

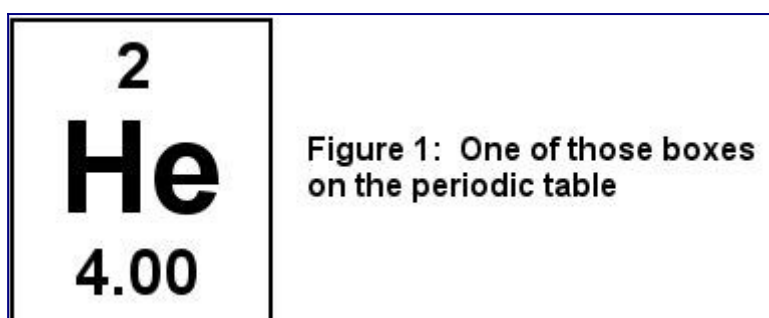
Isotopes and atomic mass

In the last section, we spent a lot of time talking about what an electron is. I mean, seriously, it was a lot of talking. When I finished with that section, I looked back and was, like, “Holy crap, that’s a lot of stuff about the electron!”

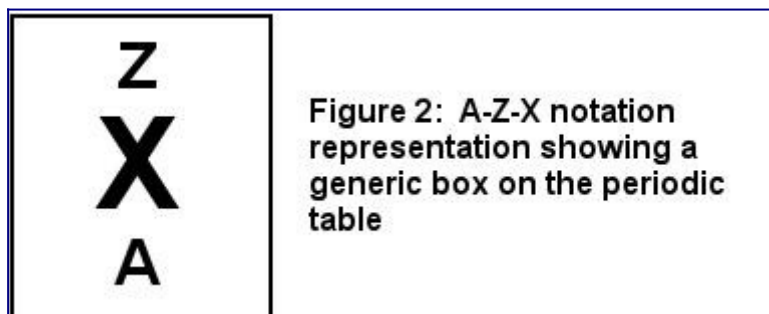
Anyhow, now that we’ve defined the proton, neutron, and especially the electron, it’s time to figure out what to do with them. As always, if you see a little superscript, look for an accompanying footnote at the bottom of the page.

The basics of using subatomic particles

When you look at the periodic table, you see a bunch of stuff that looks like this:



If we wanted to look not just at helium but at a generic box that we can use to represent all of the elements, we might set it up to look like this:



When we chemist types look at this, we interpret it in the following way:

- **The little “Z” is the atomic number.** The atomic number tells you how many protons each atom of this element has. This number defines the element, so if Z is equal to “2” (as it is in figure 1), you can be sure that the element is helium and has two protons in the nucleus.¹
- **The “X” is the atomic symbol.** This tells us the symbol of the element that’s represented. Though we don’t really need this symbol if we have the atomic number (helium will *always* have an atomic number of 2, for example), it makes our lives easier to just write it anyway.
- **The “A” is the average atomic mass of the element.**² If you were to pick a bunch of atoms of this element out of a bag, the average of the masses of each particle will be this number.³ In Figure 1, we see that A is equal to 4.00, which means that helium has an average atomic mass of 4.00 amu/atom or 4.00 g/mol.

Isotopes

Let's say that you've got an atom of helium in your pocket (for some reason). When you put it on a very tiny scale, you find that the mass of the atom is 3.00 amu.⁴ Awesome.

Let's say that a friend of yours gives you another atom of helium. When you put that on your tiny scale, you find that it weighs 4.00 amu? What's going on?

You've just learned about isotopes. As you know, neutrons are present in an atomic nucleus to keep the protons from repelling too much. Because there may be more than one possible number of neutrons that will do this, there will be more than one possible form of that element. These different forms are the isotopes.

The isotopes of an element have different masses because the mass of an atom is equal to the number of protons + the number of neutrons. All atoms of an element have the same number of protons, but if there are different numbers of neutrons, the mass of each form (i.e. isotope) will be different. In our example above, the helium atom with a mass of 3.00 amu had two protons and one neutron, while the atom with a mass of 4.00 has two protons and two neutrons.

And what about electrons? Neutral atoms have equal numbers of protons and electrons to keep the charges neutralized. However, ions are an entirely different story, and will be something to read about later!



Ask Dr. Depp!

Dear Dr. Depp,

I don't understand what all that stuff in the last paragraph meant. What happened with the protons and neutrons?

***Thanks,
Totally not imaginary person***

Dear person who actually exists,

Here's what it means:

If an atom of helium weighs 3.00 amu, this means that the sum of the protons and neutrons is equal to 3. Because helium always has two protons (the atomic number is 2), that means that this isotope must have 1 neutrons, because 2 protons + 1 neutrons = 3.

If the other atom has a mass of 4.00, then it must have two neutrons (2 protons + 2 neutrons = 4).

**Don't sniff paint,
Dr. Depp**

The notation we use when describing an isotope may take several forms:

Figure 3: Isotope notation

${}^3_2\text{He}$ The number at the top left indicates the isotopic mass (in this case, 3 amu), while the number at the bottom left is the atomic number (in this case, 2). "He," of course, means it's a helium atom.

${}^3\text{He}$ This is the same as the one above it, except that we left off the "2". After all, we didn't really need it because we know that all atoms of helium have two protons, no matter what!

helium-3 This notation just says in very plain terms, "Hey, this is helium with a mass of 3 amu."

No matter what you use, remember that the term “isotope” doesn’t mean that something is radioactive, poisonous, or ready to explode. *Every atom in the universe is an isotope of some element.* As mentioned before, the term “isotope” just signifies what flavor of the element we have, not what properties that flavor have.

**Figure 5:
Delicious, delicious
isotopes.**



What’s the deal with elements that have atomic masses that aren’t whole numbers on the periodic table?

If you refer earlier on this page, you’ll see that this number refers to the “average atomic mass”, which I said would be the average mass of the atoms that you’d find if you pulled a bunch out of the bag at random. For example, if you had an element for which 50% of the atoms weighed 12 amu and 50% weighed 14 amu, you’d correctly assume that the average atomic mass would be 13 amu.

Unfortunately for us, there are a couple of reasons this isn’t true:

- There are not equal amounts of each isotope for each element. It would be easy to find the average weight of students in your classroom if 15 weighed 50 kg and 15 weighed 70 kg – just average 60 and 80 to get an answer of 70 kg. Unfortunately, even if the only possible weights for students were 60 and 80 kg, you’d probably find that 8 out of 30 people weighed 60 kg, making the problem harder.⁵
- The isotopic masses aren’t whole numbers. We’d like to pretend that an atom of copper-63 weighs 63 amu, but the actual number (for reasons we’re not going to discuss) is actually 62.930 amu.
- Demons. Though I haven’t found it to be successful in either my personal or professional life, I maintain that many of my screw-ups are due to demons.

Fortunately, we can find average atomic masses with no problem at all. All we have to do is take a weighted average of the masses of each isotope. Which is a fancy way of saying that we’re going to account for the fact that there are uneven abundances of each isotope.

Finding average atomic mass

To find the average atomic mass of an element, use the following equation:

$$\text{average atomic mass} = \left(\text{mass of isotope 1} \right) \left(\text{abundance of isotope 1} \right) + \left(\text{mass of isotope 2} \right) \left(\text{abundance of isotope 2} \right) + [\text{and so on}]$$

Let's use this in a sample problem:

A sample problem: Copper has two stable isotopes: copper-63 has a mass of 62.930 amu and copper-65 has a mass of 64.928 amu. If a sample of copper contains 69.17% copper-63 and 30.83% copper-65, what is the average atomic mass of copper?

A sample answer:

Let's plug the values from this equation into the equation above:

- average atomic mass is what we're trying to solve for. Let's call it x.
- mass of isotope 1 (copper-63) is 62.930 amu.
- abundance of isotope 1 (copper-63) is 0.6917 (remember, we have to convert percent into decimal form by dividing by 100.
- mass of isotope 2 (copper-65) is 64.928 amu.
- abundance of isotope 2 (copper-65) is 0.3083.
- [and so forth] doesn't exist because there are only two isotopes. If there were more, than you'd have to keep doing the same thing for isotope 3, 4, and so on.

Which, gives us:

$$\begin{aligned} X &= (62.930 \text{ amu})(0.6917) + (64.928 \text{ amu})(0.3083) \\ &= 43.52 \text{ amu} + 20.02 \text{ amu} \\ &= 63.54 \text{ amu} \end{aligned}$$

Which is pretty much what you find on the periodic table. (Rounding differences account for the 0.01 amu difference).

How do they figure out isotopic masses and abundances?

They do this using mass spectrometers, which are interesting little machines that you can read about on [this page](#). Actually, they're not that little. But they are interesting. And very, very expensive.



Footnotes

1. The atomic number / number of protons is the same for all atoms of an element. The mass, however, may be different in different atoms, which leads to the idea of isotopes. Don't worry... we'll get to it.
2. The average atomic mass of an element is usually just referred to as its atomic mass. Which isn't technically true for reasons we don't really care about, but is always done anyway. It's also not wrong to call the average atomic mass the "molar mass", because one mole of the element will weigh that many grams. More about moles later.
3. Or not. There's actually a lot of controversy about this number because, depending on where the sample of interest came from, the average mass of the particles in the sample may be different. For example, differences in the isotopic makeup in a sample are common in geochemistry and can also be used in carbon dating and in identifying nuclear material. For a table illustrating this point, check [this](#) out.
4. For you purists out there, I'm well aware that the mass of a helium-3 atom will actually be 3.02 amu. I'm making a point, so just go with it.
5. Same as #3 – I just wanted to remind you.
6. It's hard to say how much a mouse lemur costs because they are so endangered that they aren't openly sold. However, the much more common ring-tailed lemur is sold as a pet in the United States at prices between \$1,000-\$4,000. I'm not going to include a link, though, because you probably shouldn't buy one.
7. Only 100 donkeys in the world produce the milk needed for this cheese, and only in May and June. Since it takes 25 liters of this milk to make 1 kilogram of cheese, it's practically a bargain!
8. Because krytrons can be used to trigger nuclear detonations, the United States government is pretty serious about keeping them from being exported. The "40 month" number comes from a case in which 300 were sold to Israel in the early 1980's, resulting in a guilty plea and subsequent prison sentence for violating the US Arms Export Control Act. [Link](#).

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