

## Chapters 11 & 12. Cracolice, 2/e. Chemistry X11.

These two chapters go together. We discuss two types of chemical bonds in terms of the electronic structure of atoms. You already know about ionic bonds, in which ions of opposite charge attract each other; the ions themselves resulted from transfer of electrons from one atom to another. Covalent bonds also involve electrons, but in this case the electrons are shared -- more or less equally -- between the bonded atoms. You will learn how the two types of bonding work, and how to predict which is likely in a given case. And you will learn about the “more or less.”

For covalent compounds (which includes organic compounds), we use the information about the electrons (both in and out of the bonds) to tell us about the shape of molecules. And we look at polarity, the property that explains why oil and water don't mix. The electronegativity (EN) table helps us identify bond polarity; **you should know** the general trend of this table. Distinguish bond polarity (Sect 11.4) and molecular polarity (Sect 12.5).

Lewis structures (electron dot formulas) are a convenient way to keep track of the electrons in chemical structures. Remember: show only the valence electrons in Lewis formulas (Sect 10.4). The octet rule guides us to proper bonding (and is based on electron configurations, Ch 10).

⇒ There is a practice quiz for Ch 11-12 at the web site.

### Priorities

Ch 11 Sect 1-4. Sect 5-6 build directly on these. We will probably ignore Sect 7-8.

Ch 12 Sect 1-3. Sect 4 builds on Sect 3, and is important. Sect 5 requires the idea of bond polarity from Ch 11 and the idea of shape from Ch 12 Sect 3. We won't explicitly cover Sect 6, but will probably choose some examples from it.

A key conceptual point is the distinction between electron geometry and molecular geometry.

**You should know:** the standard names for the basic shapes, as Cracolice presents them.

Use this statement of priorities as a guide, both in organizing and in doing problems. The sections listed are most important, and sometimes are needed for later sections.

The problems for these chapters are short. Do LOTS!

### Errata and Notes

p 295, bottom. Cracolice is correct that the  $\text{H}^+$  ion commonly exists in the hydrated form,  $\text{H}_3\text{O}^+$ . However, out of tradition and/or laziness, it is often written as  $\text{H}^+$ , and even referred to as a proton. That is,  $\text{H}^+$  and  $\text{H}_3\text{O}^+$  are two ways to represent the same chemical species. More

in Ch 18. (You should know that  $\text{H}_3\text{O}^+$  is “more correct”, but in general I don’t really care which way you write it.)

p 299. The term “molecule” is properly used only for covalent compounds. Ionic compounds exist as an “infinite” array in the solid state and as separated ions in solution. There is no NaCl molecule (except perhaps in the gas phase, which is not common for ionic compounds); NaCl is the “formula unit.” Again, this precision of language is not always observed, but you should recognize the conceptual distinction.

p 304, Sect 11.7. Again... Exceptions are less important than the rules. This is an interesting section, but not high priority.

p 320, Example 12.3. For counting the valence electrons, the set-up shown does not account for there being three O atoms.

p 336, Sect 12.6. A good quickie introduction to organic chemistry. You will probably find many familiar chemicals here. If you try to read this section, emphasize the patterns, not specific compounds. Try to learn the major types of organic compounds, such as alkanes and alcohols. (In X402 we will spend about a chapter on each of these. Join us!)

Answers...

- p 310 #8. The answer given is somewhat ambiguous. There is one bonding pair. Further, there are 6 lone pairs, 3 on each atom.
- p 342 #10. The shape of ICl has nothing to do with VSEPR.
- p 342 #21. Answer given is very weak. The problem is the use of the vague term “balanced”. You need to come up with a better explanation than this!

At several points, H is special. This special behavior follows from the fact that the first energy level holds only 2 electrons (Fig 10.8). Thus  $Z = 2$  (He) represents a full shell and a noble gas. Put in this context, H will usually seem quite reasonable. In particular, note that the EN of H is typical of elements in the middle of the PT, not those of groups 1A or 7A. [Recall Cronyn (2003), in Ch 5 handout.]

### **Further reading**

N C Craig, Correspondence with Sir Lawrence Bragg regarding evidence for the ionic bond. J Chem Educ 79:953, 8/02. A bit of history here... an attempt in 1968 to reconstruct some thinking from the 1910s, with the help of one of the original key players.

R Hoffmann, The story of O. Amer Sci 92:23, 1/04. Oxygen is a chemical of great complexity. Ordinary O is  $\text{O}_2$ , a molecule both essential and toxic to life -- and for which an ordinary Lewis structure fails to explain its properties.  $\text{O}_3$  is ozone, a key noxious ingredient of photochemical smog -- and protective agent of our upper atmosphere. And more, discussed

here in a delightful article by Nobelist Roald Hoffmann, a regular columnist in American Scientist.

T L Meek & L D Garner, Electronegativity and the bond triangle. J Chem Educ 82:325, 2/05. We discuss mainly ionic and covalent bonds -- and their intermediates -- and we discuss how electronegativity is used to predict the bond type. The chapter also notes metallic bonds (Section 11.8). To discuss all three types and their intermediate cases together it is helpful to introduce a "bond triangle", with each of the three pure bond types at one vertex. Even a browse of this article may give you a broader view of bond types.

R C Kerber, If it's resonance, what is resonating? J Chem Educ 83:223, 2/06. Sometimes we cannot find one ordinary Lewis structure that represents a molecule well. It may seem that the actual structure of the molecule is more like an average of two (or more) ordinary Lewis structures. For historical reasons, this phenomenon is called resonance -- even though nothing is really moving. This phenomenon usually just barely gets mentioned in an intro chem course; Cracolice (2/e) notes it on p 321. It will come up more if you continue with chemistry. Here is a nice article about this awkward term.

G Frenking & R Tonner, Theoretical chemistry: The six-bond bound. Nature 446:276, 3/15/07. News. How about bonds beyond "triple"? We normally say no, because within the context of the main group elements and a simple octet framework, it just doesn't happen. However, transition metal atoms have more orbitals available, and higher bonding order is possible. Recent work has provided evidence for quintuple bonds (between two Cr atoms) and possibly even sextuple bonds. This news story discusses all this, including some theoretical analyses, which suggest that six is the maximum -- at least for a homoatomic bond (two of the same atom).

### **Computer resources** (See web page for details and links.)

Movie: Formation of ionic bonds. This was also listed in the Ch 6 handout. One of the sequences shows the "electron clouds" of the atoms, illustrating an idea introduced in Ch 10. Other movies, on other Ch 11 topics, are also available from this page. These include movies on electronegativity and on atom vs ion size. I think the bonding movie listed here is the most useful as a movie, but look around, and tell me if you find others useful.

Ionic liquids -- a class of chemicals that violates one of our simple rules of thumb. The web page describes their nature and uses.

A site at Georgia Southern University has some good tutorials, with good visuals, on several topics from these chapters. The topics include: Lewis structures, VSEPR, bond and molecule polarity, intermolecular forces.

See my web page on Models, for an introduction to molecular models, which can be useful for helping you visualize the shape issues in Ch 12.