume of a mole of gas is *1 mole of any gas at STP occupies 22.4 liters. STP* in this case is not an oil or gas additive. It stands for Standard Temperature and Pressure.

 ✔ **Standard Pressure:** 1.00 atm (760 torr or mm Hg)
 ✔ **Standard Temperature:** 273 K

This relationship between moles of gas and liters gives you a way to convert the gas from a mass to a volume. For example, suppose that you have 50.0 grams of oxygen gas (O_2), and you want to know its volume at STP. You can set up the problem like this (see Chapters 10 and 11 for the nuts and bolts of using moles in chemical equations):

$$\frac{50.0 \text{ g } O_2}{1} \times \frac{1 mol\ O_2}{32.0 g} \times \frac{22.4L}{1 mol\ O_2} = 35.0L$$

You now know that the 50.0 grams of oxygen gas occupies a volume of 35.0 liters at STP. But what if the gas is not at STP? What's the volume of 50.0 grams of oxygen at 2.00 atm and 27.0 degrees Celsius? In the next section, I show you a really easy way of doing this problem. But right now, you can use the combined gas law, because you know the volume at STP:

$$P_1V_1/T_1 = P_2V_2/T_2$$

$P_1 = 1.00$ atm; $V_1 = 35.0$ liters; $T_1 = 273$ K

$P_2 = 2.00$ atm; $T_2 = 300.$ K (27 °C + 273)

Solving for V_2, you calculate the following answer:

$[P_1V_1T_2]/[P_2T_1] = V_2$

$[(1.00 \text{ atm})(35.0 \text{ liters})(300 \text{ K})]/[(2.00 \text{ atm})(273 \text{ K})] = V_2 = 19.2 \text{ liters}$

The ideal gas equation

If you take Boyle's Law, Charles's Law, Gay-Lussac's Law, and Avogadro's Law and throw them into a blender, turn the blender on high for a minute, and then pull them out, you get the *ideal gas equation* — a way of working in volume, temperature, pressure, *and* amount. The ideal gas equation has the following form:

$PV = nRT$

The P represents pressure in atmospheres (atm), the V represents volume in liters (L), the n represents moles of gas, the T represents the temperature in Kelvin (K), and the R represents the ideal gas constant, which is 0.0821 liters atm/K mol.

Using the value of the ideal gas constant, the pressure must be expressed in atm, and the volume must be expressed in liters. You can calculate other ideal gas constants if you really want to use torr and milliliters, for example, but why bother? It's easier to memorize one value for R and then remember to express the pressure and volume in the appropriate units. Naturally, you'll *always* express the temperature in Kelvin when working any kind of gas law problem.

That said, now I want to show you an easy way to convert a gas from a mass to a volume if the gas is not at STP. What's the volume of 50.0 grams of oxygen at 2.00 atm and 27.0 degrees Celsius?

The first thing you have to do is convert the 50.0 grams of oxygen to moles using the molecular weight of O_2:

(50.0 grams)(1 mol/32.0 grams) = 1.562 mol

Now take the ideal gas equation and rearrange it so you can solve for V:

$PV = nRT$

$V = nRT/P$

Add your known quantities to calculate the following answer:

$V = [(1.562 \text{ mol})(0.0821 \text{ liters atm/K mol})(300 \text{ K})]/2.00 \text{ atm} = 19.2 \text{ liters}$

This is the exact same answer you get in the preceding section, but it's calculated in a much more straightforward way.

Stoichiometry and the Gas Laws

The ideal gas equation (and even the combined gas equation) allows chemists to work stoichiometry problems involving gases. (Chapter 10 is your key to the world of stoichiometry.) In this section, you're going to use the ideal gas equation to do such a problem, using a classic chemistry experiment — the decomposition of potassium chlorate to potassium chloride and oxygen by heating:

$$2 \ KClO_3(s) \rightarrow 2 \ KCl(s) + 3 \ O_2(g)$$
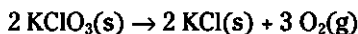
Here's your mission: Figure out the volume of oxygen gas produced at 700 torr and 27 degrees Celsius from the decomposition of 25.0 grams of $KClO_3$.

First, you need to calculate the number of moles of oxygen gas produced:

$$\frac{50.0g \ O_2}{1} \times \frac{50.0g \ O_2}{1} \times \frac{3 \ mol \ O_2}{1} = 0.3059 \ mol \ O_2$$

Next, convert the temperature to Kelvin and the pressure to atm:

$$27 \,^{\circ}C + 273 = 300 \ K$$

$$700 \ torr/760 \ torr/atm = 0.9211 \ atm$$

Now put everything in the ideal gas equation:

$$PV = nRT$$

$$V = nRT/P$$

$$V = [(0.3059 \ mol)(0.0821 \ L \cdot atm/K \cdot mol)(300 \ K)]/0.9211 \ atm = 8.18 \ liters$$

Mission accomplished.

Dalton's and Graham's Laws

This section covers a couple of miscellaneous but fine gas laws you should have a nodding acquaintance with. One relates to partial pressures and the other to gaseous effusion/diffusion. Party on.

Dalton's Law

Dalton's Law of partial pressures says that in a mixture of gases, the total pressure is the sum of the partial pressures of each individual gas.

If you have a mixture of gases — gas A, gas B, gas C, and so on — then the total pressure of the system is simply the sum of the pressures of the individual gases. Mathematically, the relationship can be expressed like this:

$$P_{Total} = P_A + P_B + P_C + \ldots$$

When working stoichiometry problems like the one in the preceding section involving the decomposition of potassium chlorate, the oxygen is normally collected over water by displacement and the volume is then measured. However, in order to get the pressure of just the oxygen, you have to subtract the pressure due to the water vapor. You have to mathematically "dry out" the gas.

Suppose, for example, that a sample of oxygen is collected over water at a total pressure of 755 torr at 20 degrees Celsius. And suppose that your job, you lucky dawg, is to calculate the pressure of the oxygen.

You know that the total pressure is 755 torr. Your first task is to reference a table of vapor pressures of water versus temperature. (You can find such a table in a variety of places, such as the *Chemical Rubber Company (CRC) Handbook.*) After looking at the table, you determine that the partial pressure of water at 20 degrees Celsius is 17.5 torr. Now you're ready to calculate the pressure of the oxygen:

$$P_{Total} = P_{Oxygen} + P_{water\ vapor}$$

$$755\ torr = P_{oxygen} + 17.5\ torr$$

$$P_{oxygen} = 755\ torr - 17.5\ torr = 737.5\ torr$$

Knowing the partial pressure of gases like oxygen is important in deep sea diving and the operation of respirators in hospitals.

Graham's Law

Place a few drops of a strong perfume on a table at one end of a room, and soon people at the other end of the room can smell it. This process is called *gaseous diffusion,* the mixing of gases due to their molecular motion.

Place a few drops of that same perfume inside an ordinary rubber balloon and blow it up. Very soon you'll be able to smell the perfume outside of the

balloon as it makes its way through the microscopic pores of the rubber. This process is called *gaseous effusion*, the movement of a gas through a tiny opening. The same process of effusion is responsible for the helium being quickly lost from rubber balloons.

Thomas Graham determined that the rates of diffusion and effusion of gases are inversely proportional to the square roots of their molecular or atomic weights. This is Graham's Law. In general, it says that the lighter the gas, the faster it will effuse (or diffuse). Mathematically, Graham's Law looks like this:

$$\frac{V_1}{V_2} = \sqrt{\frac{M_2}{M_1}}$$

Suppose that you fill two rubber balloons to the same size, one with hydrogen (H_2) and the other with oxygen (O_2). The hydrogen, being lighter, should effuse through the balloon pores faster. But how much faster? Using Graham's Law, you can determine the answer:

$$\frac{V_{H_2}}{V_{O_2}} = \sqrt{\frac{M_{O_2}}{M_{H_2}}}$$

$$\frac{V_{H_2}}{V_{O_2}} = \sqrt{\frac{32.0\,g/mol}{2.0\,g/mol}}$$

$$\frac{V_{H_2}}{V_{O_2}} = \sqrt{16}$$

$$\frac{V_{H_2}}{V_{O_2}} = 4$$

The hydrogen should effuse out four times as fast as the oxygen.