Chapter 11: The Mole: Not Just a Defective Rat Anymore

Hen mole with rice and avocado, with tamalitos on the side. Wikipedia says that mole is very popular in Guatemala.

http://commons.wikimedia.org/wiki/File:199328_mole.JPG

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If you've ever heard anything about chemistry, you've probably heard about the mole.[1] Chemists just go on and on about moles as if they're in love with moles and want to marry them. Which may be the case if you're talking about the hairless rat, but that's an issue for a psychology book, not a chemistry book. We try not to think about these things.

Figure 11.1: *A sexy, sexy mole in its native environment, which I think is in Poland somewhere. Seriously, who cares?*

http://commons.wikimedia.org/wiki/File:Kret_mole.jpg

Anyway, we're going to learn more about the chemical concept of the mole in this chapter. For more information about the other types of mole, consult my colleague Dr. Google.

Section 11.1: What's a Mole?

The term "mole" just refers to a particular number. Just as "pair" means "two" and dozen means "12", the word "mole" means 6.02×10^{23}. This is a very big number – to visualize it, imagine a whole lot of shoes. Like, a ton of shoes. This is nowhere near a mole of shoes.

The word mole is really important because it allows us to figure out how much of a substance we have, in molecules. After all, if we consider the equation $2\ H_2 + O_2 \rightarrow 2\ H_2O$, all of the terms in this equation are given in "molecules" and not in "grams." If we want to work with this reaction, we need to figure out how many molecules there are.

Q: If moles are so important, why haven't I heard of them before?

A: You haven't heard of it before because it doesn't make sense to use moles in your everyday life. "Pairs" you use, because you frequently need two of something. "Dozens" you know because it's occasionally handy to have 12 things. However, there are no items in your everyday life that you will ever need 6.02×10^{23} of. After all, a mole of pencils would weigh six thousand million million million pounds. Unless you do a lot of writing, this probably won't come in handy. On the other hand, if you've got 6.02×10^{23} of something really tiny (such as atoms or molecules), moles start being useful.

[1] This statement probably isn't true. When people think of chemistry they either think of "blowing stuff up" or "cooking crystal meth." I think that the concept of the mole is the *last* thing on people's minds. Unfortunately, this makes for a bad beginning to a chapter, so I'll stick with my original, wrong, statement.

Now, don't get the wrong idea: We're not going to use the number of molecules in any calculations – that would be stupid and pointless, given how small a molecule is. Instead, we're going to use something that works the same way as a molecule in the equations. How does this work? I'm glad you asked!

Let's say that we want to make a really big bouquet for a wedding.[2] If the bride specifies that she wants 36 daisies, 12 roses, and 24 of some other flower[3], we can either go to the florist and ask for the numbers she specified, or we could go to the florist and ask for 3 dozen daisies, a dozen roses, and 2 dozen of whatever that other flower is. Converting the raw numbers to "dozens" makes our life a little easier in this case.

In the same way, we use moles to make our lives easier when doing chemical reactions. Instead of saying that we want 1.20×10^{24} molecules of hydrogen and 6.02×10^{23} molecules of oxygen to make water, we can just say that we want 2 moles of hydrogen and 1 mole of oxygen. Converting to moles doesn't change the reaction, but it does make it a lot less annoying. In other words, those coefficients in front of the compound formulas in chemical equations don't just refer to molecules, they also refer to moles.

Chemistry Fun Fact

So, why is the number of things in a mole equal to 6.02×10^{23}? It's so that the number of grams in a mole of an element will be equal to the mass of one atom of the same element in atomic mass units (amu).

Section 11.2: Converting Moles to Grams (and Vice-Versa)

Before we can make the conversion above, we need to first explore the idea of the "molar mass" of a compound. **Molar mass** is equivalent to, you guessed it, the mass of one mole of whatever compound we're talking about.

How do we compute molar mass? To do this, we just add up the weights of all of the atoms in the compound, using our handy periodic table. For example, if we're trying to find the molar mass of H_2O, we do the following calculation:

(2 moles of H)(1.01 grams/mole) = 2.02 grams of H

(1 mole of O)(16.00 grams/mole) = 16.00 grams of O

When you add these together, you find that the molar mass of water is 18.02 g/mol. For practice, verify for yourself that the molar mass of CO_2 is 44.01 g/mol and the molar mass of CH_4 is 16.01 g/mol.[4]

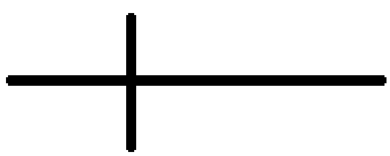
[2] Assuming we know something about flowers.
[3] I don't know anything about flowers.
[4] Depending on your teacher, he or she might want you to round your answer differently than I have. Some teachers (like me) require that you round only to the tenth of a gram, so ask before your quiz.

Once you know how to find the molar mass of a compound, it's time to do some conversions between grams of a compound and moles.

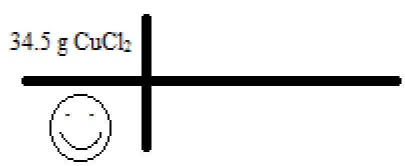
> ***Why Are We Doing This, Anyway?***
>
> *The reason we need to convert between grams and moles is that the only convenient way of measuring compounds is by mass. This is for two reasons: 1) It's very hard to count individual molecules/atoms of a compound; 2) You'd have to do an awful lot of counting to make any meaningful quantity of chemical. Rather than trying to count a whole bunch of water molecules, it makes more sense to just calculate moles from grams directly.*

Let's use the following sample problem: Convert 34.5 grams of copper (II) chloride to moles. We can solve this (or any other mole problem) by following these steps:

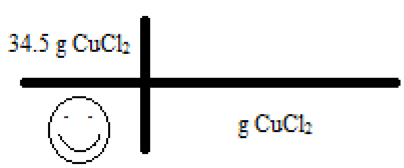
1) Find the formula and molar mass of copper (II) chloride. Given that the formula is $CuCl_2$, the molar mass is 134.45 g/mol.
2) Follow the t-chart method of solving conversion problems that you learned in section 2.2 by following the next few steps.[5]
3) Draw a t:



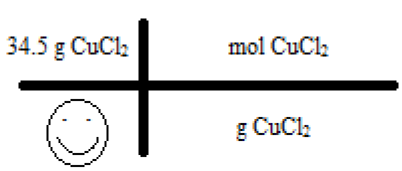
4) Put whatever you've been given in the problem in the top left:[6]



5) Put the units of whatever you've been given in the bottom right:



6) Put the units of whatever you're trying to find in the top right:



7) Put "1" in front of "moles":

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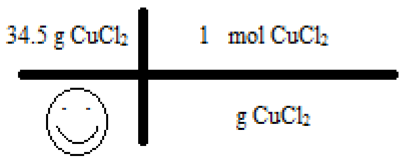
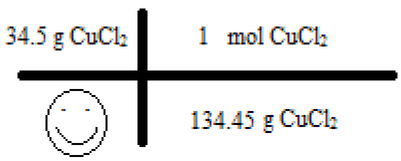
[5] This is basically a conversion problem between grams and moles, so the same method can be used.

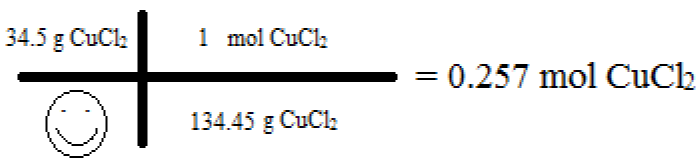
[6] Smiley face optional



8) Write the molar mass of the compound in front of "grams":



9) Solve this problem like a fraction, by multiplying the terms on top together and dividing by the term on the bottom. This will give you your answer in the unit shown in the top right:



The same process is used to convert between moles and grams, except that the units are switched around. However, if you follow the same steps, it won't make any difference and you'll still get the right answer.

Famous Moles Throughout History

Figure 11.2: *Chris Mole is the Parliamentary Under Secretary of State for Transport of the United Kingdom and a former member of Parliament. For copies of his statements and speeches, visit http://webarchive.nationalarchives.gov.uk/20100115161848/http://www.dft.gov.uk/press/ministers/chrismole?view=Alt_1 .*

http://webarchive.nationalarchives.gov.uk/20100115161848/http://www.dft.gov.uk/press/ministers/chrismole

Section 11.3: Useless Stuff We Won't Talk About

At this point, many textbooks discuss a couple of stupid topics that nobody actually needs to know. We're not going to do this, but we will discuss two of these topics and the reason for their stupidity:

- **Conversion between moles and molecules or ions:** There is no purpose whatsoever for converting between moles and molecules, as you could never count enough molecules to make even the tiniest fraction of a mole. Likewise, there's no reason to know the number of ions present in a block of an ionic compound, as we typically don't spend our time counting ions, either.

- **Empirical formulas:** These reduced versions of molecular formulas haven't been used in actual science for a very, VERY long time. Modern equipment has made calculations involving empirical formulas obsolete, so there's really no reason to learn about them.[7]

Section 11.4: Percent Composition

Let's say that you're really bored and want to find out how much calcium is in your calcium supplement.[8] To do this, you'd need to find the percent composition of calcium that's present.

Figure 11.3: *Pure calcium is a poor choice for a dietary supplement, given its strong reactivity with water.*

http://commons.wikimedia.org/wiki/File:Calcium_metal2.jpg

For the sake of argument, let's say that your calcium supplement contains calcium carbonate.[9] How much calcium is in 750.0 mg of this supplement?

To figure this out, we first need to figure out how much of the supplement is actually calcium. This means that we'll need to find out what percent of the supplement is calcium, a value known as its percent composition. We do this by dividing the mass of calcium in the compound by the molar mass of the compound:

$$\begin{array}{c}\text{percent composition}\\ \text{of calcium}\end{array} = \frac{\text{mass of calcium in } CaCO_3}{\text{molar mass of } CaCO_3} = \frac{40.08 \text{ g Ca}}{100.09 \text{ g } CaCO_3} = 40.04\%$$

Once we have a percent composition, it's easy to figure out how much calcium is actually in the supplement. If 40.04% of the supplement is actually calcium, then the total mass of calcium in the supplement is 40.04% of 750.0 mg, or 300.3 mg.

[7] About the only time you ever see empirical formulas used is in determining the chemical formulas of ionic compounds. Since you already know how to do this, why learn it again?
[8] Seriously, we're talking *really* bored.
[9] Believe it or not, most calcium supplements are really just nothing more than purified $CaCO_3$, a.k.a "chalk."

The Magical World of Hydrates

Figure 11.4: The hydration of copper (II) sulfate. Hydrates are chemical compounds that have water molecules loosely bound to them. In the case of the compound here, the anhydrate $CuSO_4$ *is white, while the hydrate* $CuSO_4 \bullet 5\ H_2O$ *is light blue. The name of this hydrate is called copper (II) sulfate pentahydrate because of the five water molecules attached to each formula unit of copper (II) sulfate.*

http://commons.wikimedia.org/wiki/File:Hydrating-copper(II)-sulfate.jpg