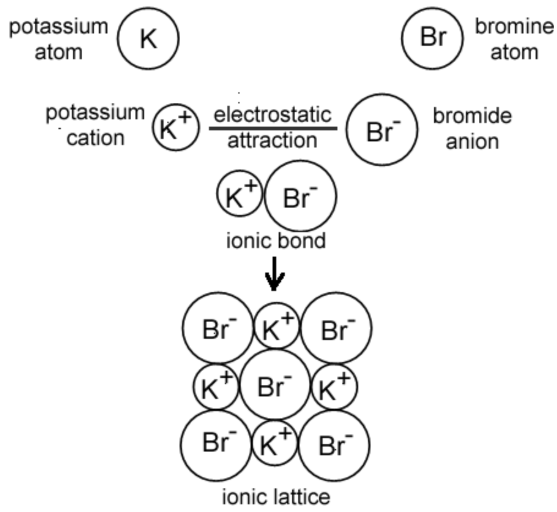
CHEM1611 Worksheet 3: Ionic and Covalent Bonding

Model 1: Ionic Bonding

The compounds formed by metals and non-metals contain *ionic* bonds. Metal atoms lose electrons to form cations. Non-metal atoms gain electrons to form anions. The interactions between these cations and non-metal anions are *ionic* and are often called *ionic bonds.* Simply, it is the coming together of opposite charges in a strict ratio based on *electrostatic attraction*.



It is important to note that ionic compounds form *extended ionic lattices* which contain an 'infinite' networks of ionic bonds.

Group 1 metals will lose one electron. Group 2 metals will lose 2 electrons. Group 17 non-metals will gain one electron. Group 16 non-metals will gain 2 electrons.

Critical thinking questions

1. What charge will a group 2 cation have?

2. What charge will a group 16 anion have?

3. What charge will a group 15 anion have?

4. Using the concepts of valence electrons and noble gas configurations, explain why the different groups gain or lose specific numbers of electrons.

Ionic lattices form in strict ratios of anions to cations to gain overall neutrality. For example, Ca^{2+} and Cl^- will form an ionic lattice in the ratio 1:2 (1 x 2 + 2 x (-1) = 0). The resulting ionic compound will have the formula $CaCl_2$, and will be called calcium chloride. The name of the cation is given first and the anion is named second *with a space between them*. The name of the anion ends in "ide".

5. How many valence electrons do the following atoms have? What charge would you expect each of their ions to have?

   Iodine, sodium, oxygen, aluminium, nitrogen, tellurium, indium, radon, tin

6. What would be the name and formula of an ionic compound made of aluminium and oxygen ions?

7. Which of the following pairs of ions is most likely to form an ionic compound and why?
   Ba^+ O^-, Ba^{2+} O^{2-} or Ba^{3+} O^{3-}

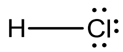
8. What are the names of the following compounds?

(a)	NaCl	(b)	Na_3N	(c)	WC
(d)	CuO	(e)	Cu_2O	(f)	Li_2S
(g)	$CaBr_2$				

Model 2: Lewis Structures

Lewis structures are used to model how the electrons are arranged in a covalent molecule. Dots represent electrons and a line between two atoms represents a bond that is formed by a pair of electrons; double and triple bonds are represented by two and three lines respectively. The Lewis Structure shows how the atoms are connected to each other, but **does not** represent bond length, angles, or 3D shape.

In a Lewis structure, H will have two electrons (*duet rule*) and most other atoms will have 8 electrons (noble gas configuration, the *octet rule*) but be cautious because atoms in the third period and higher can sometimes have more than 8 electrons. Below is one method for drawing Lewis Structures.

Step	Method	HCl
1	Count the valence electrons from all the atoms – this can be determined from the Periodic Table. For anions add and for cations subtract the appropriate number of electrons.	H has 1 electron. Cl has 7 electrons. Total = 8
2	Assemble the bonding framework. Decide which atoms are connected to each other and use a pair of electrons, represented by a line, to form a bond between each pair.	H——Cl
3	Arrange the remaining electrons as lone pairs and/or extra bonds so that each atom has 8 electrons (except H which has 2) including those within the bond.	6 electrons left to place: 
4	Check for exceptions to the octet rule, and that all other atoms are satisfied. Add the overall charge, if needed. You are now complete!	H satisfies duet rule. Cl satisfies octet rule. No charge.

Critical thinking questions

Carbon dioxide, CO_2, is a linear molecule with the arrangement of atoms shown below.

O C O

1. Step 1 for CO_2: how many valence electrons do C and O have?

2. Step 2 for CO_2: draw a bond between C and each O on the figure above.

3. Step 3 for CO_2: add the remaining electrons as lone pairs and/or extra bonds so that each C and O atom have 8 electrons.

4. Step 4 for CO_2: check your Lewis structure conforms to the octet rule for C and O. Are any other arrangements possible that do this?

5. Repeat these steps to draw the Lewis structure for carbon monoxide.

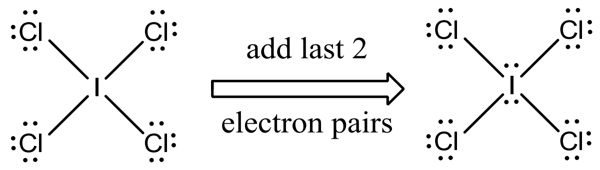
C O

6. Do you predict that the C-O bonds are stronger in CO or in CO_2? Explain.

7. Repeat these steps to draw the Lewis structure for the cyanide ion, CN^-. (*Hint*: this is an anion so do not forget to *add* an electron to the total in Step 1).

8. Do you notice any similarity between the Lewis structures of CO and CN^-?

Atoms in the third period and below can fit more than 8 electrons around themselves. Molecules containing these atoms may be exception s to the octet rule. ICl_4^-, for example, has 7 (I) + 4 × 7 (Cl) + 1 (negative charge) = 36 valence electrons. Drawing the structure such that all the atoms obey the octet rule yields the structure on the left. However, this leaves 4 electrons that have not been used! The extra electrons get placed as lone pairs on the *least electronegative* atom.

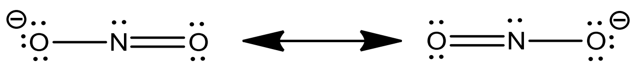


9. How do you know that iodine can be an exception to the octet rule?

10. How do you determine the relative electronegativities of the atoms? What does electronegativity mean?

11. Draw the Lewis structure for PCl_5.

Sometimes, there are two or more Lewis structures that can be drawn. For example, there are two equivalent Lewis structures for NO_2^-, as shown below. Check to make sure that each is a correct Lewis structure!



The possible structures are called *resonance structures* and are drawn with a double headed arrow, as above. The *real* bonding description is an average of the resonance structures. For NO_2^-, this means that each bond is half way between a single and a double bond and the negative charge is shared between the O atoms.

When more than one Lewis structure can be drawn, the electrons that are placed differently are said to be *delocalised*. It is always important to spot delocalisation when it can occur as it often leads to the molecule or ion being unusually stable.

12. Draw the resonance structures for ozone, O_3. (*Hint*: count the number of electrons in NO_2^- and O_3.)

13. Draw the resonance structures for the nitrate ion, NO_3^-.

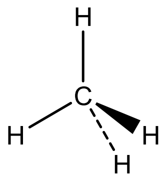
14. By first drawing their Lewis structures, arrange the following molecules in order of increasing carbon-carbon bond strength.

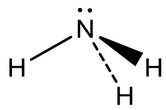
(a) C_2H_6 (b) C_2H_4 (c) C_2H_2

Model 3: Predicting Molecular Shape

The Lewis structures in the previous section provide information about how a molecule is 'put together' and how many bonding electrons and lone pairs there are. They do not provide information about the *shape* of the molecule. However, once the Lewis structure is worked out, you can easily work out the shape.

The 'Valence Shell Electron-Pair Repulsion' (VSEPR) theory states that electron pairs (both bonding and lone) will spread out as evenly as possible around the atom to reduce the repulsion between them. Lone pairs actually take up *more* space than a bonding pair, so if there is a choice, lone pairs will be as far away from each other as possible. The bonding and lone pairs control the shape of the molecule, but only the position of the atoms can be seen.

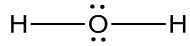
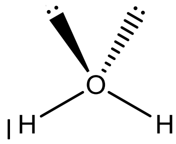
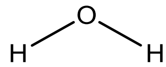




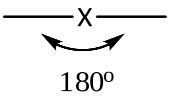
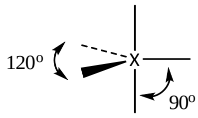
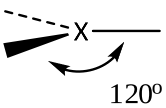
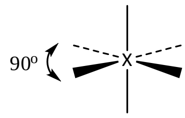
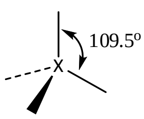
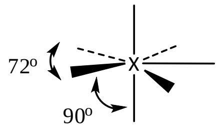
For example, there are 4 C-H bonding pairs around the C atom in CH_4. These arrange themselves in a tetrahedron with 109.5° between the bonds.

In NH_3, there are also 4 pairs around the N atom: 3 N-H bonding pairs and one N lone pair. These also arrange themselves in a tetrahedron. The atoms themselves are at the corners of a pyramid.

To work out the shape of a molecule XY_n, use the following steps:

Step	Method	H_2O	
1	Draw the Lewis structure.		
2	Count the total number of electron pairs - the sum of the number of bonds (n) and the number of lone pairs (m) around the central atom.	bonds = 2, lone pairs = 2 total = 2 + 2 = 4	
3	Arrange these pairs to maximize the distance between them. If there is a choice, maximize the distance between the lone pairs.		tetrahedral – there are no choices in where the lone pairs go
4	What is the geometry of the molecule (ignore lone pairs!)		V=shaped

Common 3D-arrangements of bonds (n) and lone pairs (m) are shown in the table below.

n + m	**Arrangement**		***n + m***	**Arrangement**	
2	linear		5	trigonal bipyramid	
3	trigonal planar		6	octahedral	
4	tetrahedral		7	pentagonal bipyramid	

Critical thinking questions

1. Using the Lewis structure of CO_2 from Model 2, work out the arrangement of the electron pairs and hence the shape of the molecule.

2. Using the Lewis structures of NH_4^+ and PCl_3 from Model 2, work out the arrangement of the electron pairs and draw the shapes of these molecules. What names would you use to describe these shapes?

3. Using the Lewis structure of ICl_4^- from Model 2, work out its shape. (*Hint*: there is a choice of where to place the lone pairs and you should choose the one in which they are as far apart as possible.)