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Errata for the first printing of Chemistry 3 are listed in this document and will be corrected in subsequent printings of the textbook.

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multiplies the scale of the unit by 1,000—that is, 1,000 times larger than the base unit scale.

Compare this incredibly logical system of units to the chaotic English system. If we want to measure something short, we use the inch unit, which is equal to one-twelfth of a foot. On the other hand, if we want to measure something with small volume, we might use the quart unit, which is equal to one-fourth of a gallon. In the English system, every alternative unit has a different relationship to the base unit, and we must remember all of those crazy numbers. We have to remember that there are 12 inches in a foot, 3 feet in a yard, and 5,280 feet in a mile, while at the same time remembering that for volume there are 8 ounces in a cup, 2 cups in a pint, 2 pints in a quart, and 4 quarts in a gallon. That's a lot to memorize! Thankfully, the majority of science operates around the metric system.

In the metric system, all we have to remember is what the prefix means. Since the *centi-* prefix means one-hundredth, then we know that 1 centimeter is one-hundredth of a meter, 1 centiliter is one-hundredth of a liter, and 1 centigram is one-hundredth a gram. Since the *kilo-* prefix means 1,000, we know that there are 1,000 meters in a kilometer, 1,000 grams in a kilogram, and 1,000 liters in a kiloliter. Doesn't that make a lot more sense?

Another advantage to the metric system is that the prefixes are all based on a factor of 10. Table 1.2 summarizes the most commonly used prefixes and their numerical meanings. The prefixes in boldface type are the ones that we will use over and over again. Memorize those 3 prefixes and their meanings before you take the test for this module. The commonly used abbreviations for these prefixes are listed in parentheses.

TABLE 1.2
Common Prefixes Used in the Metric System


| Prefix | Numerical Meaning |
|------------------|-------------------|
| micro (μ) | 0.00001 |
| milli (m) | 0.001 |
| centi (c) | 0.01 |
| deci (d) | 0.1 |
| deca (D) | 10 |
| hecta (H) | 100 |
| kilo (k) | 1,000 |
| mega (M) | 1,000,000 |

Remember that each of these prefixes, when added to a base unit, makes an alternative scale for measurement. If you wanted to measure the length of something small, you would have all sorts of options for which unit to use. If you wanted to measure the length of someone's foot, you could use the decimeter. Since the decimeter is one-tenth of a meter, it measures things that are only slightly smaller than a meter. On the other hand, if you wanted to measure the length of a sewing needle, you could use the centimeter because a sewing needle is significantly smaller than a meter. Or if you want to measure the thickness of a piece of paper, you might use the millimeter since it is one-thousandth of a meter, which is a *really* small unit.

You can see that the metric system is much more logical and versatile than the English system. That is, in part, why scientists and most countries in the world use it as their main

Summary of Comments on Chemistry 3 Book Extra Errata.pdf

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between centimeters and meters. According to table 1.2, *centi-* means 0.01. So 1 centimeter is the same thing as 0.01 meters. In mathematical form, we would say:

$$1 \text{ centimeter} = 0.01 \text{ meter}$$

Look what happens when we divide both sides by 1 centimeter:

$$\frac{1 \text{ centimeter}}{1 \text{ centimeter}} = \frac{0.01 \text{ meter}}{1 \text{ centimeter}}$$

The fraction on the right equals 1! Do you see how that works? If 1 centimeter equals 0.01 meters, then the numerator and the denominator on the right side are saying the same thing, which is why that fraction equals 1. So if we use this fraction and multiply it by a measurement, then we will not change the value of the measurement. We call this fraction a conversion factor.

Now that we know how centimeters and meters relate to one another, we can convert from one to another. First, we write down the measurement that we know:

$$15.1 \text{ centimeters}$$

We then realize that any number can be expressed as a fraction by putting it over the number 1. So we can rewrite our measurement as:

$$\frac{15.1 \text{ centimeters}}{1}$$

Now we can convert that measurement into meters by multiplying it with the conversion factor we determined above. We have to do it the right way so that the original measurement unit cancels out when multiplied by the conversion factor. Here's how we do it:

$$\frac{15.1 \text{ centimeters}}{1} \times \frac{0.01 \text{ meter}}{1 \text{ centimeter}} = 0.151 \text{ meter}$$

Given Unit Conversion Factor Wanted Unit

So 15.1 centimeters is the same as 0.151 meters. This conversion method, called the **factor-label method**, works for 2 reasons. First, since 0.01 meters is the same as 1 centimeter, multiplying our measurement by 0.01 meters over 1 centimeter is the same as multiplying by 1. Since nothing changes when we multiply by 1, we haven't altered the value of our measurement at all. Second, by putting the 1 centimeter in the denominator of the second fraction, we allow the centimeters unit to cancel (just like the 64 canceled in the previous discussion). Once the centimeters unit has canceled, the only thing left is meters, so we know that our measurement is now in meters.

This is how we will do all of our factor-label setups. We will first write the measurement we are given in fraction form by putting it over 1. We will then put the unit we do not want in the denominator of the conversion factor and put the unit we do want in the numerator. Finally, the numerical meaning of any prefixes needs to go on the opposite side of the conversion factor to get the fraction to equal 1. We will see many examples of this method, so don't be concerned if you are a little confused right now.

ANSWERS TO THE “ON YOUR OWN” QUESTIONS

$$1.1 \quad \frac{9,321 \cancel{\text{g}}}{1} \times \frac{1 \text{ kg}}{1,000 \cancel{\text{g}}} = 9.321 \text{ kg}$$

$$1.2 \quad \frac{0.465 \cancel{\text{L}}}{1} \times \frac{1 \text{ mL}}{1,000 \cancel{\text{L}}} = 465 \text{ mL}$$

$$1.3 \quad \frac{724.0 \cancel{\text{cm}}}{1} \times \frac{0.01 \text{ m}}{1 \cancel{\text{cm}}} = 7.240 \text{ m}$$

$$1.4 \quad \frac{8.465 \cancel{\text{st}}}{1} \times \frac{14.59 \text{ kg}}{1 \cancel{\text{st}}} = 123.5 \text{ kg}$$

$$1.5 \quad \frac{6.1236 \cancel{\text{L}}}{1} \times \frac{1 \text{ gal}}{3.78 \cancel{\text{L}}} = 1.62 \text{ gal}$$

$$1.6 \quad \frac{1,500 \cancel{\text{mL}}}{1} \times \frac{0.001 \cancel{\text{L}}}{1 \cancel{\text{mL}}} \times \frac{1 \text{ kL}}{1,000 \cancel{\text{L}}} = 0.0015 \text{ kL}$$

$$1.7 \quad \frac{2 \cancel{\text{km}}}{1} \times \frac{1,000 \cancel{\text{m}}}{1 \cancel{\text{km}}} \times \frac{1 \text{ cm}}{0.01 \cancel{\text{m}}} = 200,000 \text{ cm}$$

$$1.8 \quad \frac{0.01 \cancel{\text{Mg}}}{1} \times \frac{1,000,000 \cancel{\text{g}}}{1 \cancel{\text{Mg}}} \times \frac{1 \text{ mg}}{0.001 \cancel{\text{g}}} = 10,000,000 \text{ mg}$$


$$1.9 \quad \frac{0.00555 \cancel{\text{hr}}}{1} \times \frac{60 \cancel{\text{min}}}{1 \cancel{\text{hr}}} \times \frac{60 \text{ sec}}{1 \cancel{\text{min}}} = 19.98 \text{ sec}$$


That is not a long time to hold one's breath. So we would not be impressed.

$$1.10 \quad \frac{0.0091 \cancel{\text{kL}}}{1} \times \frac{1,000 \cancel{\text{L}}}{1 \cancel{\text{kL}}} \times \frac{1 \text{ mL}}{0.001 \cancel{\text{L}}} = 9,100 \text{ mL} = 9,100 \text{ cm}^3 \text{ (mL and cm}^3 \text{ equivalent)}$$

1.11 The relationship between m and mm is easy:

$$1 \text{ mm} = 0.001 \text{ m}$$

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So she did, indeed, determine the proper recipe for making table salt: 10.0 grams chlorine plus 6.5 grams sodium makes 16.5 grams table salt. To convince herself that the law of definite proportions really works, she tripled the masses of sodium and chlorine and tried again:



19.5 g sodium and 30.0 g chlorine makes 49.5 g table salt.

Once again, the experiment ended with no leftovers because the elements were added together in the proper proportion.

As a final experiment, the chemist adds 6.5 grams of sodium to 15.0 grams of chlorine. In this case, the law of definite proportions shows that she used too much chlorine because 6.5 grams of sodium needs only 10.0 grams of chlorine. There should, therefore, be 5.0 grams of chlorine left over in the end.



6.5 g sodium and 15.0 g chlorine makes 16.5 g table salt and 9.0 g leftover chlorine.

The law of definite proportions, coupled with the law of mass conservation, can be used to solve problems like example 2.2. After studying the example, answer “On Your Own” questions 2.5–2.6.

EXAMPLE 2.2

A chemist reacts 15.0 grams of calcium (Ca) with 15.0 grams of oxygen (O). This reaction makes 21.0 grams of a compound known as lime. Along with the lime, there is also some leftover oxygen. If the chemist wants to make 55 grams of lime and have no leftover oxygen or calcium, how much of each element should he use?

To solve this problem, we first need to determine the recipe for making lime. To do that, we need to figure out how much oxygen was used in the reaction. Since there was no leftover calcium, we know that all 15.0 grams of calcium were used. How much leftover oxygen was there? We can find out by using the law of mass conservation:

Total mass before reaction = Total mass after reaction

$15.0 \text{ g} + 15.0 \text{ g} = 21.0 \text{ g} + \text{Mass of leftover oxygen}$

$\text{Mass of leftover oxygen} = 30.0 \text{ g} - 21.0 \text{ g} = 9.0 \text{ g}$

Since 9.0 grams of oxygen were left over:

$\text{Mass of oxygen used} = \text{Starting mass} - \text{Mass left over}$

We can keep only 2 significant figures because our frequency had only 2. This might seem like a small amount of energy, but it corresponds to the energy of ultraviolet light, so it is enough to kill living tissue!

A light wave has a wavelength of 1.21×10^{-12} m. What is its energy?

This problem is a little harder because the only equation we can use to calculate the energy of light is equation 3.2, and it uses *frequency*, not *wavelength*. To solve this, then, we must first turn the wavelength we've been given into frequency. We can do this with equation 3.1:

$$f = \frac{c}{\lambda}$$

$$f = \frac{3.0 \times 10^8 \frac{\text{m}}{\text{s}}}{1.21 \times 10^{-12} \text{ m}} = 2.5 \times 10^{20} \frac{1}{\text{s}} = 2.5 \times 10^{20} \text{ Hz}$$

Now that we have the frequency, we can use equation 3.2:

$$E = h \times f$$

$$E = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}} \times 2.5 \times 10^{20} \text{ Hz}$$

$$E = 1.7 \times 10^{-13} \text{ J}$$

The energy is 1.7×10^{-13} J.

ON YOUR OWN

($h = 6.63 \times 10^{-34} \frac{\text{J}}{\text{Hz}}$; $c = 3.0 \times 10^8 \frac{\text{m}}{\text{s}}$;
nano- means 10^{-9})

- 3.6 If a light wave has energy of 3.4×10^{-14} J, what is its frequency?
- 3.7 If a visible light wave has energy of 3.3×10^{-19} J, what color is it?
 (You can use figure 3.10 if you need to.)

Make sure you understand how to use equation 3.2 by answering “On Your Own” questions 3.6–3.7.

HOW THE EYE DETECTS COLOR

Before we go on to discussing the Bohr model of the atom, perform experiment 3.2. Hopefully, it will give you some insight into how humans perceive color.

confused if you read about electron-dot diagrams in other books. They are the same as Lewis structures. Make sure you understand how to determine an atom's Lewis structure by studying example 4.1 and answering "On Your Own" question 4.1.

EXAMPLE 4.1

Draw the Lewis structure for the following atoms:

a. Na b. Ga c. Te

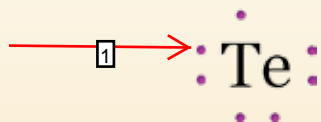
(a) Since Na is in group 1A, it has 1 valence electron. In a Lewis structure, a single electron is written on the right side of the symbol:



(b) Since Ga is in group 3A, it has 3 valence electrons. We start on the right side of the symbol and put a single dot on the right, then the bottom, then the left of the symbol:



(c) Since Te is in group 6A, it has 6 valence electrons. The first 4 are put on the sides, bottom, and top of the symbol by themselves. After that, the last 2 must be paired with others. We pair the electrons in the same order that we put them down singly; so the last 2 dots will pair with the dots on the right and bottom of the symbol:



ON YOUR OWN

4.1 Draw the Lewis structures for the following atoms:

a. Ca b. Si c. At

LEWIS STRUCTURES FOR IONIC COMPOUNDS

Now we can learn how Lewis structures help us in determining molecular structure. We will start with ionic compounds because they are the easiest to understand. Consider table salt, NaCl. Why do the elements Na and Cl come together and make a molecule?

Also, why is there 1 Na atom for every 1 Cl atom, as the formula indicates? Why aren't there 2 Na atoms for every 3 Cl atoms so that the formula is Na₂Cl₃? The answers to these questions are readily available by looking at the Lewis structures of these 2 atoms:



First, look at chlorine's Lewis structure. It is very close to the ideal valence electron configuration. All it needs to do is get 1 more electron, and then it will have 8 valence electrons. Sodium just happens to have 1 valence electron (1 dot in its Lewis structure).

Step 4: Name the compound, cation first. The ion that comes from sodium is named sodium since it is positive. Since the ion that comes from nitrogen is negative, it is called nitride. So the name is **sodium nitride**.

Once again, we did not have to be told that this is an ionic compound. We know that because it contains a metal and a nonmetal and because Na loses an electron and N gains electrons to become ions.

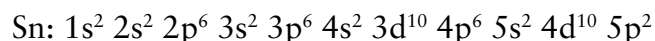
This shortcut allows you to get the chemical formula of an ionic compound very quickly, but don't let that distract you from what is really going on here. Remember that this method is just a shortcut to skip writing out the Lewis structures and moving the dots around. In nature, the chemical formulas of these compounds are determined the way we showed you with Lewis structures. This method is simply a shortcut that saves time. Make sure you understand it by answering "On Your Own" questions 4.2–4.3.

ON YOUR OWN

- 4.2 Give the chemical formulas for the following compounds:
- beryllium chloride
 - potassium oxide
 - aluminum phosphide
 - magnesium nitride
- 4.3 What is the chemical formula for the compound that forms when cesium and nitrogen are reacted together? What is its name?

HANDLING THE EXCEPTIONS IN IONIC COMPOUNDS

You might wonder about all the exceptions on the periodic table. In the list of rules on determining charge, we mentioned that Ge, Sn, and Pb are all exceptions. Why are they exceptions? Well, let's take a look at the Lewis structures, using Sn as an example.





Sn has 4 valence electrons; this is why it is in column 4A of the periodic table. Also, it has 10 electrons in the 4d orbital in addition to the full 4s and 4p orbitals it needs for a full octet. Losing d electrons will allow the **transition** metal to become more than one possible ion.


When we write out the name of the ionic compound involving a **transition metal**, we need to be very clear about which ion is present. For example, tin chloride is an ambiguous name because the tin could have any number of charges. How do we keep the multiple ion possibilities straight?


We are going to put a roman numeral in the name that represents the charge of the **transition** metal. So tin (IV) represents Sn^{+4} and tin (II) represents Sn^{+2} . We will do this for all **transition** metals. Instead of tin chloride, we should write tin (IV) chloride or tin (II) chloride; the roman numeral matches the charge on the **transition** metal ion.


This roman numeral helps us determine the chemical formula of the ionic compound. With tin (II) chloride the II represents Sn^{+2} , and chloride is Cl^{-1} . Combining the ions gives a formula of SnCl_2 . In the same way, tin (IV) chloride represents SnCl_4 because tin (IV) is Sn^{+4} and chloride is Cl^{-1} .

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these metals

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When dealing with ions of atoms which we call exceptions to the rules, you will always be given a roman numeral that signifies the positive charge on the metal. With that information, you can use the same skill you just learned to determine the chemical formulas of those ionic compounds as well.

TABLE 4.2
Common Transition Metals That Form More Than One Positive Ion

| Element | Possible Ions | Name of Ions |
|----------|-------------------------------------|-------------------------------|
| Chromium | Cr^{2+} , Cr^{3+} | Chromium (II), Chromium (III) |
| Iron | Fe^{2+} , Fe^{3+} | Iron (II), Iron (III) |
| Copper | Cu^+ , Cu^{2+} | Copper (I), Copper (II) |
| Gold | Au^+ , Au^{2+} | Gold (I), Gold (II) |
| Tin | Sn^{2+} , Sn^{4+} | Tin (II), Tin (IV) |
| Lead | Pb^{2+} , Pb^{4+} | Lead (II), Lead (IV) |

Knowing that the transition metals may have multiple charges, you can also use the chemical formula to determine the name of the compound. You do not need to memorize table 4.2. Example 4.3 gives you some practice in naming ionic compounds with transition metals.

EXAMPLE 4.3

What is the chemical name for the ionic compound CuCl_2 ?

Step 1: Identify the cation and anion and their charges. To solve this problem, we first need to determine the charges that Cu and Cl develop in an ionic compound. Cl is in group 7A, so it wants to be 1-. Cu is in group 8B, so it can possibly have multiple charges, but since Cl is 1- and there are 2 atoms of Cl for every Cu atom, then Cu must be 2+.

Step 2: Name the transition metal cation by its elemental name followed by the roman numeral. Copper (II) will be the name of the cation.

Step 3: Name the anion by its *-ide* name. Cl^{-1} is chloride.

Therefore, the compound's name is **copper (II) chloride**.


What is the chemical name for the ionic compound PbO_2 ?


Step 1: Identify the cation and anion and their charges. O is in group 6A, so it wants to be 2-. Pb is in group 4A, so it can possibly have multiple charges, but since O is 2- and there are 2 O atoms for every Pb atom, then Pb must have a charge of 4+.


Step 2: Name the transition metal cation by its element name followed by the roman numeral. Pb^{4+} is lead (IV).

Step 3: Name the anion by its *-ide* name: O^{2-} is oxide.

Therefore, the compound's name is **lead (IV) oxide**.

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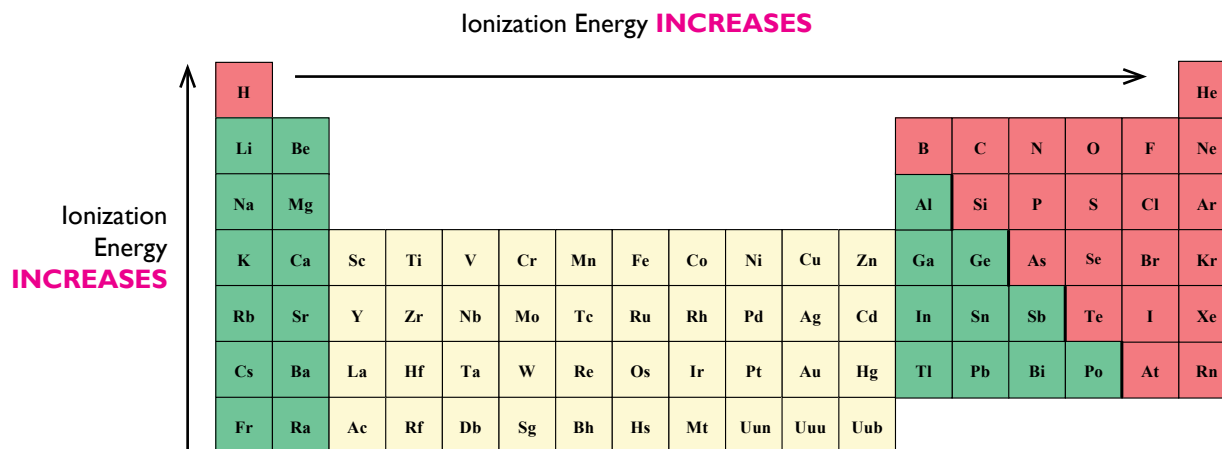
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that have multiple charges

FIGURE 4.1
The Periodic Behavior of Ionization Energy



Make sure you understand how to use this periodic property by studying example 4.4 and then answering "On Your Own" questions 4.6–4.7.

EXAMPLE 4.4

A chemist needs to find an element that readily gives up an electron. Which of the following would suit her needs best: Ca, As, or Br?

Since all of these atoms lie in the same row of the periodic table, they do not lie above or below each other. The fact that ionization energy decreases from the top to the bottom of the table is useless here. However, Ca lies to the left of As, which lies to the left of Br. Since ionization energy increases from left to right on the periodic table, Ca has the lowest ionization energy. **Ca gives up its electrons most readily and is the best choice for the experiment.**

Which of the following holds onto its electrons the tightest: Al, In, or B?

All of these elements lie in the same column, so the fact that ionization energy increases from left to right on the table is of no use. However, B lies above Al, which lies above In. Since ionization energy decreases from the top to the bottom of the table, B has the highest ionization energy. This means that **B holds on to its electrons the tightest.**

Which element has the highest ionization energy?

~~Because ionization energy increases from left to right and decreases from top to bottom, fluorine has the highest ionization energy.~~ **1**

ON YOUR OWN

- 4.6 Order the following atoms in terms of increasing ionization energy: At, Ir, Pb, Cs.
- 4.7 Which atom most readily gives up its electrons: Ca, Be, Ba, Ra?

This is because all noble gases do not want to add or lose any electrons.

ELECTRONEGATIVITY: ANOTHER PERIODIC PROPERTY

The way in which we have defined ionization energy deals only with positive ions. After all, we defined ionization energy as the energy needed to take an electron away from an atom. When electrons are taken away from atoms, the atoms become positive ions. What about negative ions? Is there a way of determining how much an atom desires to gain electrons? Of course there is. The property that determines an atom's desire to gain electrons is called **electronegativity**.

Electronegativity—A measure of how strongly an atom attracts extra electrons to itself.

As the electronegativity of an atom increases, its desire to become a negative ion increases.

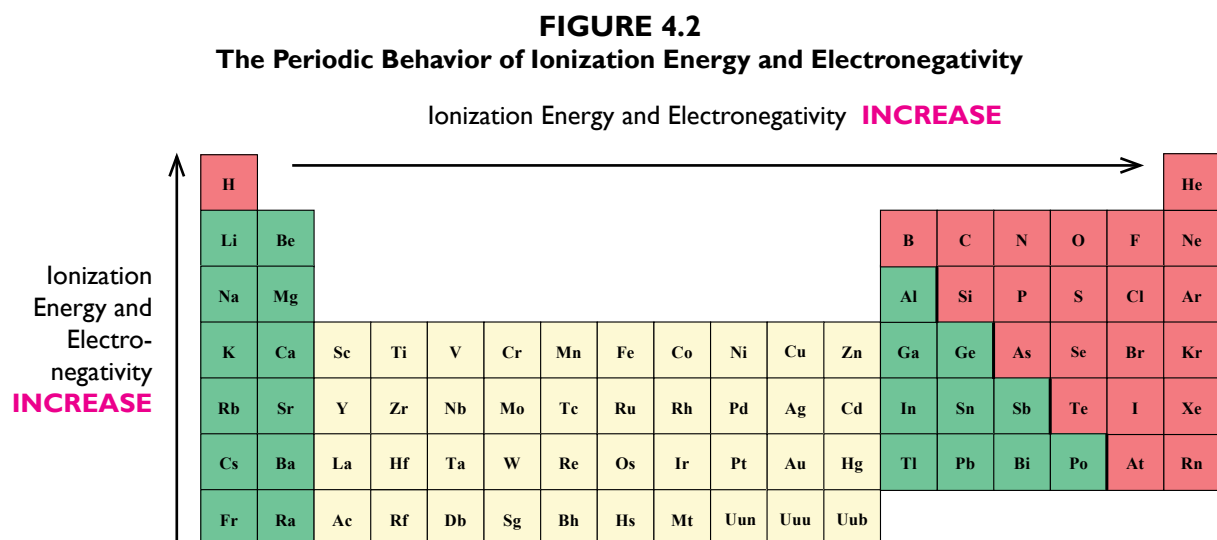
Let's determine how electronegativity varies across the periodic table. First, consider atoms in group 1A. Do they desire extra electrons? No, because they need to lose an electron to attain the ideal electron configuration. On the other hand, atoms in group 7A have a strong desire to gain an extra electron because that will give them the ideal electron configuration. Therefore, we can conclude:

In general, the electronegativity of atoms increases from left to right on the periodic table.

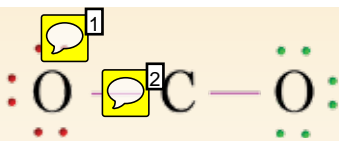
What about an atom's vertical position on the table? How does that affect its electronegativity? As already noted, atoms that are higher on the periodic table tend to have their valence electrons closer to the nucleus as compared to atoms that appear near the bottom of the table. The attraction of an atom for electrons should decrease from top to bottom on the table because the farther the valence electrons are from the nucleus, the less attraction the protons can exert on them. □

In general, the electronegativity of atoms increases from bottom to top on the periodic table.

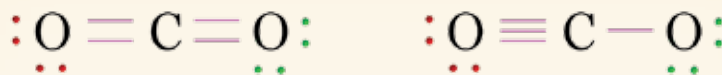
Figure 4.2 summarizes the way in which both ionization energy and electronegativity vary across the periodic table.



Noble gases (group 8A) do not follow these electronegativity periodic properties because they have the ideal electron configuration and do not want to gain electrons.

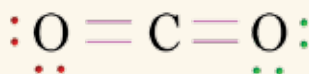


This gives carbon 6 electrons, so we will need 1 more multiple bond to give it 8. Should we double bond with the oxygen atom on the right or triple bond with the oxygen atom on the left?



For this course, choose symmetry whenever possible. In college chemistry you will learn more about how the electronegativities of the outside atoms determine what kinds of bonds are formed. However, there is a rule you should follow with multiple bonds. Group 7A atoms do not double or triple bond due to their high electronegativities.

Now all atoms in the molecule have 8 electrons, so our Lewis structure is complete.



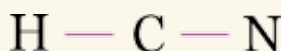
A carbon dioxide molecule, therefore, consists of 2 oxygen atoms, each double-bonded to a carbon atom.

What is the Lewis structure for HCN?

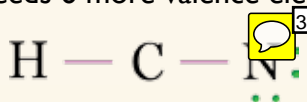
The formula tells us that we have 1 hydrogen atom, 1 carbon atom, and 1 nitrogen atom to build into our Lewis structure.



According to the procedure we are supposed to follow, we must put the carbon in the middle since it has the most unpaired electrons, and attach the others to it:




Fill in octets on outside atoms first. The H atom only wants 2 valence electrons, so once H is bonded, it is done. So the N atom needs 6 more valence electrons around it.





This uses all 10 of our valence electrons. We need to multiple bond with N to get C its 8 valence electrons.



Sharing a double bond between C and N gets C access to only 6 valence electrons. This means we will need to share another bond between C and N, making a triple bond.

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Remove the two red dots on top of this first O

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add a second pink line here to make a double bond

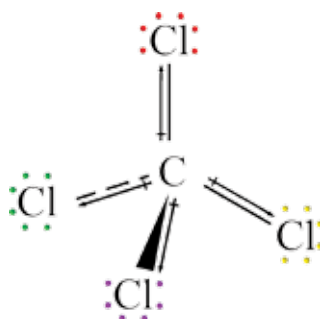
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NONPOLAR COVALENT AND POLAR COVALENT MOLECULES

Remember, the important thing about polar covalent bonds and nonpolar covalent bonds is their effect on the molecule they are in. If a molecule is polar covalent, it contains electrical charges. On the other hand, a covalent molecule can have polar bonds but be nonpolar overall (with no net charge). Since this is a relatively important thing to know about a molecule, it is important to be able to determine whether a given molecule is polar covalent or nonpolar covalent.

So what makes a molecule polar covalent? It stands to reason that a molecule must contain polar bonds to be a polar covalent molecule. After all, if the bonds do not result in fractional charges, then there is no way that a molecule can become polar covalent. It might surprise you that polar bonds are not enough to ensure that a molecule is polar covalent. This is because, under certain circumstances, the polar bonds in a molecule can all work against each other and end up canceling each other out. When this happens, the molecule has an overall net of no charges in it, despite the fact that it has polar bonds, and the molecule itself is nonpolar covalent.

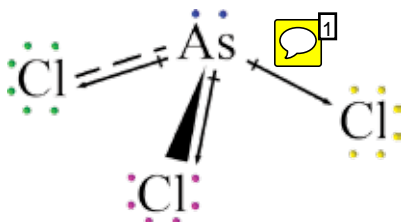
The best way to explain why this is the case is to give you an example of a molecule that has polar bonds but is nonpolar covalent. The perfect example of this is the CCl_4 molecule. The shape of this molecule is as follows:



As shown, the chlorine atoms are more electronegative than the carbon. As a result, electrons are being pulled away from the carbon and toward the chlorine atoms. However, the chlorine atoms are all pulling from exactly opposite directions. As a result, the electrons are being pulled evenly in opposite directions.

What happens when something is pulled evenly in opposite directions? It ends up sitting still, of course. That's what happens to the electrons around the carbon in this molecule. Since they are all being pulled evenly in opposite directions, they end up staying put. Although this molecule has 4 polar covalent bonds, they end up canceling each other out, making the molecule overall nonpolar covalent. In other words, since the electrons around the carbon cannot move, there are no resulting charges in the molecule.

Notice that this is not what happens in the case of AsCl_3 . This molecule has the following shape:



 Number: 1 Author: Rachael Subject: Sticky Note Date: 2/2/15, 3:29:27 PM

There needs to be a straight solid "bond" line between the As and the Cl

mass of carbon dioxide decreases significantly. Finally, carbon dioxide can also exist in the solid phase. We call solid carbon dioxide dry ice. Dry ice is formed when the molecules of carbon dioxide get as close as possible to one another. This further reduces the volume that the same mass of carbon dioxide occupies, making it a solid.

ON YOUR OWN

6.2 Cubes of frozen rubbing alcohol are put into liquid rubbing alcohol. Will the cubes float or sink?

6.3 Natural gas companies store most of their natural gas in liquid form rather than in its gas phase, even though they must deliver it to their customers as a gas. Why do they do this?

The obvious question is this: Why do the molecules in carbon dioxide get closer together or farther apart? If carbon dioxide molecules are attracted to one another, why don't they all just get as close as possible to one another, making carbon dioxide always a solid? To answer this question, we have to discuss the kinetic theory of matter. We will do that in the next section. First, answer "On Your Own" questions 6.2-6.3.

THE KINETIC THEORY OF MATTER

Did you know that right now, as you read this sentence, the molecules that make up the paper in this module and the molecules that make up the ink in these words are all moving very rapidly? You may not believe it, but it's true! The molecules or atoms that make up all of the matter we observe are always in constant motion. That is the basis of the **kinetic theory of matter**.

"Wait a minute," you might protest. "How in the world can the molecules which make up this paper be moving when the paper is sitting still?" That's a reasonable question. Atoms and molecules are so small that they cannot be seen, even with a very powerful microscope. As a result, any matter that we can actually see is made up of billions and billions and billions of atoms or molecules in random motion. For example, in the solid phase molecules (or atoms) vibrate back and forth a few billion times every second. Matter in the solid phase doesn't appear to be moving because for every molecule (or atom) that is vibrating one way, there is another molecule (or atom) vibrating another way. On average, solid matter cannot move anywhere because all its molecules or atoms are moving in different directions.

In the liquid phase, molecules and atoms do not just vibrate; they move around. If you were to look at a glass of water and track a single molecule, you would find that the molecule travels throughout the glass. It may

salt and light

George Stokes (1819–1903) was a British scientist who laid foundations for the field of fluids. He participated in the famous Gifford Lectures which sought to diffuse the knowledge of God. These lectures were given in several Scottish universities, and to be appointed lecturer was considered one of the greatest honors a Scottish scientist could obtain. Stokes said, "The evidence for the resurrection of Jesus Christ is never to be separated from a consideration of the character and teaching and works of Jesus Christ. The head and heart must go together" (Stokes 2010, 80).

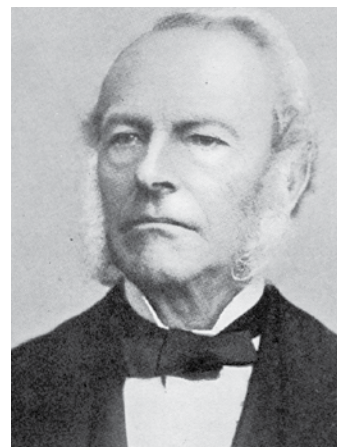



Photo: Public Domain

 Number: 1 Author: Rachael Subject: Highlight Date: 2/2/15, 3:35:24 PM
OYO question 6.2 should be answered after reading through page 220.

The equation for density is:

$$\rho = \frac{m}{V}$$

However, the equation is not much good to us the way it is currently written. Since we would like to calculate the mass of the nugget, we need an equation that starts “m=.” We can use algebra to rearrange the equation so that it reads:

$$m = \rho \times V$$

Now we can simply put in our numbers for volume and density, and we will get the mass, right? There’s only one small problem. If we use the numbers as they are given in the problem, we would not get the correct answer. Why? If we did that, it would look like:

$$m = 19.3 \frac{\text{g}}{\text{mL}} \times 0.125 \text{ L}$$

According to our rules for multiplying numbers with units, we multiply the numbers, and then we do the same with the units. Since we are calculating mass, we know that our answer should end up with the unit of grams. If we multiply these 2 units together, we don’t get grams. To get grams, the mL in the denominator would have to cancel with the L in the volume measurement. That doesn’t work. Instead, we need to change one of the units. It doesn’t matter which one, but let’s change the volume measurement into mL.

$$\frac{0.125 \cancel{\text{L}}}{1} \times \frac{1 \text{ mL}}{.001 \cancel{\text{L}}} = 125 \text{ mL}$$


Now that we have the proper units, we can put everything into the equation:


$$m = 19.3 \frac{\text{g}}{\text{mL}} \times 125 \text{ mL} = 2,412.5 \text{ g}$$

We’re still not quite done, however. According to our rules of significant figures, we can have only 3 significant figures in our final answer because there are 3 significant figures in each of our numbers. After rounding, the answer is **2,410 g**. In scientific notation, the answer is **2.41 × 10³ g**. Either expression is correct, as they both have 3 significant figures.

ON YOUR OWN

- 6.5 The density of silver is 10.5 grams per cm³. A jeweler makes a silver bracelet out of 0.200 kg of silver. What is the bracelet’s volume in mL?
- 6.6 A gold miner tries to sell some gold that he found in a nearby river. The person who is thinking about purchasing the gold measures the mass and volume of 1 nugget. The mass is 1.54 kg, and the volume is 2.080 liters. Is this nugget really made of gold? (Remember that the density of gold is 19.3 grams per mL.)

 Number: 1 Author: Rachael Subject: Highlight Date: 1/20/15, 8:10:49 AM
1.56

 Number: 2 Author: Rachael Subject: Highlight Date: 1/20/15, 8:11:10 AM
0.081

$$\#O: 6 \times 2 + 6 \times 1 = 18$$

$$\#O: 1 \times 6 + 6 \times 2 = 18$$

Remember, since O appears in 2 molecules, we add the number of O atoms that come from each molecule. This makes the equation balanced with respect to O.

$$\#H: 6 \times 2 = 12$$

$$\#H: 1 \times 12 = 12$$

The equation is balanced with respect to H.

Since the equation is balanced with respect to all its atoms, it is a **balanced chemical equation**.

ON YOUR OWN

6.7 Determine whether or not the following equation is balanced:



1.8 Determine whether or not the following equation is balanced:



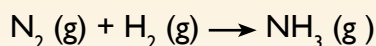
BALANCING CHEMICAL EQUATIONS

Now that you can count the atoms in a chemical equation, you are finally ready to balance chemical equations. When presented with a chemical equation that is not balanced, you must try to balance it by altering the number of molecules involved. You can alter the number of molecules involved by changing the numbers that appear to the left of each molecule. Those numbers can be changed in any way, as long as they are not made negative. *You cannot change any of the numbers that appear in the subscripts.*

If you think about what these numbers mean, this rule should make sense. When balancing chemical equations, all we are trying to do is get the number of molecules involved in the reaction to work out. The *number* of molecules is indicated by the numbers that appear to the left of the molecules. The numbers in the subscripts tell *which* molecules are involved. If you change any of those numbers, you are changing the substances in the chemical reaction, and that is not allowed. So remember, you can alter only the numbers that appear to the left of the molecules, not the subscript numbers. See if you understand these rules by following example 6.3.

EXAMPLE 6.3


Balance the following equation:



To balance this equation, we must get the same number of each type of atom on both sides of the equation. To see where we are, we will first count the number of atoms on each side:

| Reactants Side | Products Side |
|----------------------|----------------------|
| #N: $1 \times 2 = 2$ | #N: $1 \times 1 = 1$ |
| #H: $1 \times 2 = 2$ | #H: $1 \times 3 = 3$ |

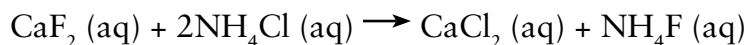
Clearly, this equation is not balanced. How can we balance it? We must change the numbers

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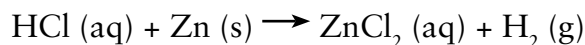
Student is welcome to try OYO 6.8, but parenthesis have not yet been discussed. We recommend you skip this problem.

PRACTICE PROBLEMS

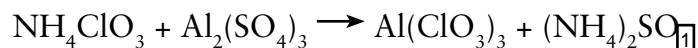
- Classify the following as physical or chemical changes:
 - A precious vase is smashed to bits.
 - Eggs, flour, sugar, milk, and vanilla are baked, making a cake.
 - Dry ice turns into gas.
 - Charcoal is burned in a grill.
 - Kool-Aid is dissolved in water.
- Lead has a density of 11.4 grams per mL. If you make a statue out of 3.45 L of lead, what is the statue's mass?
- Gold has a density of 19.3 grams per cc. If a gold nugget has a mass of 45.6 kg, what is its volume?
- Is the following equation balanced?



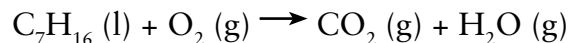
- Balance the following equation:



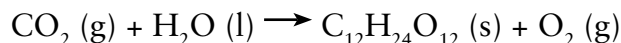
- Balance the following equation:



- In an automobile, nitrogen from the air reacts with oxygen in the engine to make nitrogen monoxide, a toxic pollutant. Write the balanced chemical equation for this reaction.
- Liquid heptene, C_7H_{10} , can react with gaseous hydrogen to make liquid heptane, C_7H_{16} . Write the balanced chemical equation for this reaction.
- Balance the following equation:

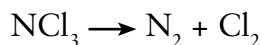


- Balance the following equation:

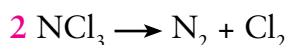


ANSWERS TO THE “ON YOUR OWN” QUESTIONS

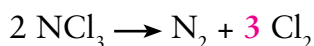
- 7.1 Nitrogen trichloride is NCl_3 , which contains 2 elements: N and Cl. Both nitrogen and chlorine are homonuclear diatomics.



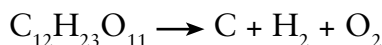
Now we just have to balance the equation. To balance the N, we need a 2 in front of NCl_3 .



To balance the Cl, we need a 3 in front of Cl_2 on the product side:



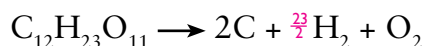
- 7.2 This molecule contains C, H, and O. Hydrogen and oxygen are both homonuclear diatomics, but carbon is not.



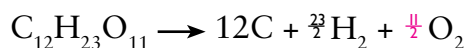
To balance, we start by making 12 C atoms on the products side:



To balance the H without disturbing the C, we need to get 23 H atoms on the products side:



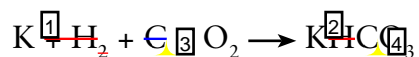
To balance the O without disturbing the C or the H, we need to get 11 O atoms on the products side:



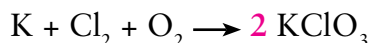
Now we just need to get rid of the fractions:





- 7.3 The compound is composed of K, Cl, and O. Cl and O are homonuclear diatomics, but K is not.





The K atoms are already balanced. To balance the Cl atoms, we need a 2 in front of KClO_3 :



 Number: 1 Author: Rachael Subject: Cross-Out Date: 1/19/15, 1:54:39 PM
there is no H in this equation

 Number: 2 Author: Rachael Subject: Cross-Out Date: 1/21/15, 3:40:30 PM
delete H

 Number: 3 Author: Rachael Subject: Inserted Text Date: 1/19/15, 1:54:23 PM
Cl (subscript 2)

 Number: 4 Author: Rachael Subject: Inserted Text Date: 1/19/15, 1:55:13 PM
l (this is a lower case "L")

7.7 The periodic table tells us that 1 scandium atom has a mass of 45.0 amu. Therefore:

$$45.0 \text{ grams Sc} = 1 \text{ mole Sc}$$

This is the conversion relationship we need for converting grams of scandium into moles of scandium:

$$\frac{50.0 \text{ g Sc}}{1} \times \frac{1 \text{ mole Sc}}{45.0 \text{ g Sc}} = 1.1 \text{ moles Sc}$$

There are **1.11 moles** in the sample.

7.8 To do these conversions, we need to know the mass of the molecule:

$$\text{Mass of } C_6H_{12}O_6 = 6 \times 12.0 \text{ amu} + 12 \times 1.01 \text{ amu} + 6 \times 16.0 \text{ amu} = 180.1 \text{ amu}$$

Remember, we are adding here. When adding (or subtracting), we do not count significant figures. Instead, we report our answer to the same decimal place as the least precise number in the problem. Most of the masses have their last significant figure in the tenths place. The mass of hydrogen has its last significant figure in the hundredths place. The answer must be reported to the tenths place. This means:

$$180.1 \text{ g } C_6H_{12}O_6 = 1 \text{ mole } C_6H_{12}O_6$$

This is the conversion relationship we need for converting grams of glucose into moles of glucose. Before we can use it, though, we need to convert kg of glucose into g of glucose. By now, you should be able to do that on your own. The conversion yields 2.50×10^3 g glucose. Now we can do the grams-to-moles conversion:

$$\frac{2.50 \times 10^3 \text{ g } C_6H_{12}O_6}{1} \times \frac{1 \text{ mole } C_6H_{12}O_6}{180.1 \text{ g } C_6H_{12}O_6} = 13.9 \text{ moles } C_6H_{12}O_6$$

There are **13.9 moles** in the sample.

7.9 This is just another conversion problem, but this time we are converting moles into grams. We still need to determine the formula of nitrogen trioxide, which is NO_3 . Now we need to find the molecular mass of NO_3 :

$$\text{Mass of } NO_3 = 1 \times 14.0 \text{ amu} + 3 \times 16.0 \text{ amu} = 62.0 \text{ amu}$$

This means:

$$62.0 \text{ grams } NO_3 = 1 \text{ mole } NO_3$$

Now we can do our conversion:

$$\frac{0.46 \text{ moles } NO_3}{1} \times \frac{62.0 \text{ g } NO_3}{1 \text{ moles } NO_3} = 29 \text{ g } NO_3$$

The mass is **29 g**.

STUDY GUIDE FOR MODULE 7

REVIEW QUESTIONS

- Can an element undergo a decomposition reaction?
- Classify the following reactions as decomposition, formation, complete combustion, or none of these:
 - $\text{C}_6\text{H}_{12}\text{O}_6 (\text{s}) + 6\text{O}_2 (\text{g}) \rightarrow 6\text{CO}_2 (\text{g}) + 6\text{H}_2\text{O} (\text{g})$
 - $2\text{H}_2 (\text{g}) + \text{O}_2 (\text{g}) \rightarrow 2\text{H}_2\text{O} (\text{l})$
 - $\text{NaCl} (\text{s}) + \text{AgNO}_3 (\text{aq}) \rightarrow \text{AgCl} (\text{s}) + \text{NaNO}_3 (\text{aq})$
 - $2\text{HBr} (\text{l}) + \text{Ca}(\text{OH})_2 (\text{s}) \rightarrow 2\text{H}_2\text{O} (\text{l}) + \text{CaBr}_2 (\text{aq})$
 - $2\text{NaHPO}_4 (\text{s}) \rightarrow 2\text{Na} (\text{s}) + \text{H}_2 (\text{g}) + 2\text{P} (\text{s}) + 4\text{O}_2 (\text{g})$
- Which of the following is *not* a formation reaction?
 - $\text{H}_2\text{CO}_3 + 2\text{KOH} \rightarrow 2\text{H}_2\text{O} + \text{K}_2\text{CO}_3$
 - $\text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{g}) \rightarrow \text{H}_2\text{CO}_3 (\text{l})$
 - $2\text{S} (\text{s}) + \text{Cl}_2 (\text{g}) + 5\text{F}_2 (\text{g}) \rightarrow 2\text{SF}_5\text{Cl}$
- Which of the following is *not* a combustion equation?
 - $2\text{C}_3\text{H}_8\text{O} + 9\text{O}_2 \rightarrow 6\text{CO}_2 + 8\text{H}_2\text{O}$
 - $\text{H}_2\text{CO}_2 (\text{s}) \rightarrow \text{CO}_2 (\text{g}) + \text{H}_2\text{O} (\text{g})$
 - $\text{CH}_2 (\text{g}) + 2\text{O}_2 (\text{g}) \rightarrow \text{CO}_2 (\text{g}) + 2\text{H}_2\text{O} (\text{g})$
 - $2 \text{Ca} (\text{s}) + \text{O}_2 (\text{g}) \rightarrow 2 \text{CaO} (\text{s})$
- What is the difference between complete combustion and incomplete combustion?
- What does a catalytic converter do for an automobile?
- Which has more mass: 100 hydrogen atoms, 4 sulfur atoms, or 1 lanthanum (La) atom?
- What do the following things have in common?
32.1 g of sulfur, 40.1 g of calcium, and 60.1 g of SiO_2
- What is Avogadro's number and what does it represent?
- What are 2 ways the following equation can be interpreted?
 $2 \text{C}_3\text{H}_8\text{O} + 9 \text{O}_2 \rightarrow 6 \text{CO}_2 + 8 \text{H}_2\text{O}$

using the relationships and terms we have used in this module

PRACTICE PROBLEMS

$$1.00 \text{ amu} = 1.66 \times 10^{-24} \text{ g}$$

$$1 \text{ mole} = 6.02 \times 10^{23} \text{ atoms or molecules}$$

1. Write a balanced chemical equation for the decomposition of RbNO_3 .
2. Write a balanced chemical equation for the formation of NaHSO_4 .
3. An automobile uses the energy gained from burning evaporated gasoline to run. Write a balanced chemical equation for the combustion of C_8H_{18} , an important component of gasoline. Include phase symbols in your equation.
4. What is the mass of an argon atom in grams?
5. What is the mass of an aluminum dichromate molecule in kg?
6. How many moles of Au atoms are in 2.56 kg?
7. How many moles of NaHCO_3 are in a 125-gram sample of the compound?
8. What is the mass of a copper (II) chloride sample if it contains 0.344 moles of the compound?
9. How many molecules are in 457 grams of tungsten (W)?
10. In the decomposition of 1.2 moles of dinitrogen pentoxide, how many moles of oxygen would be formed?

can do to one, we can do to the other. We will use this fact to help us take the next step.

In the stoichiometry problems you performed previously, you were always given a decomposition reaction and the amount of reactant, and you were always asked to determine the amount of product that was formed. This may have left you with the impression that stoichiometry must always relate the amount of reactant to the amount of product. This simply isn't true. Like mathematical equations, chemical equations can be used to relate the left side of the equation to the right side of the equation, *and vice versa*.

For example, consider the mathematical relationship below:

$$y + 3 = x$$

If you were told that $y = 4$, then you could determine the value for x , right? You would simply replace y with 4 in the equation, and you would end up determining that $x = 7$. In this case, you related the terms on the left side of the equation to the term on the right side of the equation.

Suppose that you were given the value for x instead of y . Suppose you were told that $x = 7$. Could you solve for y ? Of course you could. You would simply replace x with 7 in the equation, and you would use your algebra skills to determine that $y = 4$. In this case, you related the term on the right side of the equation to those on the left side of the equation.

Chemical equations behave in exactly the same way. We do not always have to give you the amount of reactants and have you calculate the amount of products formed. Instead, we could tell you how much product was formed, and you could use that information to determine the quantities of the reactants involved. See if you can follow this reasoning through example 8.1.

salt and light

Are you beginning to appreciate the study of chemistry? Do you ever wonder where these chemical reactions came from? Well, they came from the research of past chemists. A chemist might spend years running different chemical reactions, or he might find a way to run a known chemical reaction more efficiently. One great American chemist who spent many years in a research laboratory running chemical reactions was **Charles Stine (1882–1954)**, the director of research for DuPont Chemical Company. He developed a chemical reaction that produced a more stable explosive than dynamite. Also, he and Wallace Carothers developed DuPont's laboratory for research that produced Carothers' most famous product, nylon. Stine was a dedicated Christian and wrote his views on the Gospel and Christianity in a book titled *A Chemist and His Bible*. He wrote, "The world about us, far more intricate than any watch, filled with checks and balances of a hundred varieties, marvelous beyond even the imagination of the most skilled scientific investigator; this beautiful and intricate creation, bears the signature of its Creator, graven in its works" (quoted in Morris 1988, 84).

Photo courtesy of DuPont.



EXAMPLE 8.1

A chemist forms ammonium sulfate, $(\text{NH}_4)_2\text{SO}_4$, an excellent fertilizer, by combining ammonia (NH_3) and sulfuric acid (H_2SO_4) . If she ends up making 12.2 moles of ammonium sulfate, how many moles of ammonia were used?

As we learned before, this tells us the minimum quantity of reactants needed to make 100.0 grams of GaCl_3 . We aren't done yet because the problem asked us to calculate the number of grams of each reactant. Since we have the number of moles already, we can use the skills we learned in module 7 to convert this answer to grams.

Let's start with Ga:

$$\text{mass of Ga} = 69.7 \text{ amu}$$

This means:

$$1 \text{ mole Ga} = 69.7 \text{ grams Ga}$$

Therefore:

$$\frac{0.5675 \text{ moles Ga}}{1} \times \frac{69.7 \text{ grams Ga}}{1 \text{ mole Ga}} = 39.6 \text{ grams Ga}$$

Now we can move on to HCl:

$$\text{mass of HCl} = 1.01 \text{ amu} + 35.5 \text{ amu} = 36.5 \text{ amu}$$

Again, this means:

$$1 \text{ mole HCl} = 36.5 \text{ grams HCl}$$

Therefore:

$$\frac{1.703 \text{ moles HCl}}{1} \times \frac{36.5 \text{ grams HCl}}{1 \text{ mole HCl}} = 62.2 \text{ grams HCl}$$

So we learn from this complicated process that to make 100.0 grams of GaCl_3 , **we must start with a minimum of 39.6 grams of Ga and 62.2 grams of HCl.**

The previous example may have seemed a little long and complicated at first, but as you examine the steps we took, it should begin to make sense. After all, we know that stoichiometry allows us to relate the number of moles of one substance in a chemical equation to the number of moles of another substance in the chemical equation. To use stoichiometry, *we must have our quantities in moles*. Therefore, if a problem gives us the amount of a substance in grams, we must first convert to moles before we can use stoichiometry.

Once we have the substance's quantity in moles, we can use the chemical equation and its stoichiometric coefficients to convert the number of moles of that substance into the number of moles of any other substance in the chemical equation. If the problem asks us to put our answer into grams, then we must convert from the moles we just calculated back into grams.

In case you are having a little trouble picturing all of this in your mind, let's illustrate the procedure. Consider figure 8.1.

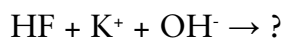
donate more than one (poly) proton (protic).

Chemists often get even more specific in referring to these acids. Phosphoric acid (H_3PO_4) is often called a **triprotic** (try' proh tik) acid because it can donate 3 protons (H^+ ions). On the other hand, sulfuric acid (H_2SO_4) is often referred to as a **diprotic** (dye' proh tik) acid because it can donate 2 protons (H^+ ions). For polyprotic acids, the number of hydrogen atoms that begin the chemical formula tells how many H^+ ions the acid will donate. H_2CO_3 (carbonic acid) is a diprotic acid.

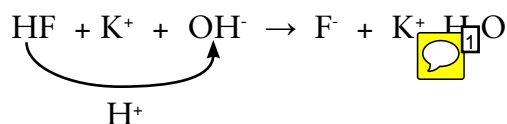
From table 9.1, acetic acid, $\text{C}_2\text{H}_4\text{O}_2$, is the only acid where the number of hydrogen atoms in the formula does not match how many H^+ ions are donated. The reason for this is that once $\text{C}_2\text{H}_4\text{O}_2$ loses 1 H^+ ion, $\text{C}_2\text{H}_3\text{O}_2^{1-}$ will be left over. This is the acetate ion, and it is very stable dissolved in water and therefore does not need or want to lose any more H^+ ions. The rest of the acids from table 9.1 will do the same thing; they will lose enough H^+ ions to get to an ion that will be stable when dissolved in water.

Now that we have this new information under our belts, let's see how to predict the products of chemical reactions involving acids and bases. For example, suppose we had hydrofluoric acid (HF) and potassium hydroxide (KOH). Clearly, HF is an acid since it has a hydrogen atom as the first atom in its chemical formula, and potassium hydroxide is a base because it contains a hydroxide ion. We know this is an acid–base reaction.

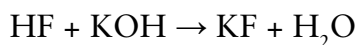
The first thing we do, then, is split up the ionic compound. This will allow us to see exactly what we're dealing with:



The acid (HF) is supposed to donate an H^+ ion. Where do you suppose that ion would like to go? Since it is positive, it will be attracted to the negatively charged hydroxide ion. The H^+ ion will be donated to the OH^- ion. When that happens, the 1+ charge on the hydrogen ion will cancel the 1- charge on the hydroxide ion, and the result will be a neutral molecule with 2 hydrogen atoms and 1 oxygen atom: H_2O . The H^+ ion came from a neutral HF molecule, however. If a positive charge is taken from a neutral substance, the substance becomes negatively charged. When the HF molecule donates its H^+ ion, it becomes an F^- ion. In the end, then, something like this happened:




That's the actual reaction. However, chemists often don't care to write out all of those ions, even though they exist in solution. Instead, we often regroup the ions into the ionic compounds that formed them. The more commonly used version of this equation would be:



This is how we can predict the reactions that occur between acids and bases.

This method for determining the reactions that occur between acids and bases can become very cumbersome when we are dealing with polyprotic acids or bases that contain more than one hydroxide ion [$\text{Mg}(\text{OH})_2$, for example]. We will use the results of the above discussion to develop a shortcut approach to determining the products of an acid–base reaction.

Notice the 2 products formed in the above discussion. The acid and base reacted to

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We are missing an addition "+" symbol between the K positive and the H2O

$$1 \text{ mole H}_2\text{SO}_4 = 98.1 \text{ g H}_2\text{SO}_4$$

$$\frac{40.2 \text{ g H}_2\text{SO}_4}{1} \times \frac{1 \text{ mole H}_2\text{SO}_4}{98.1 \text{ g H}_2\text{SO}_4} = 0.410 \text{ moles H}_2\text{SO}_4$$

Now we can calculate concentration:

$$M = \frac{\# \text{ moles}}{\# \text{ liters}} = \frac{0.410 \text{ moles H}_2\text{SO}_4}{6.8 \text{ L}} = 0.060 \frac{\text{moles H}_2\text{SO}_4}{\text{L}} = 0.060 \text{ M}$$

The concentration is **0.060 M**.

(c) To get concentration, we must have moles and liters. The problem gives us grams and mL, so we must make 2 conversions:

Molecular mass of KOH = (1 x 39.1 amu) + (1 x 16.0 amu) + (1 x 1.01 amu) = 56.1 amu

$$1 \text{ mole KOH} \boxed{=} 56.1 \text{ grams KOH}$$

$$\frac{5.55 \text{ g KOH}}{1} \times \frac{1 \text{ mole KOH}}{56.1 \text{ g KOH}} = 0.0989 \text{ moles KOH}$$

$$\frac{125.0 \text{ mL}}{1} \times \frac{0.001 \text{ L}}{1 \text{ mL}} = 0.1250 \text{ L}$$

Now we can calculate molarity:

$$M = \frac{\# \text{ moles}}{\# \text{ liters}} = \frac{0.0989 \text{ moles KOH}}{0.1250 \text{ L}} = 0.791 \frac{\text{moles KOH}}{\text{L}} = 0.791 \text{ M}$$

The concentration is **0.791 M**.


9.8 We can tell that this is a dilution problem because the chemist has a solution of high concentration that she wants to make into a solution of low concentration. That's what dilution does. The original concentration of the sulfuric acid is 10.0 M, so that is M_1 . The chemist wants the diluted solution to have a concentration of 2.5 M, so that is M_2 . She also knows that she needs to have 500.0 mL of this 2.5 M solution. Since the 500.0 mL is the desired volume of the final solution, that is V_2 . The only thing we don't know in the dilution equation is V_1 , so we can solve for it:

$$M_1 V_1 = M_2 V_2$$

$$10.0 \text{ M} \times V_1 = 2.5 \text{ M} \times 500.0 \text{ mL}$$

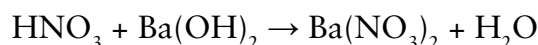
$$V_1 = \frac{2.5 \text{ M} \times 500.0 \text{ mL}}{10.0 \text{ M}} = 130 \text{ mL}$$

The chemist needs to take **130 mL of the original solution and dilute it with water until it reaches 500.0 mL of volume**. This will produce 500.0 mL of a 2.5 M solution.

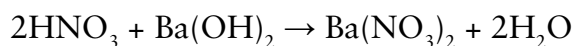
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This should be an = symbol not a subtraction "-" symbol

- 9.9 To solve any stoichiometry problem, we must first figure out the balanced chemical equation. According to the problem, our reactants are HNO_3 and barium hydroxide. What is the chemical formula of barium hydroxide? Remember from module 5 that the polyatomic hydroxide ion is OH^- . Barium is in group 2A of the periodic chart, so it develops a 2^+ charge in ionic compounds. To determine the chemical formula of barium hydroxide, then, we must switch the charges and drop the signs. Thus, the chemical formula is $\text{Ba}(\text{OH})_2$.

When $\text{Ba}(\text{OH})_2$ and HNO_3 react, a salt and water will be produced. The salt will be composed of the positive ion from the base (Ba^{2+}) and the negative ion left over when the acid gets rid of all of its H^+ ions. In this case, that will be the NO_3^- ion. Once again, to get the chemical formula of the compound that forms between these 2 ions, we switch the charges and ignore the signs to get $\text{Ba}(\text{NO}_3)_2$. Thus, the reaction is:



Of course, this equation is not balanced, so that is the first thing we must do:



Now that we have a balanced equation, we can start the stoichiometry. We need to clean up the nitric acid, so we know that it is the limiting reactant. Therefore, we need to know the number of moles of nitric acid to be able to predict how much barium hydroxide is needed. To do this, we will multiply concentration by volume:

$$\frac{12.0 \text{ moles HNO}_3}{1 \text{ L HNO}_3} \times \frac{1 \text{ L HNO}_3}{1} = 12.0 \times 10^2 \text{ moles HNO}_3$$

Notice that we must use scientific notation when reporting this answer because we must report it to 3 significant figures. Scientific notation is the only way we can do that. We can now use this information to calculate the number of moles of barium hydroxide needed:


$$\frac{12.0 \times 10^2 \text{ moles HNO}_3}{1} \times \frac{1 \text{ mole Ba}(\text{OH})_2}{2 \text{ moles HNO}_3} = 60.0 \text{ moles Ba}(\text{OH})_2$$


Now that we have the number of moles of barium hydroxide needed, we can convert back to grams:

$$\frac{60.0 \text{ moles Ba}(\text{OH})_2}{1} \times \frac{171.3 \text{ g Ba}(\text{OH})_2}{1 \text{ mole Ba}(\text{OH})_2} = 1.03 \times 10^4 \text{ g Ba}(\text{OH})_2$$

Thus, **$1.03 \times 10^4 \text{ g}$ of barium hydroxide are needed.** This answer could be reported in decimal notation as well because 10,300 g also has 3 significant figures.

- 9.10 Remember, titrations are just stoichiometry problems, so first we have to come up with a balanced chemical equation. To do that, we must first figure out the chemical formula of the magnesium hydroxide. Magnesium is in group 2A of the periodic chart, so it takes a charge of 2^+ in ionic compounds. You should have memorized

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10.0

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10.0

You should recognize the product of this reaction as carbonic acid. This is what turned the blue litmus paper red in experiment 9.1.

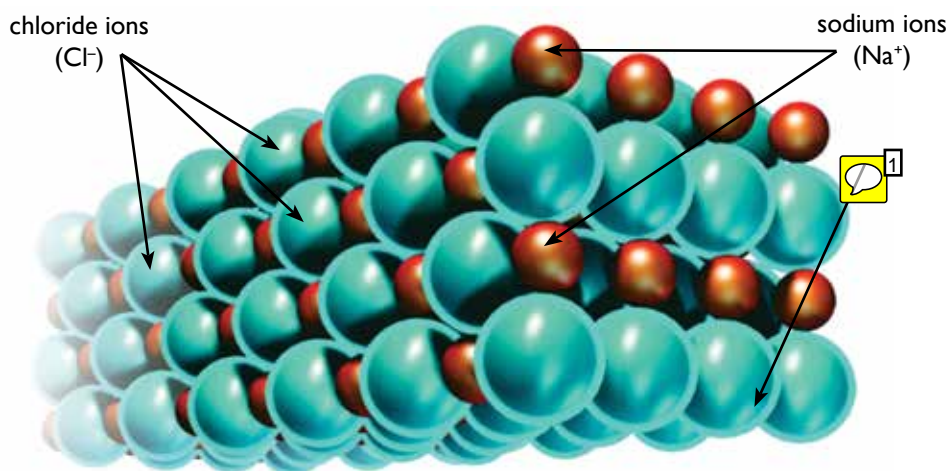
Did you know that solvents do not necessarily have to be liquids? We are very familiar with water dissolving substances like sugar and table salt. Fingernail polish remover is a liquid solvent that dissolves fingernail polish so it can be removed quickly and easily. However, can other phases be solvents? Yes, gases can be solvents! For example, methane gas is transported through pipes to homes to be burned as a fuel. However, methane is colorless and odorless, so it would be very hard to detect gas leaks that can make a very dangerous and explosive situation. So natural gas companies add a small amount of a chemical with a very strong odor so we can detect leaks by smell. In this mixture, methane is the solvent and the odorant is the solute.

HOW SOLUTES DISSOLVE IN SOLVENTS


Now that you know the terminology, you can begin to learn the science. The first thing you might ask yourself is why solutes dissolve in solvents. After all, when you buy table salt (NaCl) at the grocery store, it is a solid. When you mix it with water, why does it go into the aqueous state? Why doesn't it just stay solid? The answers to these important questions involve the polar covalent nature of the water and the ionic nature of the NaCl.

First, remember that NaCl is an ionic compound composed of Na^+ ions and Cl^- ions. Unlike covalent molecules, the ionic NaCl molecule has no chemical bond because the atoms that make it up do not share electrons. Instead, the Na atom has donated its electrons to the Cl atom. As a result, the Na ion is positive and the Cl ion is negative. *The electronic attraction between these opposite charges is the only thing that holds the Na^+ and Cl^- ions together.* Therefore, a sample of solid ionic NaCl would look something like figure 10.1.

FIGURE 10.1
A Solid Sample of NaCl
3D Illustration by Michael Dailey



Sodium chloride (NaCl) is a solid because the Na^+ and Cl^- ions are very close to one another, due to their mutual electronic attraction. As we learned in module 6, the only way to get NaCl out of its solid phase is to pull these ions away from each other. This is where the water comes in.

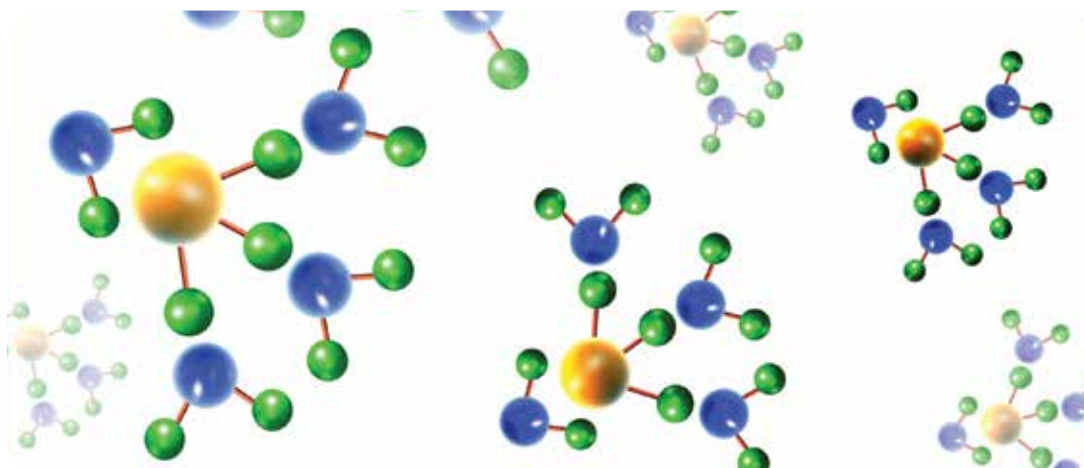
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These fractional charges attract the opposite fractional charges in the water molecules. In the same way as before, the water molecules muscle their way in between the PH_3 molecules. They surround the individual PH_3 molecules, orienting their negatively charged oxygen atoms toward the PH_3 's positively charged hydrogen atoms. In addition, the water's positively charged hydrogen atoms are oriented toward the PH_3 's negatively charged phosphorous atom. Once they are surrounded by the water molecules, the PH_3 molecules are pulled away from each other. When this happens, the compound changes from the solid phase into the aqueous phase, as shown in figure 10.5.

FIGURE 10.5
 PH_3 after It Has Been Dissolved in Water


3D Illustration by Michael Dailey




Notice *how* the water molecules surround the phosphorus trihydride molecules in figure 10.5. The water molecules orient themselves so that their oxygen atoms point toward the hydrogen atoms in the PH_3 molecules because the partial negative charge on the oxygen atom is attracted to the partial positive charges on PH_3 's hydrogen atoms. In the same way, the hydrogen atoms on the water molecules are pointed toward the phosphorus atoms because the partial positive charges on the hydrogen atoms are attracted to the partial negative charge on the phosphorus atom.

There are similarities and differences between the way in which ionic solids dissolve in water and the way in which polar covalent solids dissolve in water. In both cases, water molecules are attracted to the charges in the solute. Also, both types of solute change from the solid phase to the aqueous phase because the water molecules come between the solute molecules, surround them, and then pull them apart. The main difference between these 2 types of compounds is that while ionic compounds break apart into smaller pieces (their individual ions) when dissolved, polar covalent compounds do not. Polar covalent compounds dissolve by separating each individual molecule from its neighbors.

This same reasoning can be used to explain how any ionic or polar covalent solid dissolves in any polar solvent. What about **purely** \square valent compounds, however? How do they dissolve in other **purely** \square valent compounds? If the solvent molecules are attracted to the solute molecules, then the same thing that we have just discussed can also

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non-polar

Please note that purely covalent and non-polar covalent are the same thing.

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non-polar

happen. The solvent molecules will muscle in between the solute molecules, pulling them far apart from each other. However, the precise mechanisms for how this happens are a bit too advanced for this course.

Of course, everything we have discussed so far has been focused on solid solutes dissolving into liquid solvents. As mentioned already, solutes can also be gases or liquids. For example, when dishwashing liquid is mixed with water, the dishwashing liquid is a solute that is dissolved in water (the solvent). The way in which liquid solutes dissolve in liquid solvents is similar to the way in which solid solutes dissolve in liquid solvents.


As is the case with solid solutes, the solvent molecules must be attracted to liquid solute molecules for the solute to dissolve. This allows the solvent molecules to muscle in between the solute molecules. However, in liquid solutes, the solute molecules are not very close to each other to begin with. It isn't very hard for the solvent molecules to get between them. Remember, in solids, the molecules (or ions) are close together. As a result, solvent molecules must work harder to get between the solid solute molecules (or ions). Also, unlike what happens when solids dissolve, the molecules of a liquid solute are not drawn very far away from each other when the solvent gets between them. After all, the solute is a liquid already, so the molecules are already pretty far apart from each other and do not have to be separated very much. They have to be drawn apart only far enough to allow the solvent molecules to get between them.


Based on this information, you would expect that liquid solutes dissolve in liquid solvents more readily than solid solutes do. You probably knew this fact already, but at least now you know why this is the case. Remember that not all liquids can dissolve in other liquids. As mentioned before, the solvent molecules must be attracted to the solute molecules before the solute can be dissolved. As a result, polar covalent liquid solutes will dissolve only in polar covalent liquid solvents, and nonpolar covalent liquid solutes will dissolve only in nonpolar covalent liquid solvents. This is why oil and water don't mix. Oil is nonpolar covalent, while water is polar covalent.


What about gaseous solutes? How do they dissolve in liquid solvents? If you think about it, there is a fundamental difference between solid or liquid solutes and gaseous solutes. In the case of solid solutes, the solute molecules (or ions) must be pulled away from each other to convert them to the liquid state. In the case of liquid solutes, the molecules must also be pulled away from each other, but not by very much. In the case of gaseous solutes, the solute molecules are not pulled apart. In fact, to dissolve into a liquid solvent, the gaseous solute molecules must *be brought closer together*.

Once again, this fact should make sense if you think about it. The main difference between the phases of matter is the distance between the molecules. Molecules in the solid phase tend to be closer together; molecules in the liquid phase tend to be farther apart; and molecules in the gaseous phase tend to be very far apart. To dissolve a gas into a liquid, the gaseous molecules must be brought closer together. You can think of this as precisely the opposite of dissolving a solid solute. When dissolving a gaseous solute, the solvent molecules do not muscle in between the solute molecules. Instead, they attract the solute molecules to them, like a magnet attracts small bits of iron. As a result, the gaseous molecules come closer together.

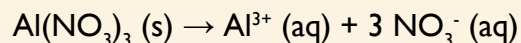
Table 10.1 summarizes the similarities and differences in the processes that dissolve solid, liquid, and gaseous solutes. It is important to understand and remember these facts.

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non-polar

 Number: 2 Author: Rachael Subject: Inserted Text Date: 2/5/15, 6:52:56 PM
non-polar

 Number: 3 Author: Rachael Subject: Inserted Text Date: 1/19/15, 2:07:18 PM
non-polar

compound, $\text{Al}(\text{NO}_3)_3$ splits up into its ions when dissolved. According to its chemical formula, each molecule of aluminum nitrate has 1 aluminum ion and 3 nitrate ions.



The total number of ions is 4, so $i = 4$.

Now that we have all of the components of equation 10.2, we can use it:

$$\Delta T = -i \times K_f \times m = -4 \times 1.86 \frac{^\circ\text{C}}{\text{m}} \times 0.080 \text{ m} = -0.60^\circ\text{C}$$

Realize that this is not the answer to the problem. The problem asks what the freezing point of the solution is. What we found was the change in the freezing point. You are supposed to have memorized that water freezes at 0.00°C . Our answer indicates that the freezing point of this solution is 0.60°C lower than that. So the answer to our problem is that the freezing point of the solution is **-0.60°C** .

What is the freezing point of a solution that is made by dissolving 100.0 grams of table sugar ($\text{C}_{12}\text{H}_{22}\text{O}_{11}$) into 950 grams of water? (K_f of water = $1.86 \frac{^\circ\text{C}}{\text{m}}$)

We start solving this problem by calculating m :

$$\frac{100.0 \text{ g } \text{C}_{12}\text{H}_{22}\text{O}_{11}}{1} \times \frac{1 \text{ mole } \text{C}_{12}\text{H}_{22}\text{O}_{11}}{342.2 \text{ g } \text{C}_{12}\text{H}_{22}\text{O}_{11}} = 0.2922 \text{ moles } \text{C}_{12}\text{H}_{22}\text{O}_{11}$$

$$\frac{950 \text{ g}}{1} \times \frac{1 \text{ kg d}}{1,000 \text{ g}} = 0.95 \text{ kg}$$

$$m = \frac{0.2922 \text{ moles } \text{C}_{12}\text{H}_{22}\text{O}_{11}}{0.95 \text{ kg water}} = 0.31 \text{ m}$$

To figure out i , we just have to think about how table sugar dissolves. According to its formula, there are no metals present, so table sugar is not ionic. It does not split up into ions. It dissolves 1 molecule at a time; therefore, $i = 1$.

Now that we have all of the components of equation 10.2, we can use it:

$$\Delta T = -i \times K_f \times m = -1 \times 1.86 \frac{^\circ\text{C}}{\text{m}} \times 0.31 \text{ m} = -0.58^\circ\text{C}$$

The solution has a freezing point that is 0.58°C lower than the freezing point of the solvent. Since the solvent is water, which normally has a freezing point of 0.00°C , this solution has a freezing point of **-0.58°C** .

This is not quite the answer we need. We were asked to figure out how many grams of oxygen were produced, so we have to convert from moles back to grams:

$$\frac{0.150 \text{ moles } O_2}{1} \times \frac{32.0 \text{ g } O_2}{1 \text{ mole } O_2} = 4.80 \text{ grams } O_2$$

Therefore, **4.80 g of oxygen** will be produced.

10.5 Once again, we can tell that this is a stoichiometry problem because we are asked to convert from the amount of one substance to the amount of another. To do stoichiometry, we must convert to moles. Since we have the number of kg of Al, we will use that as our starting point. First we must convert kg to g:

$$\frac{1.50 \text{ kg}}{1} \times \frac{1000 \text{ g}}{1 \text{ kg}} = 1.50 \times 10^3 \text{ g Al}$$

$$\frac{1.50 \times 10^3 \text{ g Al}}{1} \times \frac{1 \text{ mole Al}}{27.0 \text{ g Al}} = 55.5 \text{ moles Al}$$

Now that we have moles, we can use stoichiometry to determine how much NaOH is needed:

$$\frac{44 \text{ moles Al}}{1} \times \frac{2 \text{ moles NaOH}}{2 \text{ moles Al}} = 44 \text{ moles NaOH}$$

This tells us how much NaOH we need, but it doesn't answer the question. The question asks how many liters of a 8.10 M solution are needed. To determine this, we must remember that M means moles per liter. The molarity of a solution is a conversion relationship that allows us to relate the number of moles to the number of liters. We can therefore do the following conversion:

$$\frac{55.5 \text{ moles NaOH}}{1} \times \frac{1 \text{ L NaOH}}{8.10 \text{ moles NaOH}} = 6.85 \text{ L of solution}$$

This means that **6.85 liters of the NaOH solution** would be needed.


10.6 (a) Since the problem already gives us the number of moles of solute and the number of kg of solvent, this problem is just a straight application of equation 10.1:


$$\text{molality} = \frac{\# \text{ moles solute}}{\# \text{ kg solvent}} = \frac{4.6 \text{ moles NaOH}}{2.4 \text{ kg water}} = 1.9 \frac{\text{moles NaOH}}{\text{kg water}} = 1.9 \text{ m}$$


(b) This problem is a bit more difficult because, although we are given the number of kg of solvent, we are given the number of grams of solute. So we first must convert grams to moles:

$$\frac{125.0 \text{ g } KNO_3}{1} \times \frac{1 \text{ mole } KNO_3}{101.1 \text{ g } KNO_3} = 1.236 \text{ moles } KNO_3$$

Now that we have moles of solute and kg of solvent, we can use equation 10.1:

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55.5

 Number: 2 Author: Rachael Subject: Highlight Date: 1/19/15, 3:36:13 PM
55.5

 Number: 3 Author: Rachael Subject: Highlight Date: 1/19/15, 3:36:20 PM
8.10

$$\frac{3.00 \text{ g Mg}(\text{ClO}_3)_2}{1} \times \frac{1 \text{ mole Mg}(\text{ClO}_3)_2}{159.3 \text{ g Mg}(\text{ClO}_3)_2} = 0.0188 \text{ moles Mg}(\text{ClO}_3)_2$$

$$\frac{30 \text{ g acetic acid}}{1} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.0300 \text{ kg acetic acid}$$

$$m = \frac{0.0188 \text{ moles Mg}(\text{ClO}_3)_2}{0.0300 \text{ kg acetic acid}} = 0.626 \text{ m}$$

To figure out i , we just have to think about how $\text{Mg}(\text{ClO}_3)_2$ dissolves. Being an ionic compound, it splits up into its ions when dissolved. Each molecule of magnesium chlorite has 1 magnesium ion and 2 chlorite ions, so $i = 3$.

Now that we have all of the components of equation 10.2, we can use it:

$$\Delta T = -i \times K_f \times m = -3 \times 3.90 \frac{^\circ\text{C}}{\text{m}} \times 0.626 \text{ m} = -7.32^\circ\text{C}$$

Realize that *this is not the answer to the problem*. The problem asks what the freezing point of the solution is. What we found was the *change in the freezing point*. We are told that pure acetic acid has a freezing point of 16.6°C . Our answer indicates that the freezing point of this solution is 7.32°C lower than that. The answer to our problem is that the freezing point of the solution is $16.6 - 7.32 = 9.28^\circ\text{C}$.

10.9 To calculate boiling points, we must use equation 10.3. To do that, we must know i and m . To calculate m :

$$\frac{100.0 \text{ g CaCO}_3}{1} \times \frac{0.999 \text{ moles CaCO}_3}{100.1 \text{ g CaCO}_3} = 0.999 \text{ moles CaCO}_3$$

$$\frac{400.0 \text{ g acetic acid}}{1} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 0.4000 \text{ kg acetic acid}$$

$$m = \frac{0.999 \text{ moles CaCO}_3}{0.4000 \text{ kg acetic acid}} = 2.50 \text{ m}$$

Since calcium carbonate is an ionic compound, it dissolves by splitting up into its calcium ion and its carbonate ion. So $i = 2$.

$$\Delta T = -i \times K_b \times m = 2 \times 2.93 \frac{^\circ\text{C}}{\text{m}} \times 2.50 \text{ m} = 14.7^\circ\text{C}$$

This means that the boiling point of the solution is 14.7°C *higher* than that of pure acetic acid. The boiling point of pure acetic acid was given as 118.1°C , so the boiling point of this solution is $118.1 + 14.7 = 132.8^\circ\text{C}$.

10.10 The addition of a solute (phosphorus trihydride) to the water will elevate the boiling point. To see how the boiling point is elevated, we must use equation 10.3. To do this, we must calculate i , m , and a formula for phosphorus trihydride. The formula is PH_3 .

$$\frac{265.0 \text{ g } \text{PH}_3}{1} \times \frac{1 \text{ mole } \text{PH}_3}{34.0 \text{ g } \text{PH}_3} = 7.79 \text{ moles } \text{PH}_3$$

$$\frac{7.79 \text{ moles } \text{PH}_3}{2.00 \text{ kg water}} = 3.90 \text{ m}$$

Once again, i just refers to how the substance splits up when it dissolves. Since PH_3 is covalent, it splits up into 1 molecule at a time when it dissolves. So $i = 1$.

$$\Delta T = -i \times K_b \times m = 1 \times 0.512 \frac{^\circ\text{C}}{\text{m}} \times 3.90 \text{ m} = 2.00^\circ\text{C}$$

We must think about what this tells us. The ΔT that we just solved for tells us how many $^\circ\text{C}$ *above the normal boiling point* that the solution boils. You are supposed to have memorized that water boils at 100.0°C , so the boiling point of this solution is **102.0°C** .

Science and Creation Continued

for survival. This means the finch beaks began to decrease in size again. These 2 instances really showed that Darwin's idea was right. The population of finches could indeed adapt to changes in their surroundings from generation to generation.

Even though Darwin was right when he said that species have the ability to change and adapt to their environment, he was dead wrong when he tried to extrapolate from his data. He said that since species have the ability to change, they should be able to change into a different species. In other words, if a population of finches can, through several generations, slowly develop different beaks, why can't they also develop different wings, heads, bodies, and feet so that they change into eagles? Why can't finches, after many years of such change, develop into a completely new species? Darwin thought that this could really happen. This idea became the foundation for the theory of evolution; and it is, unfortunately, still believed by many scientists today. The problem is that Darwin made a big mistake in extrapolation. He took small changes that he observed in animals and extrapolated them into huge changes. Darwin noticed that the finches on the Galapagos Islands had changed their beaks, their feather color, and (to some extent) their body sizes to adapt to the environment of each different island. Those kinds of changes are very small compared to the kinds of changes necessary to turn a finch into a completely different species of bird! So Darwin took a small amount of data and tried to make a huge extrapolation with it. This is precisely the opposite of what a careful scientist does. A careful scientist will make only small extrapolations based on large sets of data. Darwin was not careful and, as a result, has caused all sorts of problems for both the scientific community and the population as a whole.

temperature, remains the same despite the conditions under which the gas is stored. Mathematically, we write:

$$\frac{PV}{T} = \text{constant} \quad \text{Equation 11.7}$$

As we did with both Boyle's law and Charles's law, we can restate this mathematical equation as follows:


$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} \quad \text{Equation 11.8}$$


Extrapolation is a good and necessary part of science, but you must be careful when you do it. In general, an extrapolation is good as long as the amount of extrapolation is small compared to the available data. For example, in figure 11.5, we have data that spans a temperature range of 0°C to -250°C. That's a range of 250 degrees. In the end, we had to extrapolate an additional 23.15 degrees to get all of the lines to converge. Compared to 250 degrees, an additional 23.15 degrees is pretty small, so we can say that our extrapolation is probably okay.

THE COMBINED GAS LAW

Now that you have seen both Boyle's law and Charles's law, it is time to combine them. Remember, the experiments that Boyle performed involved changing the volume of a gas and measuring its pressure or vice versa. During all of those experiments, the temperature was held constant. Likewise, when Charles performed his experiments, he changed the temperature of a gas and measured its volume or vice versa. During those experiments, the pressure exerted by the gas was held constant.

Even though the experiments performed by Boyle and Charles were quite different, we can combine their results. The product of a gas's pressure and volume, divided by the gas's

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Since we were given the partial pressure of each component, we can use equation 11.9 to calculate the total pressure:

$$P_T = P_{\text{oxygen}} + P_{\text{carbon dioxide}} + P_{\text{nitrogen}}$$

$$P_T = 456 \text{ torr} + 112 \text{ torr} + 501 \text{ torr} = 1,069 \text{ torr}$$

Now that we have both the total pressure of the mixture and the partial pressure of each component, we can use equation 11.11 to calculate the mole fraction of each component:

$$P_i = X_i \cdot P_T$$

$$X_i = \frac{P_i}{P_T}$$

Plugging in all 3 sets of numbers:

$$X_{\text{oxygen}} = \frac{456 \text{ torr}}{1,069 \text{ torr}} = 0.427$$

$$X_{\text{carbon dioxide}} = \frac{112 \text{ torr}}{1,069 \text{ torr}} = 0.105$$

$$X_{\text{nitrogen}} = \frac{501 \text{ torr}}{1,069 \text{ torr}} = 0.469$$

ON YOUR OWN

11.6 A sample of artificially made air contains 150.0 grams of nitrogen and 40.0 grams of oxygen. If the pressure of the air sample is 760.0 torr, what is the partial pressure of each gas?

11.7 A chemist produces hydrogen gas in a chemical reaction and collects it over water at 21.0°C. If the pressure of the gas inside the collection vessel is 815 torr, what is the mole fraction of the hydrogen gas in the vessel?

We hope that you are beginning to see how powerful Dalton's law can be. We can use it to determine the pressures of individual components in a mixture or, more importantly, we can use the partial pressure of a gas to determine its concentration in terms of mole fraction. Make sure you understand this by answering "On Your Own" questions 11.6–11.7.

THE IDEAL GAS LAW

The last thing we need to learn about gases is probably the most important. As we have seen in other modules, the most important

thing we can learn about a substance is the number of moles of that substance. Usually, we can get this information by knowing the identity of the substance and its mass. It is very difficult to measure the mass of a gas, and it would be nice to be able to relate the number of moles of a gas to things that are easy to measure. The ideal gas law does this. The **ideal gas law** is:

$$PV = nRT \quad \text{Equation 11.12}$$

In this equation, P is the pressure of the gas, V is its volume, T is its temperature, and n is

the number of moles of that gas. R refers to a special physical constant called the **ideal gas constant**. It is a relatively important quantity in the study of chemistry.

The ideal gas constant can have several values, depending on what units are used in equation 11.12. For our purposes, we will learn only one value:

$$R = 0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mole} \cdot \text{K}}$$

As with all centered blue facts, you are expected to memorize this. Before you do that, make sure you know what it means. The ideal gas constant never changes. However, the value of the constant is different if we use different units.

For example, suppose we use m^3 as our volume unit. If we multiply length times width times height, we get volume. If the length, width, and height are measured in meters, the volume unit will be m^3 . Also, suppose we measure pressure in Pascals. We will use a different value for the ideal gas constant because we are using different units. For those units, the value of the ideal gas constant is $8.315 \text{ (Pa} \cdot \text{m}^3)/(\text{mole} \cdot \text{K})$, which can also be expressed as $8.315 \text{ J}/(\text{mole} \cdot \text{K})$. In this course, you will usually work in liters and atmospheres. That's why you were asked to memorize the value of R in those units.

Did you know that the ideal gas constant, R , can be determined by experiment? Perform experiment 11.1 to determine how close to $R = 0.0821 \text{ (L} \cdot \text{atm)} / (\text{mole} \cdot \text{K})$ you can get.

EXPERIMENT 11.1

PURPOSE: To use an ideal gas to determine the value of R .

MATERIALS:

- Kitchen sink or large bucket
- Towels or napkins
- ~~Graduated cylinder~~ ¹
- Butane lighter (A cigarette lighter with a clear case will work best.)
- Mass scale
- Thermometer
- Safety goggles
- Weather report that contains the atmospheric (sometimes called barometric) pressure for the day

QUESTION: Can the value of R be determined by experiment?

HYPOTHESIS: Write a hypothesis about how close to the theoretical value of R you can get.

PROCEDURE:

1. Fill your kitchen sink or large bucket with warm water. Allow the water to sit for half an hour so that it will attain room temperature.
2. Obtain the current atmospheric pressure where you live. You can get this information at almost any weather website like www.weather.com. If you can't find a weather report or access the Internet, assume that the atmospheric pressure is 1.00 atm. Convert the measurement to atm, if needed.

A 250 mL beaker or a measuring cup with mL markings should be substituted throughout this experiment. A 100 mL cylinder will not be large enough.

3. Submerge the butane lighter underwater and allow it to sit there for 1 minute. This will allow any air bubbles in the lighter to escape. The air bubbles could get trapped with the gas later in this experiment, causing your volume measurement to be incorrect.
4. Remove the butane lighter and dry it off with a towel or napkin. Get as much water off the butane lighter as possible.
5. Measure the mass of the butane lighter. Record the mass in your notebook.
6. After the water in the sink has sat for at least half an hour, measure the temperature of the water with your thermometer, and record your result.
7. Submerge the graduated cylinder into the sink or bucket and allow it to fill completely with water with no air bubbles. You may have to tip the cylinder up slightly to allow the air to escape.
8. Hold the graduated cylinder upside down in the water. If some of the cylinder sticks up above the water level, that is okay.
9. Submerge the butane lighter under the water and put the mouth of the lighter directly under the opening of the graduated cylinder.
10. Push the red button on the lighter to allow the gas to escape into the cylinder. If some of the gas escapes the cylinder and moves into the atmosphere, then you will need to pull the lighter out, dry it off, and measure the mass again. Replace the mass from step 5 with this measurement.
11. As the graduated cylinder is filling with gas, the water level will drop. Stop filling the cylinder when the water level is around ~~80 to 90 mL~~
12. Pull the lighter out of the water and set it aside.
13. With the cylinder opening still submerged, move the cylinder up or down in the water until the water level inside the cylinder is even with the water level outside of the cylinder. This will ensure that the pressure of the gas inside the container is equal to atmospheric pressure.
14. Read the volume of the gas when the water levels are even and record your measurement.
15. Dry off the lighter with a towel and measure its new mass. Record this mass in your notebook.
16. In your notebook, use $PV = nRT$ to solve for R .
17. To find the moles of butane, n , you need to find the mass of butane expelled from the lighter. Take the mass of the lighter from step 5 and subtract the mass measured in step 15. The result will be the mass in grams of butane in the cylinder. You will need to find the molar mass of butane, C_4H_{10} , to convert the grams of butane to moles of butane.
18. You now have all of the measurements needed to solve for R . Record the result you obtain for R in your notebook and compare it to the known value of R .
19. Clean up and return everything to the proper place.

CONCLUSION: Write something about what happened in this experiment and whether it confirmed your hypothesis.

Looking at the ideal gas equation, you can see that it relates 4 pieces of information about a gas. It relates a gas's volume, pressure, temperature, and number of moles. If we are given 3 of those 4 pieces of information, we can use equation 11.12 to determine the fourth. Notice how this works in example 11.5

200 mL. If your mass scale cannot measure more than 0.5 g differences, then double the volume of gas collected to 400 mL.

$$T_{\text{final}} = 10.49^{\circ}\text{C} + 24^{\circ}\text{C} = 34^{\circ}\text{C}$$

Notice that in this case, we are adding, so we look at decimal place. The value for ΔT goes out to the hundredths place, but the value for T_{initial} goes out to the ones place. So the answer can go out only to the ones place. The chunk of copper, then, reaches a final temperature of **34°C**.

See if you really understand this by answering “On Your Own” questions 12.3–12.4.

CALORIMETRY

Even though we mentioned that you needn't memorize any of the specific heats listed in table 12.1, there is one specific heat that you do need to memorize: the specific heat of water. This is an easy one to memorize because it relates back to the definition of a calorie. Remember that we defined a calorie as the amount of heat

necessary to raise 1 gram of water 1.0 degree Celsius. If that is the definition of a calorie, what does that tell us about the specific heat of water?

The specific heat of water is $1.000 \frac{\text{cal}}{\text{g} \cdot ^{\circ}\text{C}}$ or $4.184 \frac{\text{J}}{\text{g} \cdot ^{\circ}\text{C}}$

These 2 numbers are very important to chemists, so you need to memorize them.

So how is specific heat measured? ~~Is there an instrument that will allow us to measure specific heat? Unfortunately, there is no such instrument! But by~~ looking at equation 12.1, we see that 3 measurements are needed to calculate specific heat. If we measure the mass of an object, its temperature change, and how much energy it has experienced, then we can calculate heat with those measurements. How can we measure mass then? We use a balance to measure grams! Okay, that was easy, but how can we measure ΔT ? We need 2 different temperatures. We measure temperature with a thermometer. This works well for liquids and gases, but what about solids like gold or copper? How can we measure the temperature of solids? Thermometers don't work well for this.

What are we going to do? We will measure the temperature of the solid's environment. If we allow the metal to sit in one place for a while, then heat will transfer between the metal and its environment until they are the same temperature. To determine the specific heat, then, we will need 2 different environments at 2 different temperatures. Then ΔT will be the difference between the 2 temperatures.

But now we have reached the hard part of determining the specific heat of an object. We need to know how much energy that object is holding. In equation 12.1, we need to know q . There is no such instrument that can measure heat, so how are we going to measure q ? We are going to measure how much heat the object's environment absorbs because this amount of heat had to come from the object! The experimental process that chemists use to do this is called **calorimetry** (kal uh rim' uh tree) and is illustrated by figure 12.3.

ON YOUR OWN

12.3 A 432 g aluminum rod is heated to an unknown initial temperature. If, while cooling down to a temperature of 24.0°C , the rod releases 612 Joules of energy, what was its initial temperature? (Hint: When an object loses energy, its q is negative. We will discuss this fact in detail shortly.)

12.4 If a 16.0 g object requires 85 J of heat to raise its temperature by 22°C , what is the specific heat of the object?

PRACTICE PROBLEMS

1. The average person uses 2,500.0 Cal of energy per day. How many Joules is that? Remember that a food calorie (Cal) is 1,000 chemistry calories (cal).
2. How many Joules does it take to heat 15.1 kg of glass from 15°C to 45°C? (Refer to table 12.1.)
3. 124.1 g of an unknown substance absorbs 50.0 kJ of heat and increases its temperature by 36.3°C. What is its specific heat? (Remember that k is the abbreviation for kilo, so kJ stands for kiloJoules.)
4. A 245 g piece of copper at room temperature (25°C) loses 456.7 Joules of heat. What will its final temperature be?
5. Review question 7 mentioned that drinking ice-cold water is a way of burning excess Calories. Calculate how many Calories are burned when a 12-ounce (3.40 x 10² g) glass of water at 0.0°C is warmed up to body temperature (37.0°C).
6. How many grams of silver are needed to absorb 1.00 kJ of energy changing the temperature from 33.0°C to 41.8°C?
7. A calorimeter is filled with 150.0 g of water at 24.1°C. A 50.0 g sample of a metal at 100.0°C is dropped in this calorimeter and causes a temperature increase of 5.4°C. What is the specific heat of the metal? Ignore the calorimeter in this problem.
8. Now let's take the calorimeter into consideration for problem 7. The calorimeter has a mass of 2.0 g and a specific heat of 3.5 J/(g · °C). Recalculate the specific heat of the metal.
9. A 150.0 g sample of copper [0.3851 J/(g · °C)] was heated to 200.0°C. It is dropped into a 5.0 g calorimeter [c = 50.00 J/(g · °C)] with 175.0 g of liquid. The liquid was initially at 25.0°C, and the temperature increased to 31.2°C by the end. What is the specific heat of the liquid?
10. A 345.1 g sample of copper at 100.0°C is dropped into a 4.5 g calorimeter made of an unknown substance. If the calorimeter has 150.0 grams of water in it and the temperature changed from 24.2°C to 25.1°C, what is the specific heat of the calorimeter? (Refer to table 12.1.)

then ΔG has to be negative. How can this happen? If ΔH is negative and ΔS is positive, the first term in equation 13.6 will be negative and the second term ($T \cdot \Delta S$) will be positive. If a negative term has a positive term subtracted from it, then the result will *always* be negative. Therefore, whenever a reaction is exothermic (has a negative ΔH) and also has a positive ΔS , it will *always* be a spontaneous reaction. On the other hand, if ΔH is positive and ΔS is negative, ΔG will *always* be positive. Therefore, reactions that are endothermic and have a negative ΔS will *never* be spontaneous.

This is one qualitative way of determining whether or not a reaction is spontaneous. If we know whether the reaction is exothermic or endothermic, we can look at the equation and determine the sign of its ΔS . If the reaction is endothermic and ΔS is negative, the reaction will never occur. If the reaction is exothermic and ΔS is positive, the reaction will always be spontaneous.

The 2 possibilities we just discussed are not the only possibilities for a chemical reaction. Suppose the ΔH of a chemical reaction is negative and the ΔS is also negative. In this case, the first term in equation 13.6 will be negative, but the $T \cdot \Delta S$ term will also be negative. This means that we will have one negative number subtracted from another negative number. Sometimes, the result will be positive (indicating a nonspontaneous reaction), and sometimes the result will be negative (indicating a spontaneous reaction). In those cases, we have to calculate a value of ΔG to determine whether or not the reaction is spontaneous.

Of course, calculating a value for ΔG isn't too hard since we already know how to calculate ΔH and ΔS . One way to calculate ΔG is to calculate ΔH and ΔS for a reaction and plug the result into equation 13.6. There is another way to calculate ΔG . Because ΔG is a state function, we can use Hess's law. Earlier in this module, our job was made easier by using standard enthalpies and standard absolute entropies. Similarly, chemists have also calculated standard Gibbs free energies of formation (ΔG_f°). Using ΔG_f° from the table below with Hess's law, we can calculate standard Gibbs free energies (ΔG°).

TABLE 13.5
Standard Gibbs Free Energies of Formation in kJ/mole

| Substance | ΔG_f° | Substance | ΔG_f° | Substance | ΔG_f° | Substance | ΔG_f° |
|-----------------------------------|--------------------|----------------------|--------------------|-------------------------------------|--------------------|-------------------------------------|--------------------|
| CH ₄ (g) | -50.8 | NO ₂ (g) | 240 | CaSO ₄ (s) | -1,320 | C ₂ H ₆ O (l) | -175 |
| C ₆ H ₆ (l) | 125 | H ₂ O (l) | -237 | CaCO ₃ (s) | -1,129 | CaCl ₂ (s) | -750 |
| CO (g) | -137 | H ₂ O (g) | -229 | H ₂ CO ₃ (aq) | -623 | HCl (aq) | -131 |
| CO ₂ (g) | -394 | CaO (s) | 38.2 | CH ₃ OH (l) | -166 | Ca(OH) ₂ (s) | -897 |

The numbers in this table will allow us to use equation 13.7 as a means of determining ΔG :

$$\Delta G^\circ = \sum \Delta G_f^\circ (\text{products}) - \sum \Delta G_f^\circ (\text{reactants}) \quad \text{Equation 13.7}$$

Remember, the ΔG_f° of any element in its elemental form is 0, as is its ΔH_f° . Since we have 2 different methods for calculating ΔG , does it matter which we use? Yes, it does. You see, the values given for ΔG_f° in table 13.5 are all for standard conditions: 298 K and 1 atm of pressure (pressure of Earth's atmosphere). So if we are working with a reaction at stan-

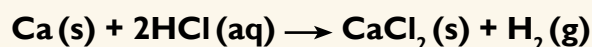
$$T \times 0.147 \frac{\text{kJ}}{\text{mole} \cdot \text{K}} < 114.2 \frac{\text{kJ}}{\text{mole}}$$

$$T < \frac{114.2 \frac{\text{kJ}}{\text{mole}}}{0.147 \frac{\text{kJ}}{\text{mole} \cdot \text{K}}}$$

$$T < 777 \text{ K}$$

For any temperature less than **777 K**, this reaction is spontaneous.

What is the ΔG for the following reaction at standard conditions? Is it spontaneous?



This problem asks us for the ΔG of a reaction at 298 K and 1 atm. This means we can use table 13.5 and equation 13.7 to answer it. Recognizing that the ΔG_f° 's for both Ca (s) and H₂ (g) are 0, Equation 13.7 becomes:

$$\Delta G = (1 \text{ mole} \times -750 \frac{\text{kJ}}{\text{mole}}) - (2 \text{ moles} \times -131 \frac{\text{kJ}}{\text{mole}}) = 490 \text{ kJ}$$

This reaction, then, is **spontaneous** at 298 K.

ON YOUR OWN


13.11 The ΔH° of a certain chemical reaction is 20.0 kJ, and the ΔS° is 123 kJ/mole. At what temperatures is this reaction spontaneous?

13.12 Is the following synthesis of ethyl alcohol a spontaneous reaction at 298 K?

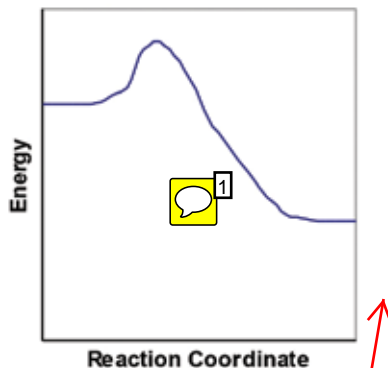


As we have cautioned you in the past, do not get so wrapped up in solving problems that you forget the chemical significance of what you are doing. Remember, ΔG is so important because it takes inventory of all of the entropy of both the chemical reaction and its surroundings. By doing that, we can use it to determine whether or not any chemical equation is consistent with the second law and therefore whether or not it will occur

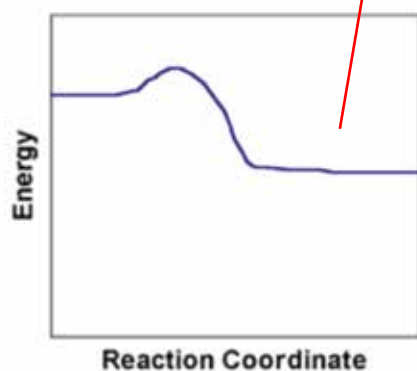
naturally. Finish this module with "On Your Own" questions 13.11–13.12.

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For reaction B, we know that it was easy to start. This means that the difference in energy between the reactants and the intermediate state is small. In other words, this energy diagram has a small hump. Secondly, although the reaction is exothermic, it doesn't release much energy. The reactants will have a higher energy than the products, but not by much:





For reaction C, we know that it was hard to start. This means that the difference in energy between the reactants and the intermediate state is large. In other words, this energy diagram has a large hump. Second, although the reaction is endothermic, it absorbs a lot of energy. The reactants have a lower energy than the products by a large distance.:




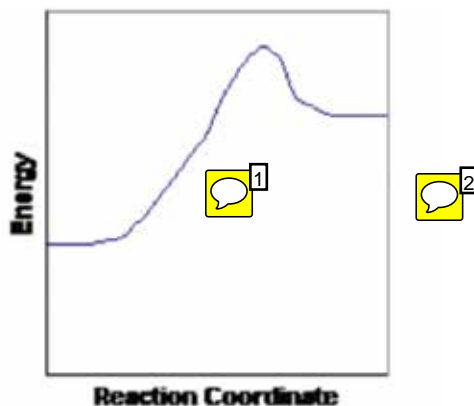
For reaction D, we know that it was easy to start. This means that the difference in energy between the reactants and the intermediate state is small. In other words, this energy diagram has a small hump. Second, although the reaction is endothermic, it doesn't absorb much energy. The reactants will have a lower energy than the products, but not by much:



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13.9 (a) This reaction has 1 gas molecule on the reactant side of the equation and 0 on the product side. This means that the reactants have more entropy than the products, so ΔS is **negative**.

(b) In this reaction, there are 2 molecules of gas on the product side and 3 on the reactant side. The reactants have higher entropy than the products, so ΔS is **negative**.

(c) This reaction has no molecules or atoms in the gas phase, so we must look at the liquid (or aqueous) phase. There are 3 molecules in the aqueous phase on the product side and only 1 molecule in the aqueous phase on the reactant side. As a result, ΔS is **positive**.

13.10 Table 13.4 has all of the information we need, so we just set this up like a Hess's law problem:

$$\Delta S^\circ = (1 \text{ mole} \times 92.9 \frac{\text{J}}{\text{mole} \cdot \text{K}}) - (1 \text{ mole} \times 38.2 \frac{\text{J}}{\text{mole} \cdot \text{K}}) - (1 \text{ mole}) \times 214 \frac{\text{J}}{\text{mole} \cdot \text{K}} = -159 \frac{\text{J}}{\text{K}}$$

13.11 This problem gives us ΔH and ΔS and asks us at what temperature the reaction is spontaneous. In other words, we need to see for what temperatures ΔG is negative.

$$\Delta H - T\Delta S < 0$$

To use this equation, though, we need to get our units consistent:

$$\frac{123 \text{ J}}{\text{K}} \times \frac{1 \text{ kJ}}{1,000 \text{ J}} = 0.123 \frac{\text{kJ}}{\text{K}}$$

Now we can use the equation:

$$20.0 \text{ kJ} - T \times 0.123 \frac{\text{kJ}}{\text{K}} < 0$$

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The diagram that fits here can be found on page 470. It is the one labeled Diagram III. Please use it and add the same words as seen on these other ones: "Energy and Reaction Coordinate" All the other words and numbers are NOT needed.

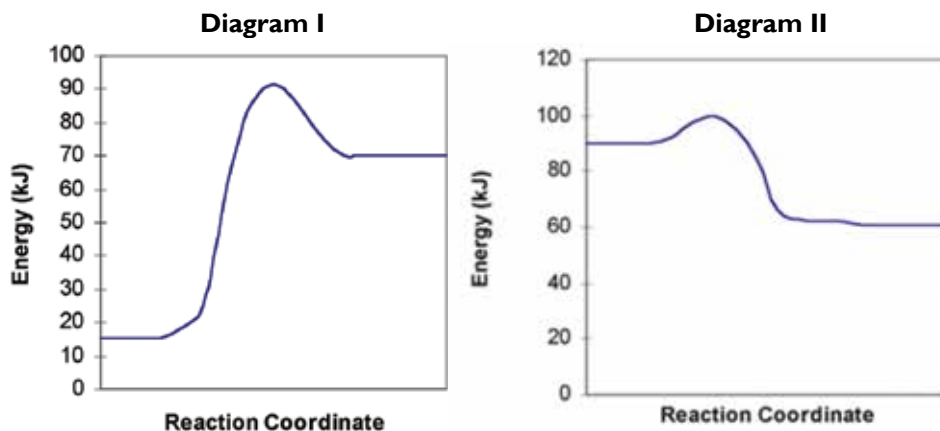
STUDY GUIDE FOR MODULE 13

REVIEW QUESTIONS

1. In a chemical reaction, where is the potential energy? Where is the kinetic energy?
2. If a chemical reaction has a positive ΔH , will the beaker that contains the reaction feel hot or cold once the reaction is finished?
3. Why is Hess's law a more exact way of determining ΔH than the technique that uses bond energies?
4. What is a state function? Give 2 examples.
5. Which of the following substances will have a ΔH_f° of 0?

NaOH (aq) , $\text{Na}^+(\text{aq})$, $\text{O}_2(\text{g})$, $\text{O}(\text{g})$, $\text{Cl}_2(\text{g})$, $\text{H}_2(\text{l})$

6. Which of the following diagrams indicates an endothermic reaction? What is the ΔH of both reactions?

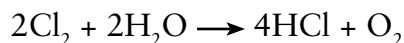


7. Draw an energy diagram for a reaction that has a large activation energy and a ΔH of 0.
8. If you have 2 50.0 kg blocks of copper, and one is at a temperature of 50°C while the other is at a temperature of 25°C , which has more entropy?
9. If a chemical reaction is exothermic but has a negative ΔS , what could you do to the temperature to make it possible to run the reaction?

PRACTICE PROBLEMS

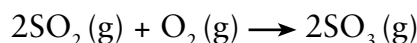
(You will need to use tables 13.2–13.5)

1. The chlorine smell of pool water is not, in fact, due to chlorine. When chlorine hits water, some of it reacts with the water in the following way:



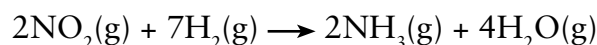
The smell of a chlorinated pool is often the smell of the HCl produced in this reaction. What is the ΔH of this reaction?

2. Benzene (C_6H_6) is a toxic liquid with a ΔH_f° of 49.00 kJ/mole. The best method for disposing of benzene is to burn it in the presence of excess oxygen. What is the ΔH° for this reaction? All phases are gases except for benzene, so use Hess's law.
3. When coal is burned in a power plant, it usually contains a small amount of sulfur contamination. The sulfur burns along with the coal to make SO_2 . This sulfur dioxide then begins to travel up the smokestack, where it reacts with oxygen as follows:



If the ΔH° of this reaction is -198 kJ, what is the ΔH_f° of $\text{SO}_2(\text{g})$?

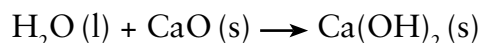
4. Using the ΔH you got in problem 2, calculate the number of kJ of energy released when 125.0 grams of benzene are burned in excess oxygen.
5. Suppose you synthesized ammonia with the following reaction:



How much energy will be released if 70.0 g of NH_3 are formed?

6. What is the sign of ΔS for the following reactions?
- $\text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l}) + \text{CO}_2(\text{g}) \rightarrow \text{NaHCO}_3(\text{aq}) + \text{HCl}(\text{aq})$
 - $\text{NH}_4\text{Cl}(\text{s}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{HCl}(\text{g}) + \text{NH}_4\text{OH}(\text{g})$
 - $\text{SF}_6(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{SO}_2(\text{g}) + 3\text{F}_2(\text{g})$

7. What is the ΔS° of the following reaction?



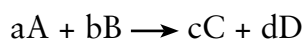
8. If the ΔH of a certain reaction is -1,023 kJ/mole and the ΔS is -324 J/(mole·K), what is the temperature range for which this reaction is spontaneous?

the equation we gave you at the beginning of this module for calculating the rate of a reaction? If you don't, turn back a few pages and review equation 14.1. This equation is very limited in its result of calculating the rate, R . The rate calculated using equation 14.1 is good only ~~for the time that was in the denominator.~~ So if we calculate R after 30 seconds, then that R from equation 14.1 is only good at that 30-second mark.

Why is that? What is happening to the concentrations of the reactants as time increases? Our experiment results showed that R gets slower as concentrations get smaller. So we should expect that an R calculated at 45 seconds would be smaller than an R calculated at 30 seconds.

Is there a way to figure out what the rate would be at any time during a chemical reaction? Yes. Chemists can write mathematical equations to represent the rate of a chemical reaction at any time. These mathematical equations are based on the concentration of the reactants in the chemical equation since the concentration of the reactants in a chemical reaction has a profound influence on the rate of the chemical reaction.

We will use the following generic chemical equation to represent all chemical reactions.



This expression represents a general chemical equation that has reactants A and B, which react to form the products C and D. The lowercase letters (a, b, c, and d) represent the stoichiometric coefficients for each of the substances (A, B, C, and D) in the chemical equation.


For this generalized chemical reaction, we can develop a mathematical equation that we call the **rate equation**:

$$R = k[A]^x[B]^y \quad \text{Equation 14.2}$$

In this equation, R represents the rate of the chemical reaction, $[A]$ represents the concentration of reactant A, $[B]$ stands for the concentration of reactant B, and k is called the **rate constant** for the chemical reaction. The letters x and y represent exponents for the concentrations of the reactants.

If all of this seems rather confusing, don't worry. We first have to pick the rate equation apart bit by bit to understand each of its components. Let's start with the 2 concentration terms. If we look at equation 14.2, what happens to the reaction rate (R) when the concentration of either reactant ($[A]$ or $[B]$) increases? Based on the equation, we would expect R to increase. So this equation accurately predicts a fact that we have already learned: When reactant concentration increases, the rate of the chemical reaction increases. Remember that when using reactants to calculate the reaction rate (equation 14.1), we used a negative sign to indicate the loss of reactants over time. The rate equation (equation 14.2) addresses only the overall speed of the reaction; we don't add a negative sign when using reactants.

Of course, chemists are detail-oriented people. We not only want to know that reaction rate increases when reactant concentration increases, but we also want to know how much it increases. For example, if we double the concentration of a reactant in a chemical equation, will the reaction rate double? Will it triple? Will it quadruple? In fact, all of those things are possible. For some chemical reactions, doubling the concentration of a reactant doubles the reaction rate. For others, doubling the concentration of a reactant quadruples the reaction rate. For other reactions, doubling the concentration

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as an average over the time interval for which delta t was calculated.

The ionization constant, then, is just the equilibrium constant of this reaction. Remember that we must ignore water here since it is in its liquid phase:

$$K_b = \frac{[\text{OH}^-][\text{CH}_6\text{N}^+]}{[\text{CH}_5\text{N}]}$$

Since the value given for this ionization constant is less than 1, **this would be considered a weak base.**

ON YOUR OWN

15.11 What is the equation for the acid ionization constant of HClO_2 ?

15.12 What is the equation for the base ionization constant of PH_3 ?

THE pH SCALE

Since the strength of an acid is really dependent on how many H^+ ions it can make in an aqueous solution, chemists have developed a special scale to help us classify acids and bases. This scale is called the **pH scale**. The pH scale was developed by the Danish biochemist S. P. L. Sørensen in 1909 and is

an abbreviation in French for *pouvoir d'hydrogene* or in English, **power of hydrogen**.

The pH scale ranges from 0 to 14 and gives an idea of the concentration of the hydrogen ion. Having a scale of 0 to 14 is a lot easier to work with than a range of actual concentration of 1 to 10^{-14} (100 to 10^{-14}). The pH scale is a convenient way to describe the concentration of hydrogen (H^+) and hydroxide ions (OH^-). When a solution has a pH of 0, we consider it a very strongly acidic solution. Therefore, it has a lot of H^+ ions. When a solution has a pH of 14, it is strongly basic and has a lot of OH^- ions. When a solution is purely neutral (equal acid-base properties), it has a pH of 7. Generally speaking:

Solutions with pH 0-1.9 are considered strongly acidic; solutions with pH 2-6.9 are weakly acidic; solutions with pH 7.1-12 are weakly basic; and solutions with pH 12.1-14 are strongly basic. A pH of 7 indicates a neutral solution.

In other words, the lower the pH, the more acidic a solution is; the higher the pH, the less acidic (and therefore more basic) the solution is.



If you purchased the MicroChem kit mentioned in the Student Notes, you can perform experiment 11 to get more experience with the pH scale and pH indicators.

It's important for you to understand that pH is a measure of the *overall* acidity of a solution. For example, the pH of a 1.0 M solution of acetic acid is about 2.3. As mentioned before, acetic acid is a weak acid. However, the pH of a 0.0050 M solution of HCl is *also* 2.3, despite the fact that HCl is a strong acid. Why does the solution made with a weak acid have the same pH as a solution made with a strong acid? Look at the concentrations. The weak acid has a concentration of 1.0 M, while the strong acid has a concentration that is 200 times smaller. That's what's nice about pH. It is a measure of the acidity of a *solution*, which takes both the strength of the acid and

M

ON YOUR OWN

16.1 What is the oxidation number of phosphorus in each of the following substances? **1**

- P
- P^{3-}
- P_2O_5

Answer “On Your Own” question 16.1 to make sure you understand this part of calculating oxidation numbers.

DETERMINING OXIDATION NUMBERS

Of course, you need to be able to calculate the oxidation numbers of *all* atoms in a molecule. We really shouldn't have to tell

you any of them. We just had you answer “On Your Own” question 16.1 as a warm-up to the real thing, which is often a bit difficult. You see, the best way to calculate the oxidation numbers of all atoms in a molecule is to draw a Lewis structure of the molecule and then give the shared electrons to the most electronegative atoms. Then we can count the electrons on each atom and determine the charge they would have under those conditions. This is what we did in the previous section to show you the oxidation numbers of H and Cl in HCl.

The problems with figuring out oxidation numbers this way are pretty enormous. First, many of the compounds that participate in the reactions we are going to study have extremely difficult Lewis structures, so first-time chemists often get very frustrated figuring them out. Also, even though you learned in module 4 how to tell which atoms are more electronegative than others, the rules you learned are often broken, making it very difficult to determine which atoms get the shared electrons. In the end, the Lewis structure method for determining oxidation numbers cannot really be taught to chemistry students at either the high school or college level.

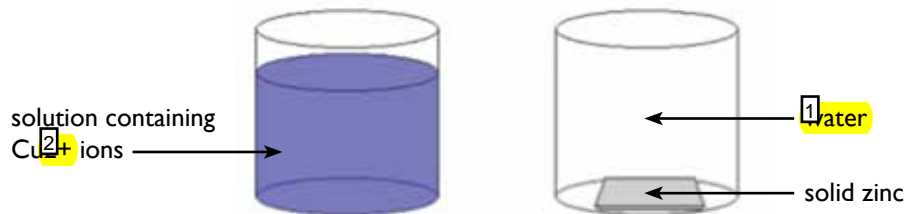
Are we stuck then? Not really. Chemists have developed rules of thumb that allow us to determine the oxidation numbers of certain atoms. Once you learn those rules, you can usually use those oxidation numbers to begin your calculations, which will then help you calculate the oxidation numbers of other atoms. The first 4 rules we will use in this text are:

1. When a substance has only 1 type of atom in it (F_2 , O_3 , Ca^{2+} , or Mg for example), the oxidation number for that atom is equal to the charge of the substance divided by the number of atoms present.
2. Group 1A metals (Li, Na, K, Rb, Cs, and Fr) always have oxidation numbers of +1 in molecules that contain more than 1 type of atom.
3. Group 2A metals (Be, Mg, Ca, Sr, Ba and Ra) always have oxidation numbers of +2 in molecules that contain more than 1 type of atom.
4. Fluorine always has a -1 oxidation number in molecules that contain more than 1 type of atom.

What's nice about these 4 rules is that they are *always true*. The problem is that after these 4 rules, things get a little fuzzy because there are a few exceptions to the next rules. In general, though, the next rules *usually* apply:

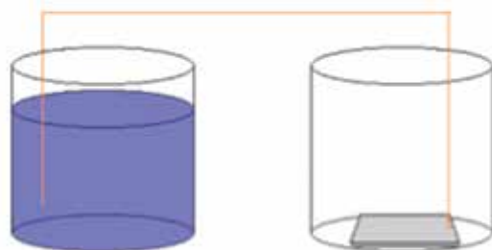
5. When grouped with just 1 other atom that happens to be a metal, H has an

Rather than mixing the 2 chemicals, what if we just put them in 2 separate beakers and placed the beakers next to each other?

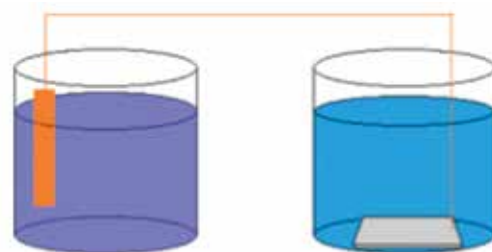


Would anything happen? Of course not! For the reaction to occur, the reactants would have to get close enough together to exchange electrons. Since they can't even touch each other, there's no way to get the electrons transferred. Or is there?

Suppose we were to put the 2 chemicals in contact, but not directly. Instead of mixing the chemicals together, we connect them with a wire called a conductor that can transport the electrons.



All that is left to do now is to include the products. As zinc loses its electrons, Zn^{2+} is produced, but Zn^{2+} cannot exist by itself. Zinc ions can exist in water, so we need to fill the zinc beaker with water. Also, the copper ions will be turning into solid copper. This process can be helped out by putting a piece of solid copper in the beaker.



Believe it or not, under these circumstances, *the reaction would proceed!* To get this reaction to work, all we need to do is get electrons from the zinc to the Cu^{2+} . The wire allows this to happen because it conducts electrons. Electrons will flow out of the zinc electrode, through the wire, into the copper strip, and into the Cu^{2+} solution. When they leave the zinc, the zinc becomes Zn^{2+} . As the electrons arrive in the Cu^{2+} solution, the Cu^{2+} turns into Cu. The reaction occurs even though the reactants are in different containers! If we watched this setup for a while, we would notice that the zinc strip would slowly disappear as it turned into Zn^{2+} and dissolved into solution. We would also notice that the reaction would stop.

When the first few electrons leave the solid zinc, it turns into Zn^{2+} . As a result, the

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BOTH arrow and water need to move down to the third illustration where the water is actually in the beaker.

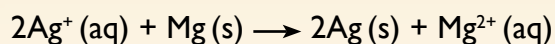
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superscript 2+

work. First you need to be comfortable with analyzing a Galvanic cell. Given the chemical equation, you need to be able to draw a figure similar to figure 16.3, label the anode and cathode, and identify the electron flow. This will take a bit of practice, so study example 16.5, perform experiment 16.2, and answer “On Your Own” questions 16.6–16.7.

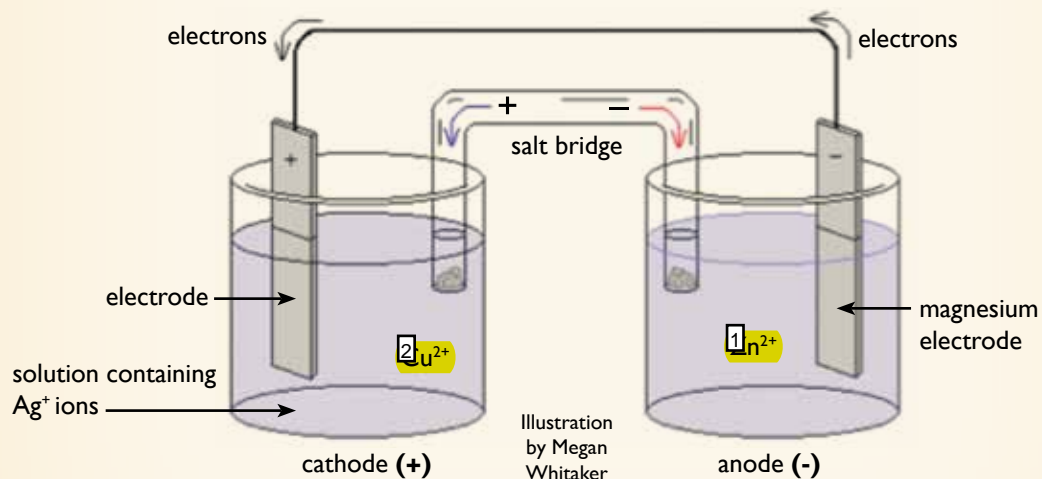
EXAMPLE 16.5

A Galvanic cell runs on the following chemical reaction:



Draw a diagram of the Galvanic cell, labeling all reactants and products, the electron flow, the anode, the cathode, and the positive side and negative side of the Galvanic cell.

The first thing that we have to do is look at the reactants and see which is being oxidized and which is being reduced. Since Ag^+ turns into Ag , it must be gaining electrons, so it is reduced. This means that the container that holds the Ag^+ will have the electrons flowing into it. The Mg , on the other hand, loses electrons because it goes from an oxidation number of 0 to an oxidation number of +2. This means that the Mg is oxidized and the container that holds it will have electrons leaving it. Since the electrons are flowing from the Mg to the Ag^+ , the container that holds Mg is negative (electrons are repelled by negative charge) and is therefore the anode. The container that holds the Ag^+ is positive (electrons are attracted to the positive charge), and is therefore the cathode. In the equation, Ag^+ is aqueous. This means that the Ag^+ is dissolved, so the solution that holds it will be clear. The Mg , on the other hand, is solid, so there should be a magnesium electrode in its beaker. That will be the source of the solid magnesium. As always, there must be a salt bridge as well. In the end, then, the picture should look something like this:



We did not label the electrode that is in the Ag^+ solution because it does not need to be any specific metal. It just needs to conduct electricity. The electrode on the right must be magnesium because solid magnesium is a reactant in the equation. The presence of solid magnesium helps the Mg^{2+} ions turn into Mg solid. Since the only other reactant needed is Ag^+ , and since Ag^+ is in the solution, there is no need for the electrode in that solution to be silver. It can be any metal. Usually, we do use the metal version of the ion in that ion's beaker, but since silver is expensive,

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Replace with: Mg²⁺
superscript the 2+


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replace with: Ag⁺
superscript the +

EXTRA PRACTICE PROBLEMS FOR MODULE 5

1. Give the chemical formulas for the following compounds:
 - a. lithium sulfite
 - b. ammonium phosphate
 - c. calcium carbonate
 - d. aluminum dichromate
2. Name the following compounds:
 - a. NH_4ClO_2
 - b. $\text{Mn}(\text{NO}_2)_2$
 - c. $\text{Be}_3(\text{PO}_4)_2$
 - d. NiPO_4
3. Determine the shape of a CS molecule. Give its bond angle and draw a picture of it.
4. Determine the shape of a SiI_4 molecule. Give its bond angle and draw a picture of it.
5. Determine the shape of a NF_3 molecule. Give its bond angle and draw a picture of it.
6. Determine the shape of a CCl_2S molecule. Give its bond angle and draw a picture of it.
7. Determine the shape of a SF_2 molecule. Give its bond angle and draw a picture of it.
8. Identify each of the molecules in problems 3–7 as polar covalent or **purely** 1] covalent.
9. Classify each of the following molecules as ionic, polar covalent, or **purely** 2] covalent:
 - a. SiFCl_3
 - b. SiS_2
 - c. AlF_3
 - d. N_2
 - e. CCl_4
 - f. PF_3
10. Which of the substances in problem 9 would you expect to dissolve in water?

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nonpolar

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nonpolar