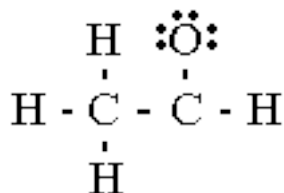
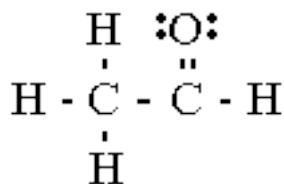


**Step 3.** The structure in Step 2 shows one of two carbon atoms and the oxygen atom without octets. Place lone pairs using the remaining 6 electrons around the oxygen atom.



**Step 4.** Multiple bonds are necessary to give carbon an octet, so a lone pair is moved from the oxygen atom to form a second bond between C and O.



### 3. Resonance Structures

#### Problem

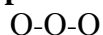
Draw all resonance structures for  $\text{O}_3$ .

#### Solution

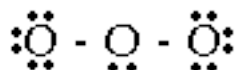
**Step 1.** Add valence electrons.

3 O at 6 electrons each = 18 electrons

**Step 2.** Place a bond between atoms. This uses 4 electrons. 14 electrons remain.



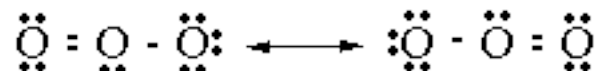
**Step 3.** Place lone pairs of the remaining 14 electrons around the oxygen atoms, starting with the terminal atoms. After each terminal O atom has 8 electrons, 2 electrons remain and are placed on the central atom.



**Step 4.** Use multiple bonds to obtain octets around each O.



**Step 5.** Electron delocalization leads to two resonance structures for  $\text{O}_3$ .



## 4. Lewis Structures

### Problem

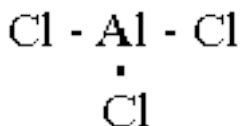
Draw the Lewis Dot Structure for  $\text{AlCl}_3$ .

### Solution

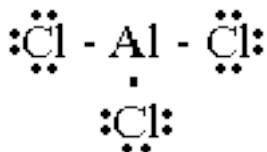
**Step 1.** Add valence electrons.

$$\begin{array}{rcl} 1 \text{ Al at 3 electrons} & = & 3 \text{ electrons} \\ 3 \text{ Cl at 7 electrons each} & = & 21 \text{ electrons} \\ \text{Total} & = & 24 \text{ electrons} \end{array}$$

**Step 2.** Place a bond between the atoms to represent the sharing of 2 electrons. 18 electrons remain.



**Step 3.** Place lone pairs of the remaining 18 electrons around the outside, terminal atoms. Each chlorine atom should have an octet of electrons. Neither Al nor Cl can form multiple bonds, so even though Al has only 6 electrons, this is the final Lewis structure.



## 5. Lewis Structures

### Problem

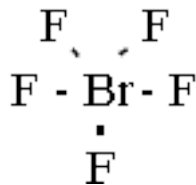
Draw the Lewis Dot Structure for  $\text{BrF}_5$ .

### Solution:

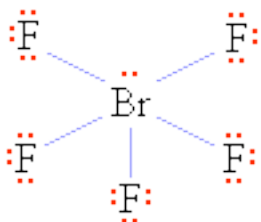
**Step 1.** Add valence electrons.

$$\begin{array}{rcl} 1 \text{ Br at 7 electrons} & = & 7 \text{ electrons} \\ 5 \text{ F at 7 electrons each} & = & 35 \text{ electrons} \\ \text{Total} & = & 42 \text{ electrons} \end{array}$$

**Step 2.** Place a bond between Br and F atoms to represent the sharing of 2 electrons. 32 electrons remain.



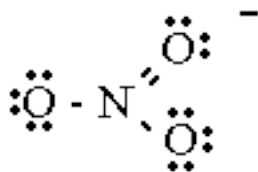
**Step 3.** Place lone pairs of the remaining 32 electrons around the 5 fluorine atoms. Each F atom has an octet of electrons. The remaining electrons are placed on the central atom.



## 6. Oxidation Numbers and Formal Charge

### Problem

Determine the oxidation number and formal charge for N in the nitrate ion,  $\text{NO}_3^-$ .



### Element      Electronegativity Values

H	2.1
B	2.0
C	2.5
N	3.0
O	3.5
F	4.0
Cl	3.0

### Solution

**Step 1.** According to oxidation number rules, all valence electrons are considered to be held by the more electronegative atom. Therefore, each O atom has an oxidation number of  $-2$ . There are three oxygens for a total of  $-6$ . The charge on the ion is  $-1$ . Solving for  $x$ , where  $x$  is the oxidation number for nitrogen:

$$x + (3)(-2) = -1$$

$$x = +5$$

Nitrogen has an oxidation number of  $+5$ .

**Step 2.** The formal charge is based on the assumption that electrons are shared equally between covalently bonded atoms.

$$\text{Formal charge} = (\text{group number in periodic table}) - (\text{number of lone electrons})$$

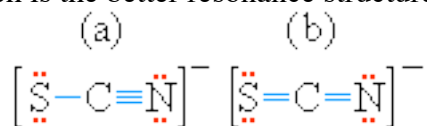
$$- \frac{1}{2} (\text{number of bonding electrons})$$

For nitrogen, formal charge =  $5 - 0 - 4 = +1$ . In this case, N has a formal charge of  $+1$ .

## 7. Resonance Structures

### Question

Which is the better resonance structure for thiocyanate ion,  $\text{SCN}^-$ ?



### Solution

**Step 1.** Determine the formal charges on each element in both resonance structures.

$$\text{Formal charge} = (\text{group number in periodic table}) - (\text{number of lone electrons}) - \frac{1}{2} (\text{number of bonding electrons})$$

	Formal Charge	
	Structure (a)	Structure (b)
Sulfur	+1	0
Carbon	0	0
Nitrogen	-2	-1

**Step 2.** The structure that has the fewest number of atoms with a formal charge is the better structure. Consequently, structure (b) is the correct one.

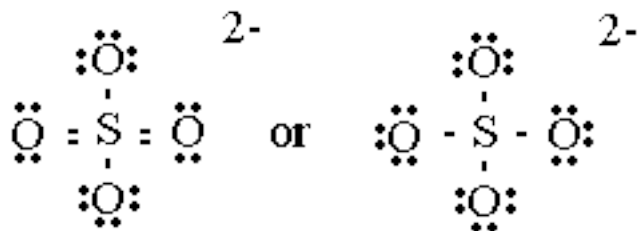
## 8. Using VSEPR: Predicting Geometries

### Question

Using the VSEPR theory, what is the geometry of  $\text{SO}_4^{2-}$ ?

### Solution

**Step 1.** Draw the Lewis Dot Structure. There is more than one.



**Step 2.** Both structures have 4 bonded atoms and no lone pairs on the central S atom. Therefore, both the electron pair geometry and the molecular geometry are tetrahedral.

(Note: the VSEPR theory applies to both ions and molecules.) If the central S had one or more lone pairs, the electron pair and molecular geometries would differ.

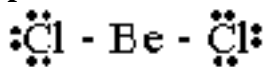
## 9. Using VSEPR: Predicting Geometries

### Question

Using the VSEPR theory, what is the geometry of  $\text{BeCl}_2$ ?

### Solution

**Step 1.** Draw the Lewis Dot Structure.



**Step 2.** Neither Be nor Cl is able to form multiple bonds, so the central atom Be remains electron deficient. There are no lone pairs on the central atom, Be. The electron pair geometry is linear. This molecule is linear.

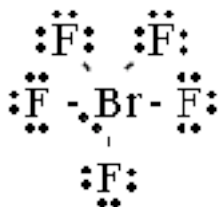
## 10. Using VSEPR: Geometry and Bond Angles

### Problem

Based on the VSEPR theory, predict the electron-pair geometry for  $\text{BrF}_5$  and F-Br-F bond angles.

### Solution

**Step 1.** Draw the Lewis structure of the molecule.



**Step 2.** Determine the number of bonded pairs and lone pairs of electrons around the central atom. In this molecule, there are 5 bonded atoms and 1 lone pair around Br. There are 6 structurally significant electron pairs.

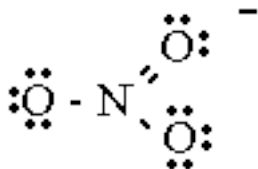
**Step 3.** The electron pair geometry is octahedral. The ideal molecular geometry is square pyramidal.

**Step 4.** The ideal bond angles in an octahedral structure are 90 degrees. The angles in  $\text{BrF}_5$  should be compressed slightly from the ideal 90 degrees angle due to the lone pair of electrons on Br.

## 11. Bond Properties and Resonance

### Problem

$\text{NO}_3^-$  has three resonance structures and a bond order of 1.3. Predict the bond length for N-O.



N-O single bond length = 136 pm

N-O double bond length = 115 pm

N-O triple bond length = 108 pm

### Solution

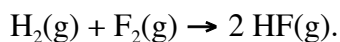
Bond length depends partly on bond order. An N-O bond order of 1.3 suggests that the N-O bond length is a value between that for a single and double N-O bond. The bond length is 122 pm, which is in fact the length of a single N-O bond added to one third the difference between single and double N-O bond lengths.

Note: in order to use the  $\Delta H_{\text{formation}}$  data, all reactants must be in the gas phase.

## 12. Calculating $\Delta H$

### Problem

Calculate  $\Delta H$  for



using the bond energy data below.

$$\text{H-H} = 436 \text{ kJ/mol}$$

$$\text{F-F} = 159 \text{ kJ/mol}$$

$$\text{H-F} = 569 \text{ kJ/mol}$$

### Solution

In this reaction, we need to break the bonds between H-H and F-F and will form 2 H-F bonds. Because all the reactants and products are in the gas phase, we can relate  $\Delta H$  for the reaction to the strengths of bonds broken and formed.

$$\Delta H = (\text{energy of bonds broken}) - (\text{energy of bonds formed})$$

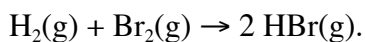
$$\Delta H = [(1 \text{ mol})(436 \text{ kJ/mol}) + (1 \text{ mol})(159 \text{ kJ/mol})] - (2 \text{ mol})(569 \text{ kJ/mol})$$

$$\Delta H = -543 \text{ kJ}$$

## 13. Calculating $\Delta H$

### Problem

Calculate  $\Delta H$  for



using the bond energy data given below.

$$\text{H-H} = 436 \text{ kJ/mol}$$

$$\text{Br-Br} = 192 \text{ kJ/mol}$$

$$\text{H-Br} = 368 \text{ kJ/mol}$$

### Solution

In this reaction, we need to break the bonds between H-H and Br-Br and will form 2 H-Br bonds.

$$\Delta H = (\text{energy of bonds broken}) - (\text{energy of bonds formed})$$

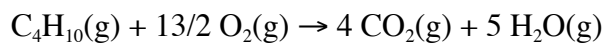
$$\Delta H = [(1 \text{ mol})(436 \text{ kJ/mol}) + (1 \text{ mol})(192 \text{ kJ/mol})] - (2 \text{ mol})(368 \text{ kJ/mol})$$

$$\Delta H = -108 \text{ kJ}$$

## 14. Calculating $\Delta H$

### Problem

Calculate  $\Delta H$  for



using the bond energy data given below.

$$\text{C-C} = 331 \text{ kJ/mol}$$

$$\text{C-H} = 414 \text{ kJ/mol}$$

$$\text{O=O} = 498 \text{ kJ/mol}$$

$$\text{C=O} = 803 \text{ kJ/mol}$$

$$\text{O-H} = 464 \text{ kJ/mol}$$

### Solution

In this reaction, find that 3 C-C, 10 C-H, and  $13/2$  O=O bonds are being broken, and 8 C=O and 10 O-H bonds are being formed. Thus

$$\Delta H = (\text{energy of bonds broken}) - (\text{energy of bonds formed})$$

$$\Delta H = [(3 \text{ mol})(331 \text{ kJ/mol}) + (10 \text{ mol})(414 \text{ kJ/mol}) + (13/2 \text{ mol})(498 \text{ kJ/mol})] - [(8 \text{ mol})(803 \text{ kJ/mol}) + (10 \text{ mol})(464 \text{ kJ/mol})]$$

$$\Delta H = -2694 \text{ kJ}$$

## 15. Bond Polarity

### Problem

Arrange the following covalent bonds in order of increasing polarity:

O-H   I-Br   C-F   P-H   S-Cl.

#### Electronegativity Values

<b>H</b>	2.1
<b>B</b>	2.0
<b>P</b>	2.1
<b>I</b>	2.5
<b>C</b>	2.5
<b>S</b>	2.5
<b>Br</b>	2.8
<b>N</b>	3.0
<b>O</b>	3.5
<b>F</b>	4.0
<b>Cl</b>	3.0

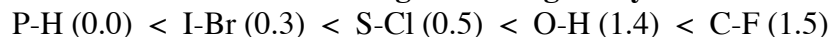
## Approach

Consider the differences in electronegativity values for each bonded pair of atoms.

## Solution

The polarity of a bond increases with increasing difference in electronegativity of the bonded atoms. The covalent bonds, therefore, have the following order of increasing polarity with the electronegativity differences shown in parentheses:

→ **Increasing electronegativity** →



## 16. Molecular Polarity

### Question

Is  $\text{NF}_3$  polar or nonpolar?

### Solution

$\text{NF}_3$  has the same pyramidal structure as  $\text{NH}_3$ . Fluorine is more electronegative than N, thus each bond is polar. The  $\text{NF}_3$  molecule is asymmetrical and polar.

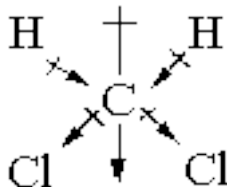
## 17. Molecular Polarity

### Question

Is  $\text{CH}_2\text{Cl}_2$  polar or nonpolar?

### Solution:

In  $\text{CH}_2\text{Cl}_2$ , chlorine atoms are the most electronegative, followed by carbon atoms, then hydrogen atoms.

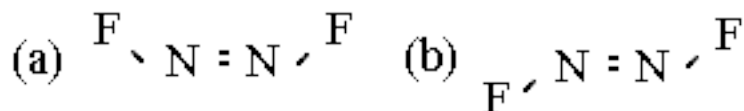


There is a net movement of electron density away from H atoms and toward Cl atoms. The asymmetric arrangement of the atoms, and the negative end of the bond, is toward the two Cl atoms. The positive end toward the 2 H atoms makes  $\text{CH}_2\text{Cl}_2$  a polar molecule.

## 18. Molecular Polarity

### Problem

There are two different molecules with the formula  $\text{N}_2\text{F}_2$ . Is either molecule polar?



### Solution

Fluorine atoms are more electronegative than nitrogen atoms. The negative ends of the bonds are toward the two fluorine atoms and the positive ends are toward the two nitrogen atoms. In molecule (a), the two dipoles do not cancel each other, thus this molecule is polar. In molecule (b), however, the two dipoles are opposite in direction and do cancel each other, making this molecule nonpolar.