

Empirical and molecular formulas

If you're reading this tutorial, it's probably because you have to do some kind of calculation using empirical formulas. Something where your teacher said something like "If you have [whatever] percent of [some element] and [a different] percent of [a different element], what is the empirical formula of that compound?" It's just as much fun as it sounds, which isn't saying much.

Well, if you're having problems with this kind of problem, quit worrying. You're about to learn some amazing things.

What are empirical and molecular formulas?

A molecular formula is a chemical formula that tells you how many of each atom are in a chemical compound. For example, if you see the chemical compound H_2SO_4 , you can be sure that one molecule of this compound contains two atoms of hydrogen, one atom of sulfur, and four atoms of oxygen. In other words, this is the kind of formula you've been using your whole life.



Even babies love molecular formulas!

An empirical formula, on the other hand, is a type of formula that tells you the ratios of the elements in a chemical compound. Or, to put it another way, it reduces down the numbers after each element. For example, hydrogen peroxide has a molecular formula of H_2O_2 (i.e. two atoms each of hydrogen and oxygen), which means it has an empirical formula of HO (both of the "2"'s reduce to a 1). Likewise, the empirical formula of C_2H_6 is CH_3 , and the empirical formula of H_2SO_4 is, well, H_2SO_4 .

When should I use each type of formula?

Molecular formulas are the type of formula that you write about 90% of the time when you're doing chemistry stuff. The other 10% is made up of...

Structural formulas.

Wait, shouldn't empirical formulas be in there somewhere? You know, given the fact that this tutorial has the words "empirical formulas" in it?

No. It shouldn't. Empirical formulas haven't been used in actual chemistry for the better part of a century. When you're forced to do calculations, you're basically being forced to do calculations that no chemist has done in a real lab for a very long time. It's as if you were learning terms like "the bee's knees" and "23-skiddoo" in your English class – it's not really teaching you something *wrong*, it's just *irrelevant*.

So, why are we learning about empirical formulas, anyway?

I've always been honest with you, so it's time that you learned the reality of introductory chemistry: *The people who are in charge of determining what you learn are morons.*

This is not to say that your teacher is a moron. I've never met him or her, but I'd guess that he or she is probably reasonably intelligent, attractive, and charming. This is a characteristic that all chemistry teachers have in common.



Ladies...

No, the truth is that the people in charge of determining chemistry curricula are pencil pushers at the state board of education who have never been told what actual chemists do. As a result, they still believe that the state of experimental chemistry is somewhere back in 1874, and think that you should be taught accordingly. It was only last year that we teachers were able to convince those guys that atoms weren't "tiny demons", which we thought was quite the improvement.

Anyway, briefly put, if you want things to get better, tell your parents to quit voting for politicians who believe the sun goes around the earth.

In which I actually help you figure out what's going on

OK... I've gotten that out of my system, and are ready to start with the instruction. Let's have a look at the sort of problem that you're probably facing:

Example: I've got a compound that has a composition of 80.0% carbon and 20.0% hydrogen. What is the empirical formula of this compound?

Solution: Follow these steps to solve this problem:

- **Step 1: Assume you have 100.0 grams of this compound.** If you do this, you can conveniently say that you have 80.0 grams of carbon and 20.0 grams of hydrogen in this compound.
- **Step 2: Figure out how many moles of each element you have.** This requires that you do fancy mole calculations, which will look like this:

80.0 grams C	1 mol C	$= 6.66 \text{ mol C}$
12.01 grams C		

20.0 grams H	1 mol H	$= 19.8 \text{ mol H}$
1.01 grams H		

- **Step 3: Figure out the ratio of these two elements to each other.** Or, to put it another way, just look at the numbers of moles you found in step 2, and divide them both by the smaller one. In our example:

$6.66 \text{ mol C} / 6.66 \text{ mol C} = 1.00$
$19.8 \text{ mol H} / 6.66 \text{ mol C} = 2.97$

(Note: Because of real-world experimental errors, we usually assume that small mistakes have caused these numbers to be a little off. For example, though the number for H is shown as 2.97 above, it's probably a pretty safe bet to say that it's close enough to 3 as to be the real answer. However, if the answer had turned out to be 2.7 or 3.2, this is far enough off to suggest that there's a math error hiding around here somewhere.. Use common sense).

- **Step 4: Write the empirical formula.** So, given that the calculations above show that the ratio for carbon is 1 and the ratio for carbon is 3, this means that our empirical formula for this compound is CH_3 .

That's all there is to it!

Finding molecular formulas from empirical formulas

Let's say that you were given the same problem above, and given a second calculation to perform. For example, your teacher says that:

Part 2: If the actual molar mass of this compound has been determined to be 30.14 g/mol, what is the molecular formula of this compound?

Answer: This one's not too bad. You just need to...

- **Step 1: Find the molar mass of the empirical formula** (i.e. treat the empirical formula as if it were a valid compound). In the case of CH_3 , the molar mass would be 15.04 g/mol (12.01 for carbon, 3.03 for the three hydrogen atoms).
- **Step 2: Divide the molar mass that's given to you in the problem by the molar mass that you found for the empirical formula.** Since the problem gave you a molar mass of

30.14 g/mol and the mass you found for the empirical formula was 15.04 g/mol, you'll get a ratio of 1.997.

- **Step 3: Multiply the subscripts in the empirical formula by the ratio you just computed.** If your empirical formula is CH_3 , just multiply the number of atoms by 2 to give you a molecular formula of C_2H_6 .
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A note

If your teacher is clever, you'll be given masses that are just a little off from what they are in the real world. For example, in the problem above I told you that the molar mass of the compound was 30.14 g/mol. It turns out to actually be 30.07 g/mol. However, experimental error is something that you'll have to deal with in the real world, so it's not unreasonable to throw things off a little bit. As long as you follow the steps above, you shouldn't really notice things like this anyway.

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