

Lewis structures of ions

Now that you've mastered the art of drawing the Lewis structures of neutral covalent compounds, it's time to draw the Lewis structures of polyatomic ions. If you haven't mastered the art of drawing Lewis structures yet, you can check out the tutorial by clicking [HERE](#). Seriously, unless you check out that tutorial, this probably won't make a lot of sense.

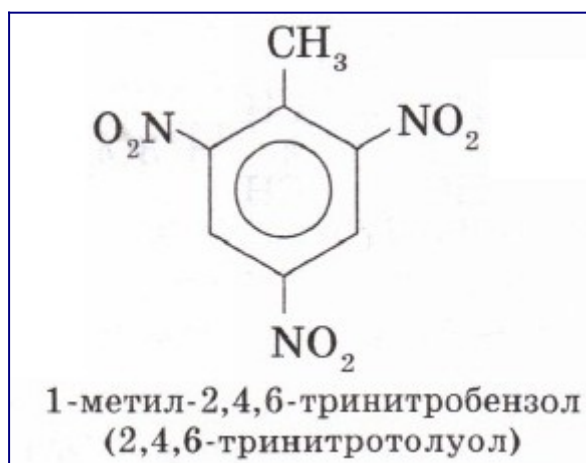
Let's get started!

What do polyatomic ions have to do with covalent compounds?

Well, it depends. In a very strict sense, polyatomic ions have nothing to do with covalent compounds. After all, covalent compounds don't contain any ions at all, which is why we need different sections for ionic and covalent compounds.

However, it turns out that the polyatomic ions that are present in ionic compounds have atoms which are covalently-bonded to each other. In other words, ionic compounds sometimes have covalent ions in them.

You'd think this wouldn't make any difference, and if you're looking at the basic properties of ionic compounds you'd be right¹. However, when you dissolve an ionic compound in water, those polyatomic ions are now free to wander around and react with whatever covalent compounds are present. As a result, the structures of these ions turn out to be important after all.²



The formation of TNT involves the nitrate ion. I've included a foreign version of this diagram for our Russian friends. России-замечательная страна!³

How to draw the Lewis structures of a polyatomic ion

Before you can do this, you really need to understand the method for drawing Lewis structures that I talked about in the [previous tutorial](#). Go ahead... I'll wait for you to come back.

To understand how to do this, let's use the example of the hydroxide ion: OH-

Step 1: Find the number of valence electrons.

This is done in the same way as for neutral ions, except that when you're done, you need to change the number of electrons to compensate for the charge on the ion. Here's how:

- If it's an anion, add the negative charge to the number of valence electrons.
- If it's a cation, subtract the negative charge from the number of valence electrons.

Let's find the valence electrons for OH⁻:

Oxygen: 1 atom x 6 valence electrons = 6 valence electrons

hydrogen: 1 atom x 1 valence electron = 1 valence electron

$$6 + 1 = 7 \text{ valence electrons}$$

The additional step: OH⁻ has a -1 charge, so we need to add one to the number we found, giving us a total of **8 valence electrons**.

Step 2: Find the number of octet electrons.

The rules for doing this are pretty much the same as we saw before, with a few small modifications:

- hydrogen always wants two electrons
- beryllium always wants four electrons
- boron wants six electrons in neutral compounds, but 8 electrons in polyatomic ions
- all other elements always want eight electrons

In the case of the OH⁻ ion:

oxygen: 1 atom x 8 octet electrons = 8 octet electrons

hydrogen: 1 atom x 2 octet electrons = 2 octet electrons

$$\text{Grand total: } 8 + 2 = \mathbf{10 \text{ octet electrons}}$$

Step 3: Find the number of bonding electrons.

Again, this is done the same way as it is for neutral compounds, where you need to subtract the valence electrons from the octet electrons. For us:

$$10 \text{ octet electrons} - 2 \text{ valence electrons} = \mathbf{2 \text{ bonding electrons}}$$

Step 4: Find the number of bonds.

Divide the number of bonding electrons by 2 to get the number of bonds:

$$2 \text{ bonding electrons} / 2 = \mathbf{1 \text{ bond}}$$

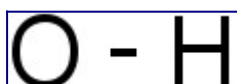
Step 5: Draw it!

When drawing polyatomic ions, we do pretty much the same thing as we did with neutral covalent compounds. However, the rules for doing this are a little different. I love making charts, so here's a chart that shows what I mean:

Family	# of bonds in neutral covalent compounds	# of bonds in polyatomic ions
carbon's family	Bonds four times	Bonds four times
B or nitrogen's family	Bonds three times	Bonds three or four times
oxygen's family and beryllium	Bonds twice	Bonds once or twice
halogens or hydrogen	Bonds once	Bonds once

This chart is basically the same thing you saw in step five in the first Lewis structure tutorial, except that it the right column gives a somewhat different set of rules when drawing the Lewis structures of polyatomic ions.⁴

For the example of OH⁻, we've found that there's one bond. Since there are two atoms, it's pretty simple to see that this bond lies between them:



Step 6: Add lone pairs

The basic process is the same as for neutral compounds, except that the number of electrons everything wants reflects the rules shown in #2 above. (In case you don't feel like scrolling up, hydrogen and the halogens always want 2 electrons, beryllium wants 4 electrons, boron wants 6 electrons in neutral compounds and 8 in polyatomic ions, and everybody else wants 8 electrons).

Looking at OH⁻, we see that oxygen has only two electrons around it (the pair of electrons it shares with H), and hydrogen also has those same two electrons around it. While hydrogen only wants two electrons, oxygen wants eight. As a result, we need to add six electrons in the form of lone pairs:



Awesome!

Step 7: Figure out which atom is charged

Recall that the whole reason we're doing a tutorial here is that some covalent things have either positive or negative charges. As a result, we've got to figure out which atom is positively or negatively charged.

Here's how you do it: Count the number of electrons that each atom actually owns in this Lewis structure. In our example, oxygen owns seven electrons. Why seven? Let's discuss:

The three lone pairs are on oxygen and have nothing to do with bonding. Thus, oxygen owns all six of these electrons outright.

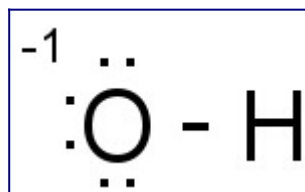
The bond it has with hydrogen is formed when one electron from oxygen and one electron from hydrogen combine to form a covalent bond. Though the bond has a total of two electrons, only one of those came from the oxygen atom. As a result, all bonds count as one electron in this step, rather than two.

Now that we know how many electrons oxygen owns, let's compare that to how many it usually has:

- Oxygen usually has six valence electrons, making it neutral.
- In the OH⁻ ion, oxygen actually owns seven electrons. This means that it has one more electron than it usually does in its neutral state.
- Because electrons are negatively-charged, oxygen has a -1 charge.

Hydrogen, on the other hand, doesn't have any charge at all. In the OH⁻ ion it owns one electron (one of the electrons from the bond with oxygen) and has no lone pairs. Because hydrogen usually has only one electron, it's neutral.

Now that we know oxygen has a -1 charge, let's finish our drawing:



That's all we have to do: Put a "-1" over oxygen. If you wanted to just put a "-" over oxygen, that's fine too, as we assume any charge without a number must be one. In the case of something with a charge of "-2", just write "-2".

Addressing a common student complaint:

When they first learn how to draw Lewis structures, most students complain that Lewis structures suck. And, in all fairness, this method *is* a real pain in the butt and totally sucks. However, it also happens to work, and when you're just learning chemistry, that's really the most important thing to think about.

Now, do *I* draw Lewis structures this way? Nope. You've got to remember, though, that I've been doing this for 25 years, so I don't *need* to use this method anymore. As you do more examples, you'll start to get the hang of things and won't have to use this method anymore, either. Until then, just keep with it.

Another example: NH_4^+

Step 1: Count the valence electrons

Nitrogen: 5 valence electrons x 1 atom = 5 valence electrons

Hydrogen: 1 valence electron x 4 atoms = 4 valence electrons

Subtract 1 valence electron because it's a cation with +1 charge

Total: $5 + 4 - 1 = 8$ valence electrons

Step 2. Count octet electrons

Nitrogen: 8 octet electrons x 1 atom = 8 octet electrons

Hydrogen: 2 octet electrons x 4 atoms = 8 octet electrons

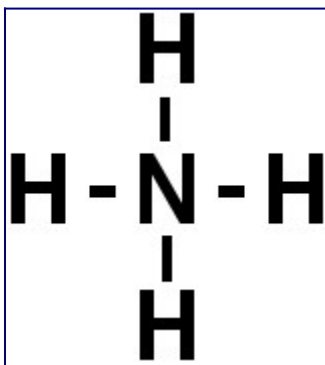
Total: $8 + 8 = 16$ octet electrons

Step 3. Find bonding electrons: $16 - 8 = 8$ bonding electrons

Step 4. Find number of bonds: $8 / 2 = 4$ bonds

Step 5. Draw it:

Since nitrogen wants three or four bonds in polyatomic ions, we can see that this is A-OK:

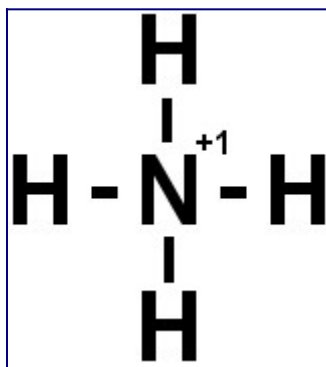


Step 6: Add lone pairs

In this case, nitrogen already has eight electrons around it and hydrogen has its desired two electrons. As a result, no more electrons are needed.

Step 7: Assign charge

Hydrogen never has charge on it, which would correctly imply that nitrogen must have the +1 charge in this ion. However, if you didn't think of that, you can see that nitrogen owns four electrons in this ion (one from each bond) and that it usually has five valence electrons to keep neutral. As a result, it has one fewer electron than it usually has, giving it a +1 charge:



I'm not going to include any examples here, as the next tutorial will discuss both polyatomic ions and resonance structures (which are closely related in many cases). Check over there for examples.

Footnotes:

1. Mostly right, anyway. It's pretty hard to say that the locations of charges and/or resonance structures are totally irrelevant to the properties of ionic compounds, but it's certainly the case that they're less important than the part where the ions interact.
 2. To put it another way, you can think of polyatomic ions in solution as acting a lot like any other covalent compound. Organic chemists use this a lot when building new compounds.
 3. As measured by the burliness of Vladimir Putin, anyway.
 4. Please keep in mind that these are general rules and don't represent infallible secrets to the universe. If you use these rules, however, you should be in pretty good shape for anything that your teacher throws at you.
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Figure credits:

1. Russian TNT diagram: By Bauka91 91 (Own work) [CC BY-SA 3.0 (<http://creativecommons.org/licenses/by-sa/3.0>)]%5D, via Wikimedia Commons.