CHE 1301

Basic Principles of Modern Chemistry I

Week 5

Hi! Thanks for checking out the weekly resources for Chemistry 1301! This resource covers topics typically taught by professors during the 5th 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: Lewis Structure, Ionic Compounds, Covalent Compounds, Molecular Geometry

TOPIC OF THE WEEK: Lewis Structures

What is a Lewis structure?

It's scientists' go-to representation of a covalently bonded molecule. It shows the electrons present and where they are located (in bonds, as lone pairs, etc.)

How to draw one:

ex: H_2O

1. Determine the number of valence electrons for each atom.
    a. Valence electrons are electrons in an atom's outermost energy level. Because they are furthest from the nucleus, they are most accessible, which means that they are the electrons that will be exchanged/shared in chemical reactions. This is why chemists have

H: 1
H: 1
+ O: 6
8

a specific type of diagram called a Lewis structure that shows a molecule's valence electrons.

   b. *# of valence electrons = column (group) number on the periodic table*

2. Determine the total number of valence electrons.
   a. Add up values from step 1.
   b. *If drawing Lewis structure of an ion:*
      i. *Add 1 for each negative charge (an anion has gained negatively charged electrons)*
      ii. *Subtract 1 for each positive charge (a cation has lost negatively charged electrons)*
3. Put the atom with the greatest bonding capacity in the center.
   a. *Bonding capacity = number of electrons it would take to form an octet (or in the case of hydrogen, a duet)*
   b. **Why do atoms want to form octets?** This goes back to the number of electrons that can fit on each energy level. Atoms "want" to have full energy shells. When n=1, there are 2 spots for electrons—that's why hydrogen "wants" to have 2. When n=2, there are 8. Most of the atoms that you'll see will "want" 8 electrons, which is called the **octet rule**. As you know, in higher energy levels, atoms have access to more orbitals (see last week's resource on quantum numbers), so there are some exceptions to this rule! Check out this link for more info: [9.9: Exceptions to the Octet Rule - Chemistry LibreTexts]

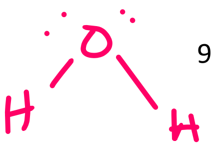
O needs 2 bonds
H needs 1 bond

H–O–H
e^-s: 2+2=4
✓ duet ✓ duet

4. Put the other atoms around it and draw covalent bonds (represented by lines) in between. Keep track of the number of electrons that you've used—each bond takes 2 electrons, because each atom is sharing one of its electrons with the other.
5. Give all **terminal atoms** (all the atoms except the one in the center) enough electrons to have an octet (or duet, if your terminal atom is hydrogen).
6. Put all remaining electrons around the central atom.
   a. # of remaining electrons = total electrons (step 2) – electrons used (step 4)

8-4=4
H–Ö–H
octet ✓

7. Use double/triple bonds to give the central atom an octet, if it does not have one. Here's an example of a Lewis structure that requires use of this step: [Lewis diagram of the cyanide ion (worked example) (video) | Khan Academy]
8. *probably won't have to worry about this too much* If there are multiple possible Lewis structures, choose the one with the best formal charge.
   a. Formal charge = number of electrons an atom has in its natural state – number of electrons it has as lone pairs – half the number of electrons it has as bonds. See this video: [Formal charge (video) | Khan Academy]



b. Choose the structure with (1) the least number of formal charges and (2) the formal charges that are furthest apart.

9. You may be asked to adjust your drawing to include accurate representation of electron geometry. See Highlight 3!

Highlight 1: Ionic Compounds

Ionic compounds' goal is to have a total neutral charge! Keep this in mind when figuring out how many of each ion will be part of one compound.

Examples of cations	Examples of anions
Alkali metals and **alkaline earth metals** (first 2 groups of the periodic table) like to lose electrons, because this is the easiest way for them to *obtain full energy shells.* Charge = + group number	**Halogens** and other nonmetals like to gain electrons, because this is the easiest way for them to obtain *full energy shells.* Charge = group number minus 8
Polyatomics: ammonium (NH_4^+), hydronium (H_3O^+), mercury (Hg_2^{2+})	Polyatomics: See this video for a helpful mnemonic! [Nick the Camel - YouTube] Names of oxyanions: However many oxygens ____ate has, ____ite will have one less. Hypo___ite will have one less than that. And Per___ate will have one more than ____ate: Hypo___ite ___ite ___ate Per___ate
Transition metals Charge = depends; these metals have the possibility of forming ions of different charges. The name of an ionic compound with a transition metal will have the charge of the transition metal listed as a Roman numeral in parentheses after its name: copper (II) chloride.	

Writing **formula** given **name**:

1. Identify charge of each ion.
2. Multiply as necessary to balance.

Ex. Nickel (II) hydrogen sulfate → Ni(HSO4)2

Writing **name** given **formula**:

[name of cation] SPACE [name of anion, minus its last syllable]-ide

Ex. $CaCl_2$ → calcium chloride

Highlight 2: Covalent Compounds

Writing **name** given **formula**:

[prefix][name of element with lowest group #] SPACE [prefix][name of other element minus last syllable]-ide

When naming a covalent compound, you must use prefixes that tell the reader how many of each element are in the compound. In ionic compounds, this wasn't necessary, because it could be calculated using the charges.

Prefixes:

1-mono

2-di

3-tri

4-tetra

5-penta

Etc.

Naming acids (a special type of covalent compound):

Binary acids:

Hydro[first syllable of name of other element]-ic SPACE acid

Oxoacids:

[oxoacid name minus -ate or -ite][-ic if the oxoacid was an ___ate, or -ous if the oxoacid was an ___ite] SPACE acid

Formula of an acid:

1. Write the formula of the anion.
2. *Number of hydrogens = charge of anion*. Write the hydrogens first, and the anion second.

Always put $_{(aq)}$ at the end!

Check out this quick video for an overview! [How To Name Acids - The Fast & Easy Way! - YouTube]

Highlight 3: Molecular Geometry

Here is a link explaining molecular geometry, which is the spatial arrangement of electrons around atoms: [Electron Geometry VS Molecular Geometry - Difference between Electron and Molecular geometry, Determination of electron Geometry and Molecular Geometry along with FAQs (byjus.com)]

(Key point: The atom will try to keep electrons as far away from each other as possible. This is why it arranges them the way it does!)

First, count the number of electron groups around the atom in question.

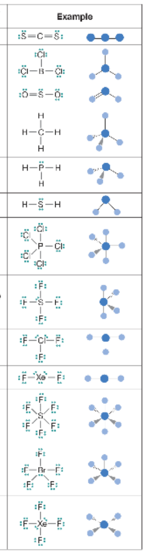
Now, there are two things you can describe: electron geometry and molecular geometry.

Electron geometry describes the locations of all the pairs of electrons surrounding an atom. This includes lone pairs and bonds.

Molecular geometry describes the locations of pairs of electrons that are connected to other atoms (in other words, the bonds).

Much more in the Week 7 resource! For now, here's a table summarizing molecular and electron geometries to know. Check out Week 7 for tips on memorizing it!

Electron Groups	Bonding Groups	Lone Pairs	Electron Geometry	Molecular Geometry	Approximate Bond Angles	Example
2	2	0	Linear	Linear	180°	
3	3	0	Trigonal planar	Trigonal planar	120°	
3	2	1	Trigonal planar	Bent	<120°	
4	4	0	Tetrahedral	Tetrahedral	109.5°	
4	3	1	Tetrahedral	Trigonal pyramidal	<109.5°	
4	2	2	Tetrahedral	Bent	<<109.5°	
5	5	0	Trigonal bipyramidal	Trigonal bipyramidal	120° (equatorial) 90° (axial)	
5	4	1	Trigonal bipyramidal	Seesaw	<120° (equatorial) <90° (axial)	
5	3	2	Trigonal bipyramidal	T-shaped	<90°	
5	2	3	Trigonal bipyramidal	Linear	180°	
6	6	0	Octahedral	Octahedral	90°	
6	5	1	Octahedral	Square pyramidal	<90°	
6	4	2	Octahedral	Square planar	90°	



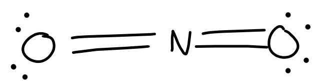
Check Your Understanding

1. Draw the Lewis structure for NO2.
2. Name CaCl2.
3. Name CF4
4. What is the formula for hydrobromic acid?

Things You May Struggle With

1. *What's the difference between ionic and covalent compounds?* **Ionic compounds** contain elements attracted to each other because one has lost an electron and one has gained an electron. **Covalent compounds** share electrons.
2. How do I tell which is which? If the atoms are close to each other on the periodic table, they have about the same number of protons. This means that one is probably not strong enough to take an electron from the other. As a general rule, metal + nonmetal is an ionic compound. Nonmetal + nonmetal is a covalent compound.

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. 
2. Calcium chloride
3. Carbon tetrafluoride
4. $HBr_{(aq)}$