CHE 1301

Basic Principles of Modern Chemistry I

Week 6

Hi! Thanks for checking out the weekly resources for Chemistry 1301! This resource covers topics typically taught by professors during the 6th week of classes.

Visit our website, https://baylor.edu/tutoring, to sign up for appointments and check out additional resources for your course! You'll find helpful links with the following titles:

- "Online Study Guide Resources" – The pace of your course may vary slightly from what's shown in this document. If you don't see the topics you're learning right now, use this link to find the weekly resources for the rest of the semester.
- "How to Participate in Group Tutoring" - See if there is a Chemistry 1301 group tutoring session being hosted this semester. These are weekly question/answer sessions taught by our master tutors!
- "View tutoring times for your course" or "Schedule a private 30-minute appointment!"

You can also give us a call at (254)710-4135, or drop in! Our hours are Monday-Thursday 9 am – 8 pm on class days.

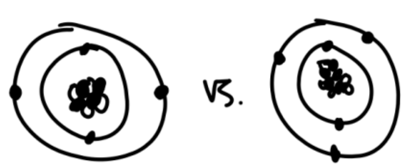
KEY WORDS: Electronegativity, Formal Charges, Resonance, Orbital Diagrams, Orbital Energy Levels

TOPIC OF THE WEEK: Bond Formation (Electronegativity cont.)

Electronegativity: Which atom is best at keeping electrons, and how does this decide bond type?

Periodic trend: Electronegativity generally increases across a period and decreases across a group. In other words, the most electronegative elements are furthest right and furthest up. The most electronegative element is fluorine.

So why are the trends the way they are? Well, if we're comparing two atoms, and one has one more proton than the other, which will be better at keeping its electrons close?



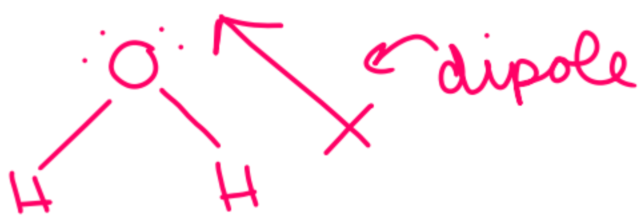
Usually, it's the one with more protons, because protons are what "hold on" to electrons. This is why electronegativity increases across a period. But what if that one electron is on the next energy level?

In this case, the atom on the left will be better at holding on to its electrons, because the electrons are nearer to the nucleus and easier to “keep safe.”

Why does it matter which atoms are best at holding on to electrons? This is what decides what type of bond will be formed.

If one atom is a lot “stronger” than the other, it will just take the latter atom’s electron! This is called an **ionic bond**. After an electron has been stolen, it is the force between the opposite charges that keeps the atoms together.

If one atom is a little stronger than the other, the two will share the electron, but it will spend more time with the stronger atom. This is called a **polar covalent bond**. This unevenness is called a “**dipole**.” Since electrons are negatively charged, this results in a **partial negative charge** on the more electronegative atom. “Partial” means that the atom isn’t as negatively charged as it would be if it were to receive a whole other electron—it doesn’t have quite a -1 charge—but it is more negatively charged than it would normally be.



If their strengths are about the same, they will share the electron equally. This is called a **nonpolar covalent bond**. The bond between hydrogen and carbon is an important example to know.

A couple related terms to know:

Oxidation: Atom loses an electron. Named this way because forming bonds to oxygen often cause electrons to migrate away from the atom…oxygen is very electronegative!

Reduction: Atom gains an electron

For more info on electronegativity and related periodic trends, check out this video! [The Periodic Table: Atomic Radius, Ionization Energy, and Electronegativity - YouTube]

Highlight 1: Formal Charges

Formal charge can be defined as *effectively* how much the number of protons in an atom’s nucleus is balanced out by the number of electrons around the nucleus. In mathematical terms…

(number of electrons that an atom **normally** has: see periodic table) minus (number of electrons that are spending time around in atom **right now**, when it is bonded to some other atoms)

Or, (atomic number) – (number of electrons in lone pairs around the atom) – (half the number of bonds around the atom…we'll assume that, since two electrons are being shared between two atoms, it's basically like one electron "belongs" to each atom)

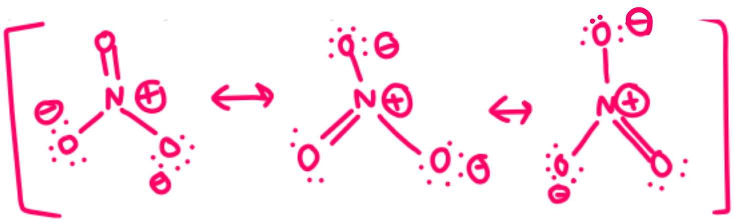
Here's an easy way to remember the formula: **Formal charge = atomic # -- (dots + sticks)**

How To Calculate The Formal Charge of an Atom - Chemistry - YouTube

Resonance: After single bonds have formed, which atoms will have the remaining electrons?

Resonance, like electronegativity, results in partial positive and negative charges on atoms.

Sometimes, multiple Lewis structures are possible for an atom. If this is the case, the electrons will balance themselves out between these possible structures. Here's an example (see the Check Yourself section and Answers at the end of this doc to see where the negative and positive charges are coming from!)



You can see that the three structures above are all valid Lewis structures for NO3. (See last week's resource for more on Lewis structures!)

All of the bonds are somewhat like double bonds, because if just one of the bonds were double, its electrons would be pushed away from the center of the atom, towards the oxygen atom, by the electrons from the nearby overloaded oxygen atoms. And then, if the new oxygen atom were to become overloaded with electrons, it would push them back toward the center…so the actual location of the electrons ends up being something in between all of the different options. Don't worry about this too much yet—that last point is definitely the one to remember!

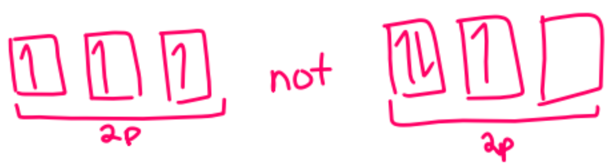
Highlight 2: Orbital Diagrams: Box Notation

Orbital diagram: One way to write electron configuration—can be compared to the notation "$1s^2 2s^2 2p^2$" that you'll recognize! Basically, each orbital is represented by a rectangle. The number of boxes grouped together depends on how many orbitals are possible for a given energy level, or how many ml values are possible for a given n and l (for example, there are 3 possible 2p orbitals). An orbital can hold two electrons; two arrows can go in each rectangle. One of them faces up, and one faces down—this represents the fact that one of them has ms=+0.5 and the other has ms=-0.5. Here is the orbital diagram for carbon:

The order in which they are filled depends on the following 3 rules:

1. **Aufbau principle**: The order in which electron orbitals are filled depends on how much energy it takes to sustain electrons in that orbital! An atom will fill lowest energy orbitals first. See Highlight 3!
2. **Pauli exclusion principle**: It is impossible to have two electrons with the same 4 quantum numbers--each electron must have its own space. See example below:

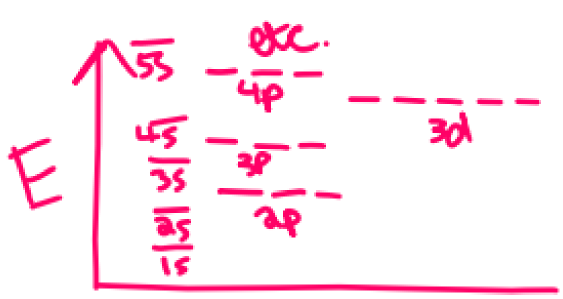
3. Hund's rule: When filling orbitals that are sustained by the same amount of energy (when you have more than one option according to the Aufbau principle), put one electron in each orbital before you start completing the pairs.



https://www.youtube.com/watch?v=C6afrc1QS6Y

Highlight 3: Orbital Diagrams: Energy Diagrams

We've discussed that electrons fill orbitals based on which has the least energy to sustain (see Electron Configuration from September 5's resource). The amount of energy that it takes to sustain an orbital can be represented by an energy diagram! **The y-axis represents the amount of energy**—the further up you get, the more energy it takes to put electrons in one of those orbitals. A line represents an orbital that can be filled by electrons. Here's what they look like:

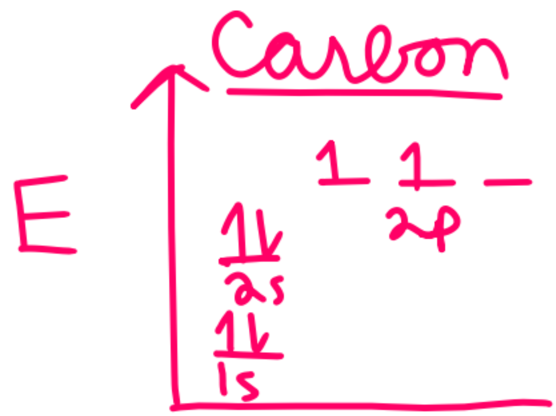


You'll notice that, as you read the orbitals listed on an energy diagram from bottom to top (sometimes you might have to skip around left to right), the order in which they are filled is the same as the order in which you've been writing out their electron configurations!

To organize things a bit, s orbitals are written on the left side of the chart. The p orbitals are in a column next to them, then d, then f.

Steps for drawing an energy diagram for a specific element:

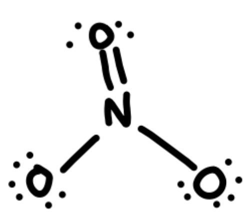
1. Draw the orbitals up to the highest row that will have electrons from that atom. For example, if you have an atom with 4 2p electrons, the lines that you draw will be 1s, 2s, and 2p.
   a. If there are multiple orbitals available for a certain row, draw multiple lines. For example, since 2p has 3 available orbitals, draw 3 lines with the same energy (same height).
2. Draw arrows to represent each electron. Remember the Aufbau principle and Hund's rule! (see Highlight 2)



https://www.youtube.com/watch?v=bgAuj4EEmw4

Check Your Understanding

1. Order the following in order of increasing electronegativity difference: polar covalent bond, nonpolar covalent bond, ionic bond. *Hint: The greater the electronegativity difference, the "stronger" one atom is than the other.*
2. Calculate the formal charge on each atom for the following compound:



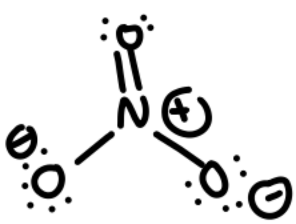
3. Draw the orbital diagram for oxygen.
4. Draw the energy diagram for oxygen.

Things You May Struggle With

- *Aren't some resonance structures the same thing? Wouldn't they be the exact same drawing if you just rotated them?* It's a weird concept for sure! The reason that the resonance structures shown above, for example, are not identical, is because each of the oxygen atoms has to "take a turn" having the double bond.
    - Don't worry about this too much yet—just know that it's possible for different Lewis structures to exist for a molecule, and that if they do, the "true" Lewis structure will be a *hybrid* between them.
- *What is the difference between the Aufbau principle, Hund's rule, and the Pauli exclusion principle?* The Aufbau principle describes the order in which orbitals become available—it orders them in terms of increasing energy. Hund's rule describes the order in which electrons fill orbitals at each energy level. The Pauli exclusion principle is the reason that orbitals become "full"—they don't have unlimited space, because each electron needs its own space.

That's all this week! Please reach out if you have any questions and don't forget to visit the Tutoring Center website for further information at www.baylor.edu/tutoring. Answers to Check Your Learning are below.

1. Nonpolar covalent bond, polar covalent bond, ionic bond
2. N +1, O -1, O -1, as shown below:



$$F.C._N = 5-(0+4)=1$$
$$F.C._O = 6-(4+2)=0$$
$$F.C._O = 6-(6+1)=-1$$
$$F.C._O = 6-(6+1)=-1$$

3.