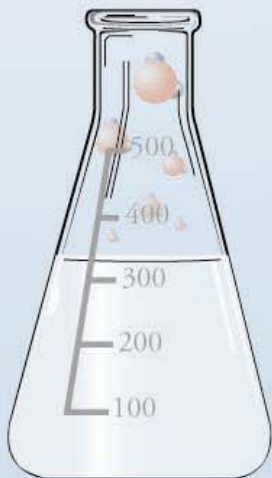


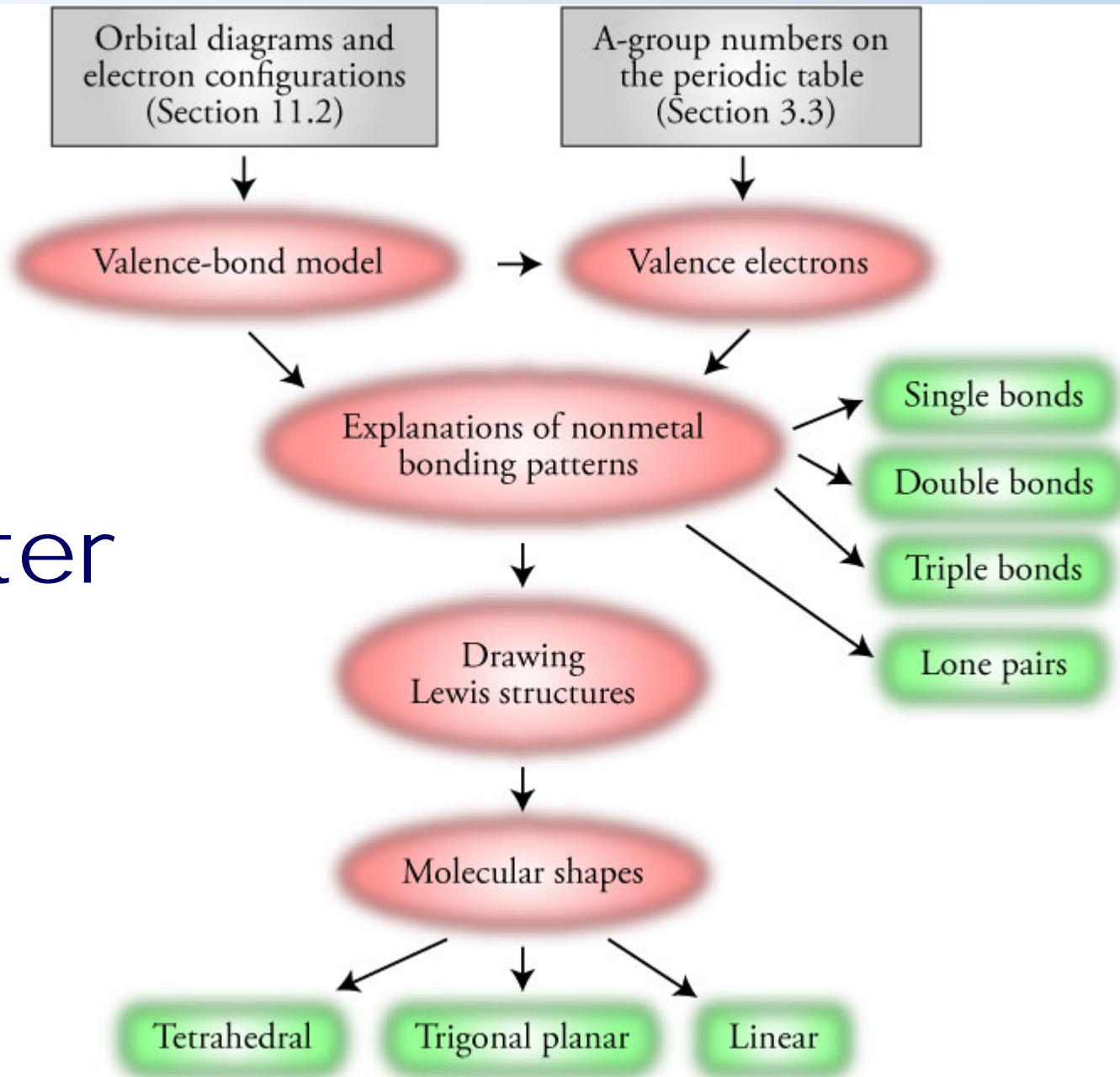


# Chapter 12

## Molecular Structure

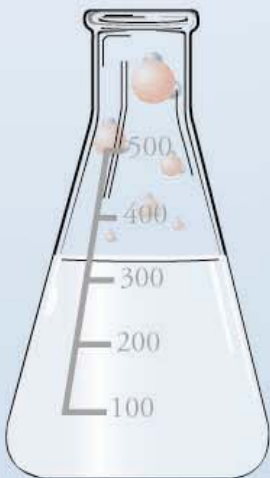


# Chapter Map



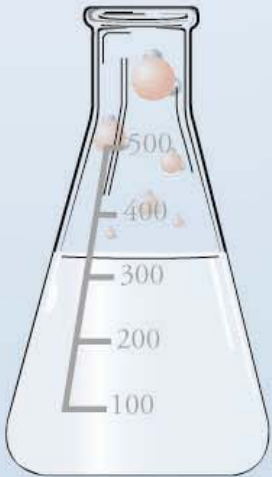
# Models – Advantages and Disadvantages (1)

- They help us to *visualize*, *explain*, and *predict* chemical changes.
- Because a model is a *simplified* version of what we think is true, the processes it depicts are sometimes described using the phrase *as if*. When you read, “It is as if an electron were promoted from one orbital to another,” the phrase is a reminder that we do not necessarily think this is what really happens. We merely find it useful to talk about the process as if this is the way it happens.



# Models – Advantages and Disadvantages (2)

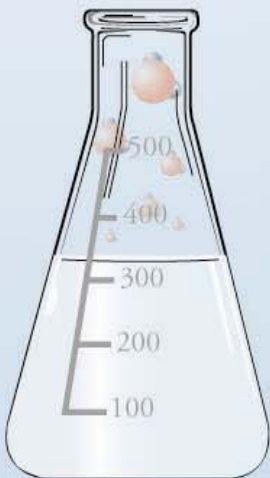
- One characteristic of models is that they ***change with time***. Because our models are simplifications of what we think is real, we are not surprised when they sometimes fail to explain experimental observations. When this happens, the model is altered to fit the new observations.





# Assumptions of the Valence-Bond Model

- Only the highest energy electrons participate in bonding.
- Covalent bonds usually form to pair unpaired electrons.



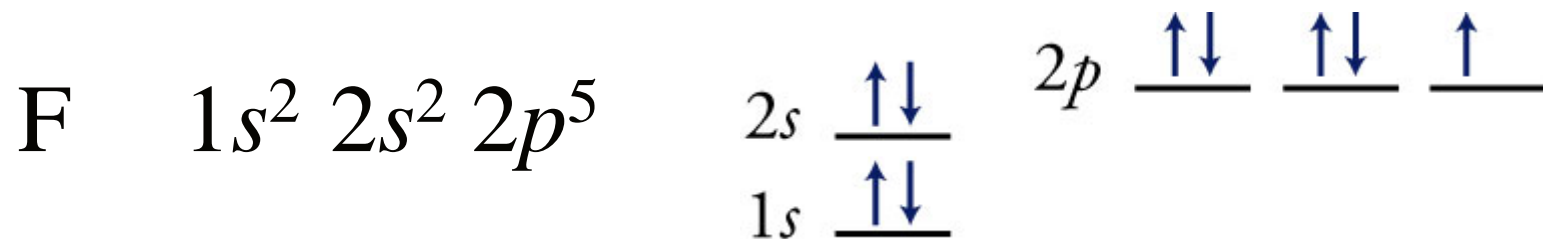
# Valence Electrons

- **Valence electrons** are the highest-energy *s* and *p* electrons in an atom.

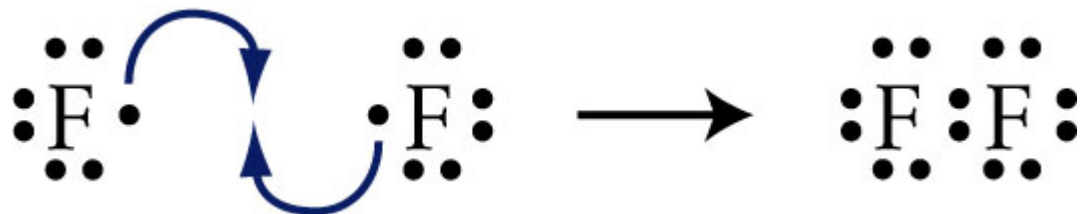
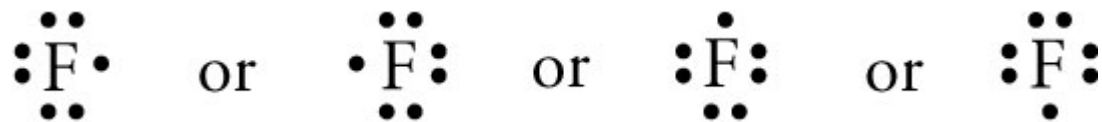
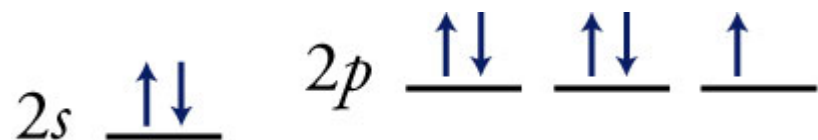
One valence electron  
|

Number of valence electrons equals the A-group number							
					8A		
1A	2A	3A	4A	5A	6A	7A	8A
1 H							2 He
	5 B	6 C	7 N	8 O	9 F	10 Ne	
			15 P	16 S	17 Cl	18 Ar	
			33 As	34 Se	35 Br	36 Kr	
				52 Te	53 I	54 Xe	

# Fluorine

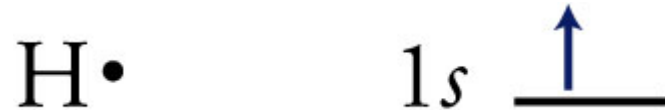


Valence electrons



# H<sub>2</sub> Formation

- The unpaired electron on a hydrogen atom makes the atom unstable.

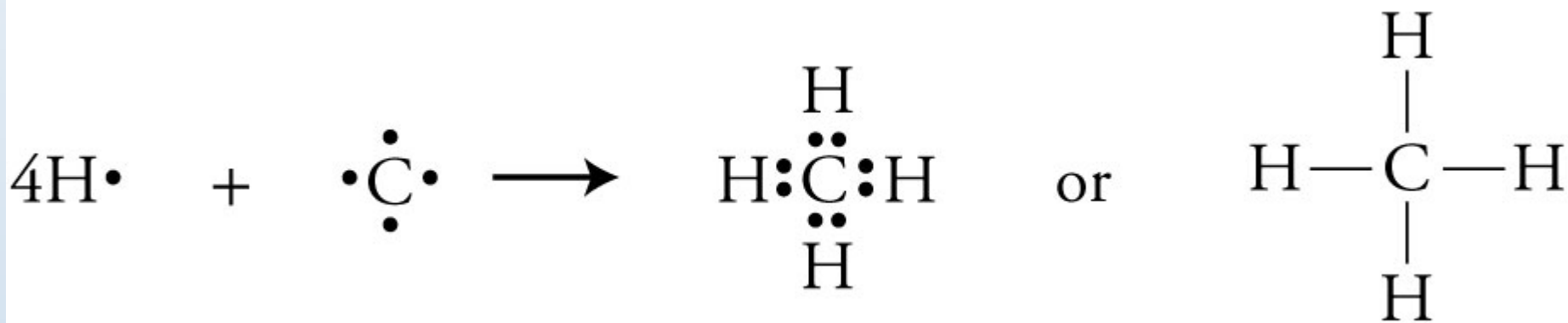
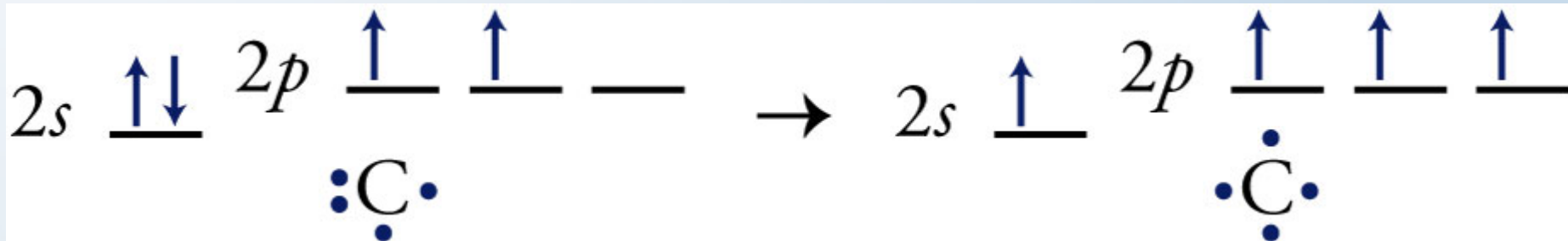


- Two hydrogen atoms combine to form one hydrogen molecule.

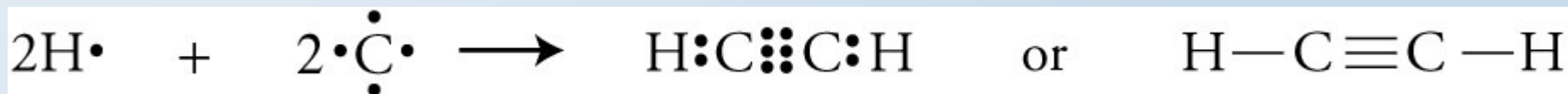
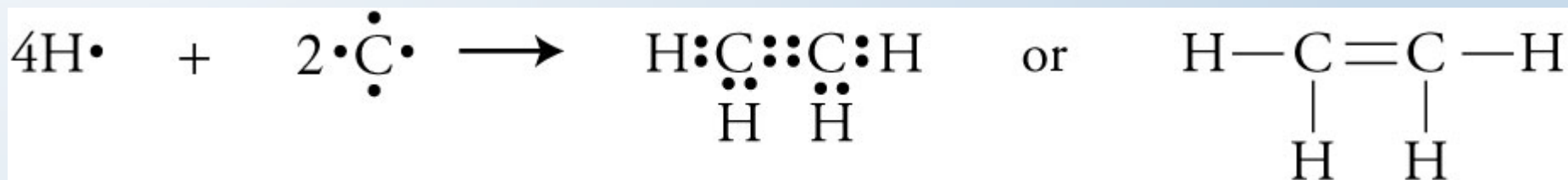




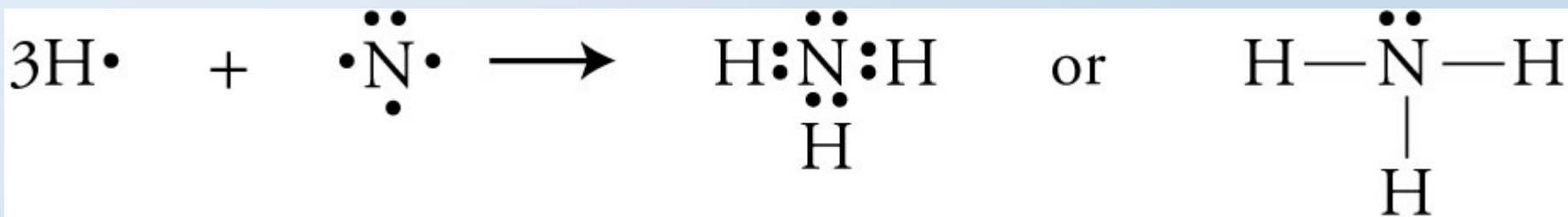
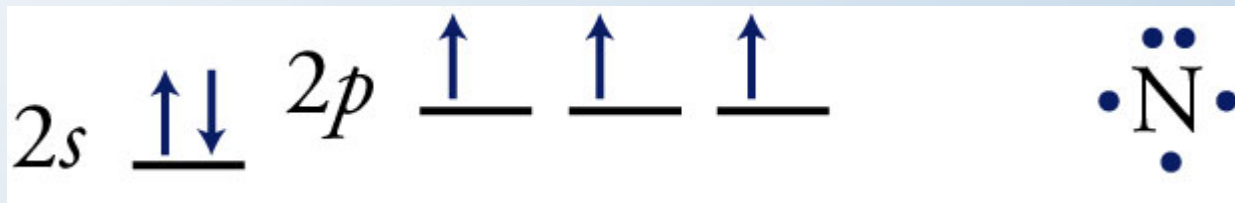
# Carbon - 4 bonds



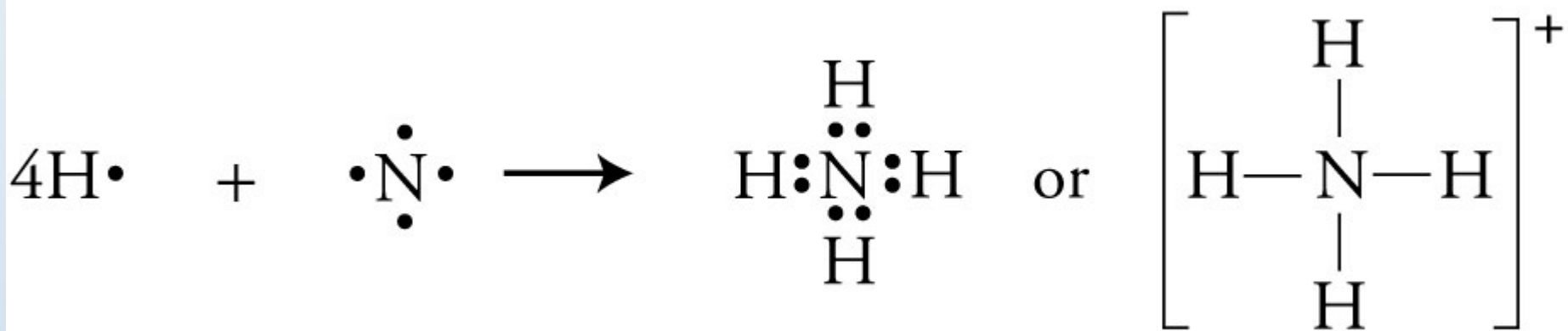
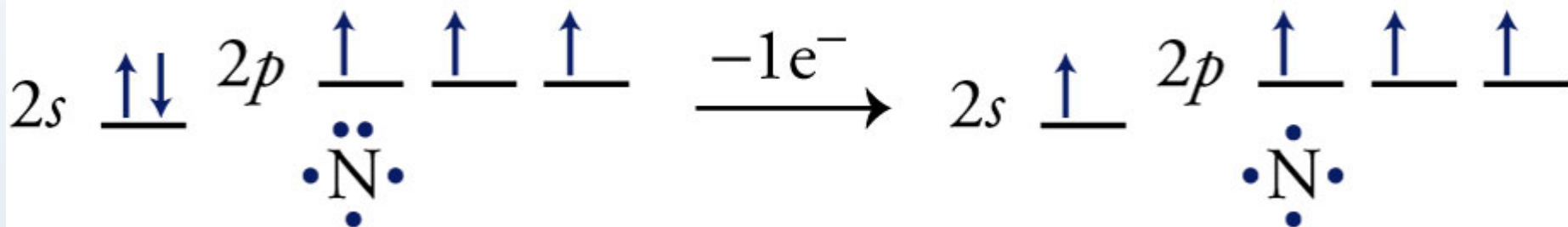
# Carbon – Multiple Bonds



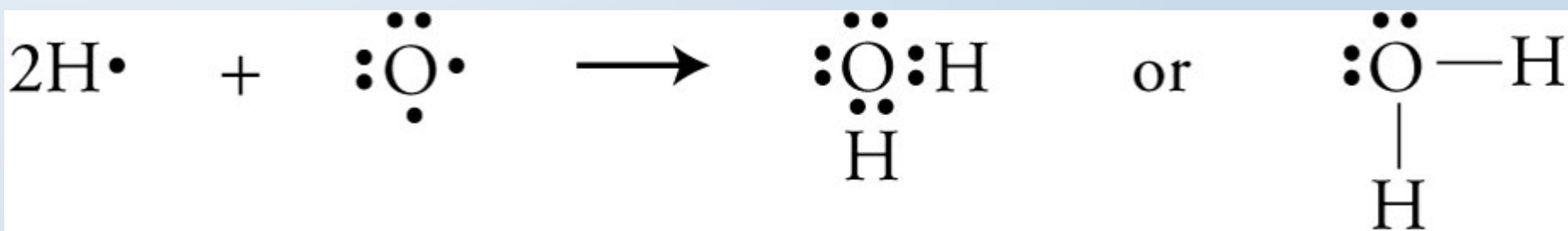
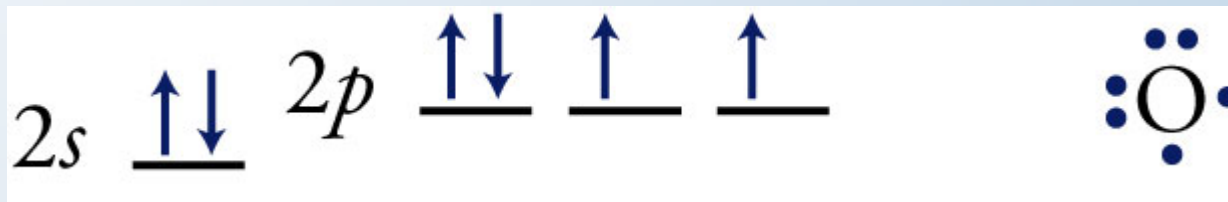
# Nitrogen – 3 bonds & 1 lone pair



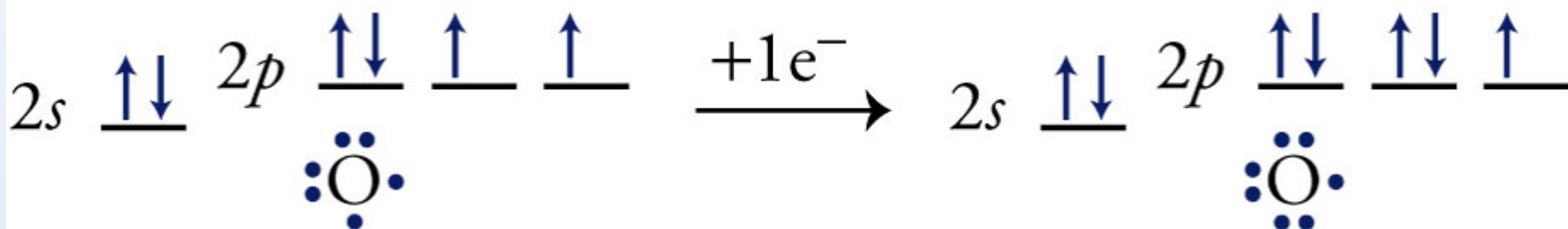
# Nitrogen – 4 bonds



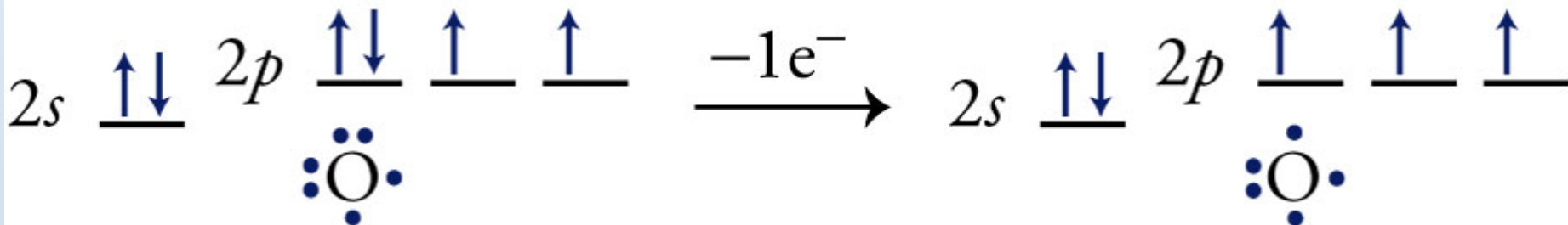
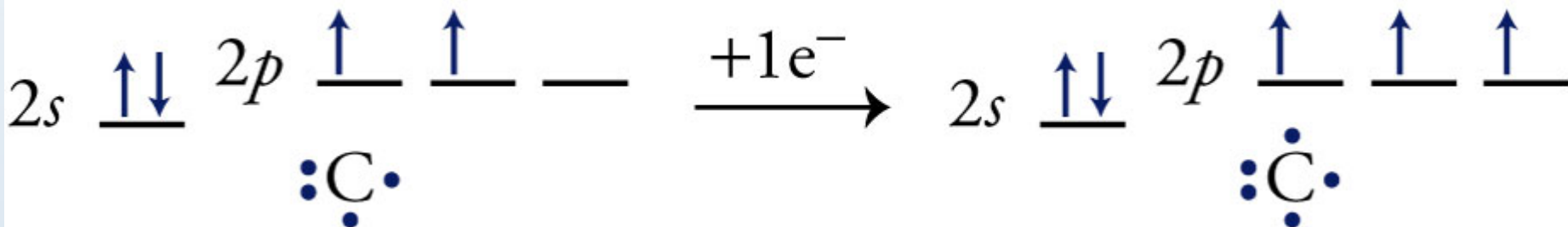
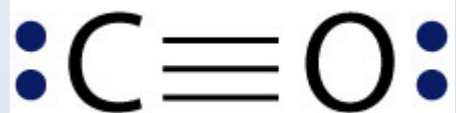
Oxygen – 2 bonds & 2 lone pairs



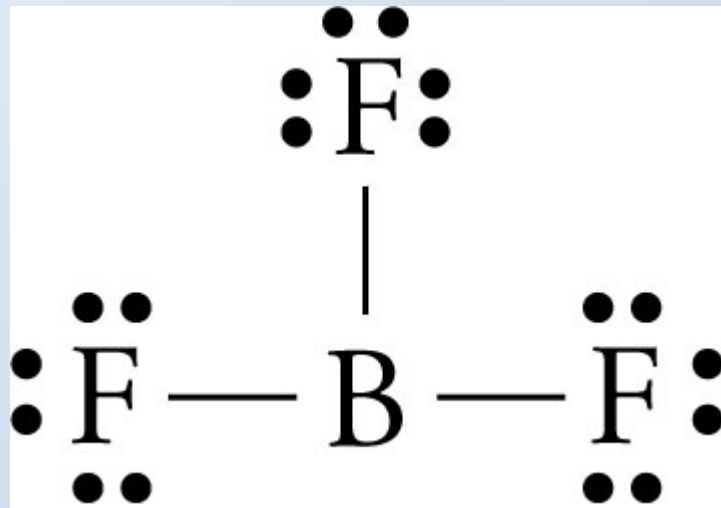
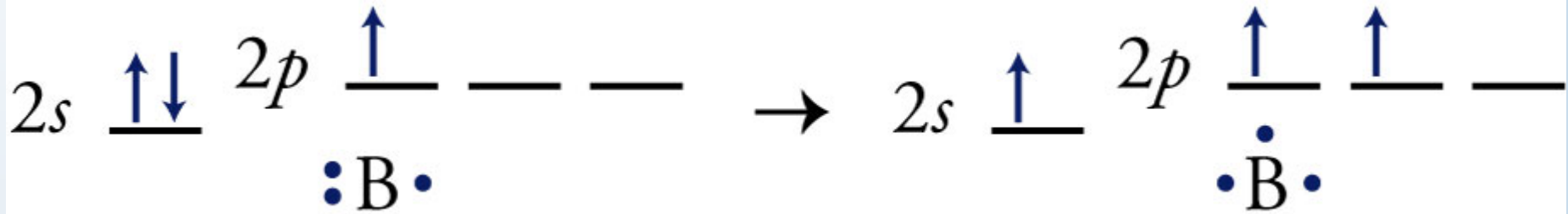
# Oxygen – 1 bond & 3 lone pairs



Carbon – 3 bonds & 1 lone pair  
 Oxygen – 3 bonds & 1 lone pair

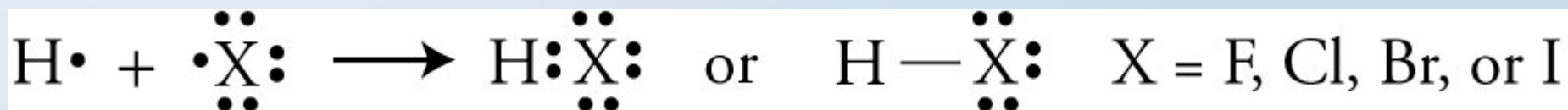
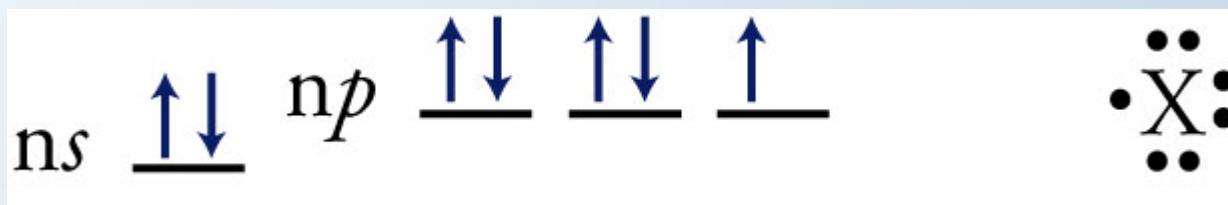


# Boron – 3 bonds





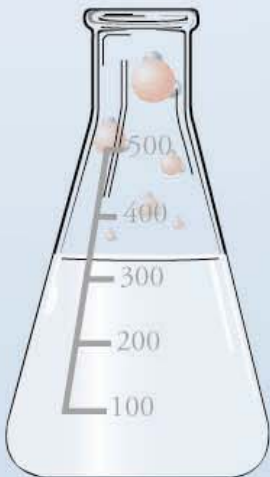
# Halogens – 1 bond & 3 lone pairs





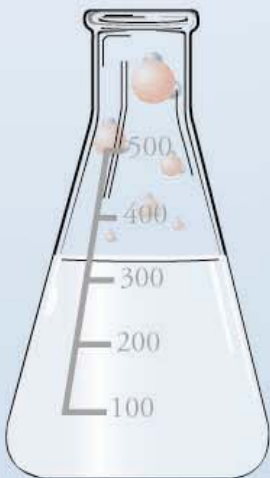
# Drawing Lewis Structures (1)

- **Step 1:** Determine the total number of valence electrons for the molecule or polyatomic ion.
  - For uncharged molecules, the total number of valence electrons is the sum of the valence electrons of each atom.
  - For polyatomic cations, the total number of valence electrons is the sum of the valence electrons for each atom minus the charge.
  - For polyatomic anions, the total number of valence electrons is the sum of the valence electrons for each atom plus the charge.



# Drawing Lewis Structures (2)

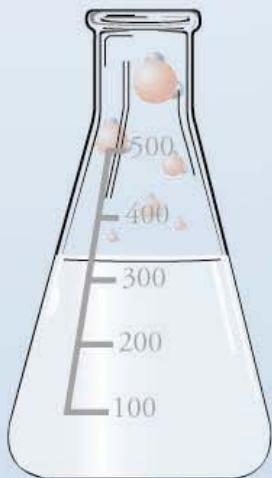
- **Step 2:** Draw a reasonable skeletal structure, using single bonds to join all the atoms.
  - Try to arrange the atoms to yield the most typical number of bonds for each atom.
  - Apply the following guidelines in deciding what element belongs in the center of your structure.
    - Hydrogen and fluorine atoms are never in the center.
    - Oxygen atoms are rarely in the center.
    - The element with the fewest atoms in the formula is often in the center.
    - The atom that is capable of making the most bonds is often in the center.
  - Oxygen atoms rarely bond to other oxygen atoms.
  - The molecular formula often reflects the molecular structure.
  - Carbon atoms commonly bond to other carbon atoms.



A vertical column of water molecules (H<sub>2</sub>O) is positioned on the left side of the slide. Each molecule consists of one large red sphere (oxygen) and two smaller black spheres (hydrogen) bonded to it. The molecules are arranged in a descending staircase pattern from the top left towards the bottom left.

# Drawing Lewis Structures (3)

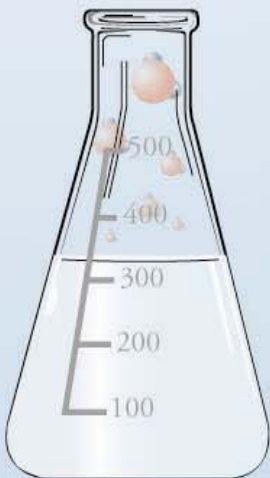
- **Step 3:** Subtract 2 electrons from the total for each of the single bonds (lines) described in Step 2.



A decorative border on the left side of the slide consists of several water molecules. Each molecule is represented by a large red sphere (oxygen) and two smaller black spheres (hydrogen) arranged in a bent shape. The molecules are scattered vertically from the top left towards the bottom left.

# Drawing Lewis Structures (4)

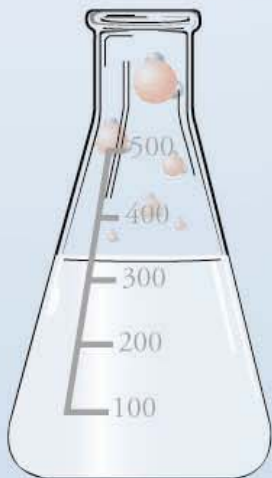
- **Step 4:** Try to distribute the remaining electrons as lone pairs to obtain a total of eight electrons around each atom except hydrogen and boron.
  - In a reasonable Lewis structure, carbon, nitrogen, oxygen, and fluorine always have eight electrons around them.
  - Hydrogen will always have a total of two electrons from its one bond.
  - Boron can have fewer than eight electrons but never more than eight.
  - The nonmetallic elements in periods beyond the second period (P, S, Cl, Se, Br, and I) usually have eight electrons around them, but they can have more.
  - The bonding properties of the metalloids arsenic, As, and tellurium, Te, are similar to those of phosphorus, P, and sulfur, S, so they usually have eight electrons around them but can have more.





# Drawing Lewis Structures (5)

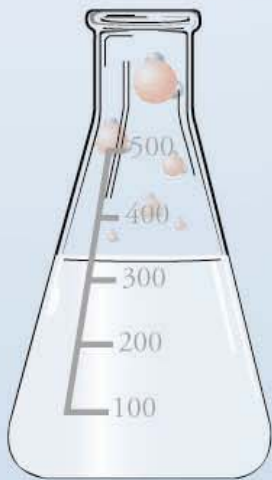
- **Step 5:** Do one of the following.
  - If in Step 4 you were able to obtain an octet of electrons around each atom other than hydrogen and boron, and if you used all of the remaining valence electrons, go to Step 6.
  - If you have electrons remaining after each of the atoms other than hydrogen and boron have their octet, you can put more than eight electrons around elements in periods beyond the second period. (You will not need to use this procedure for any of the structures in this text, but if you take more advanced chemistry courses, it will be useful.)
  - If you do not have enough electrons to obtain octets of electrons around each atom (other than hydrogen and boron), convert one lone pair into a multiple bond for each two electrons that you are short.



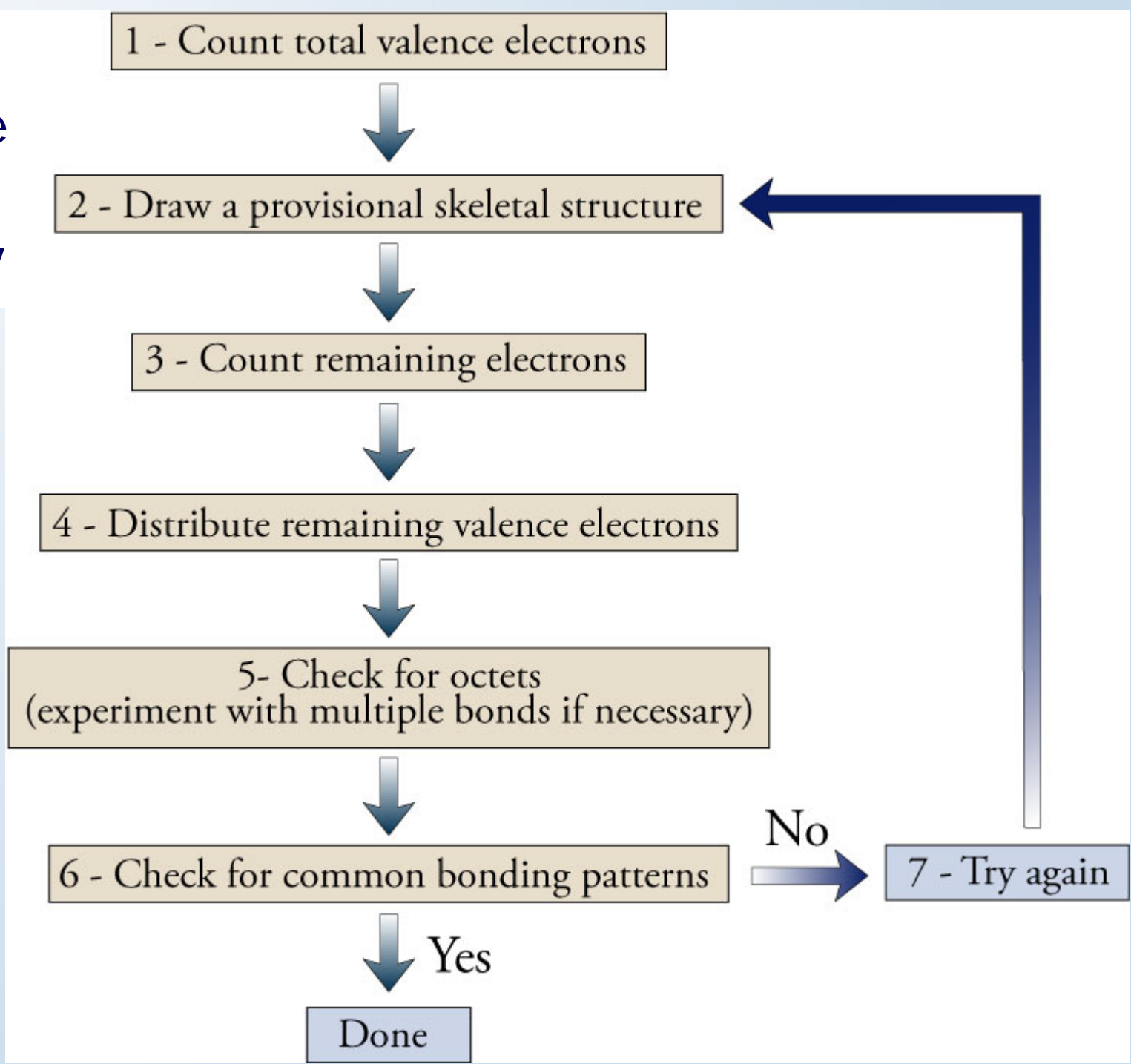
A decorative border on the left side of the slide consists of several water molecules (H<sub>2</sub>O) represented as red spheres (oxygen) with two smaller black spheres (hydrogen) attached. The molecules are arranged in a vertical column, with some appearing to be inside a flask at the bottom.

# Drawing Lewis Structures (6 & 7)

- **Step 6:** Check your structure to see if all of the atoms have their most common bonding pattern.
- **Step 7:** If necessary, try to rearrange your structure to give each atom its most common bonding pattern. One way to do this is to return to Step 2 and try another skeleton. (This step is unnecessary if all of the atoms in your structure have their most common bonding pattern.)



# Lewis Structure Drawing Summary





A series of water molecules, each consisting of one red oxygen atom and two black hydrogen atoms, are shown falling from the top left towards a flask at the bottom left. The flask is a standard Erlenmeyer flask with a scale on its side, ranging from 100 to 500. The water level is indicated at approximately 350. The molecules are arranged in a vertical line, suggesting a flow or a sequence of events.

# Resonance

- We can view certain molecules and polyatomic ions as if they were able to resonate—to switch back and forth—between two or more different structures. Each of these structures is called a ***resonance structure***. The hypothetical switching from one resonance structure to another is called ***resonance***.

# To draw Resonance Forms

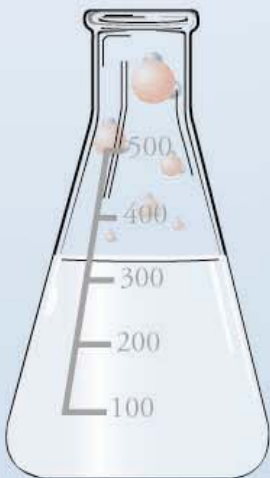
It is as if this lone pair forms a second bond...

...pushing the electrons in this bond off to form a lone pair.

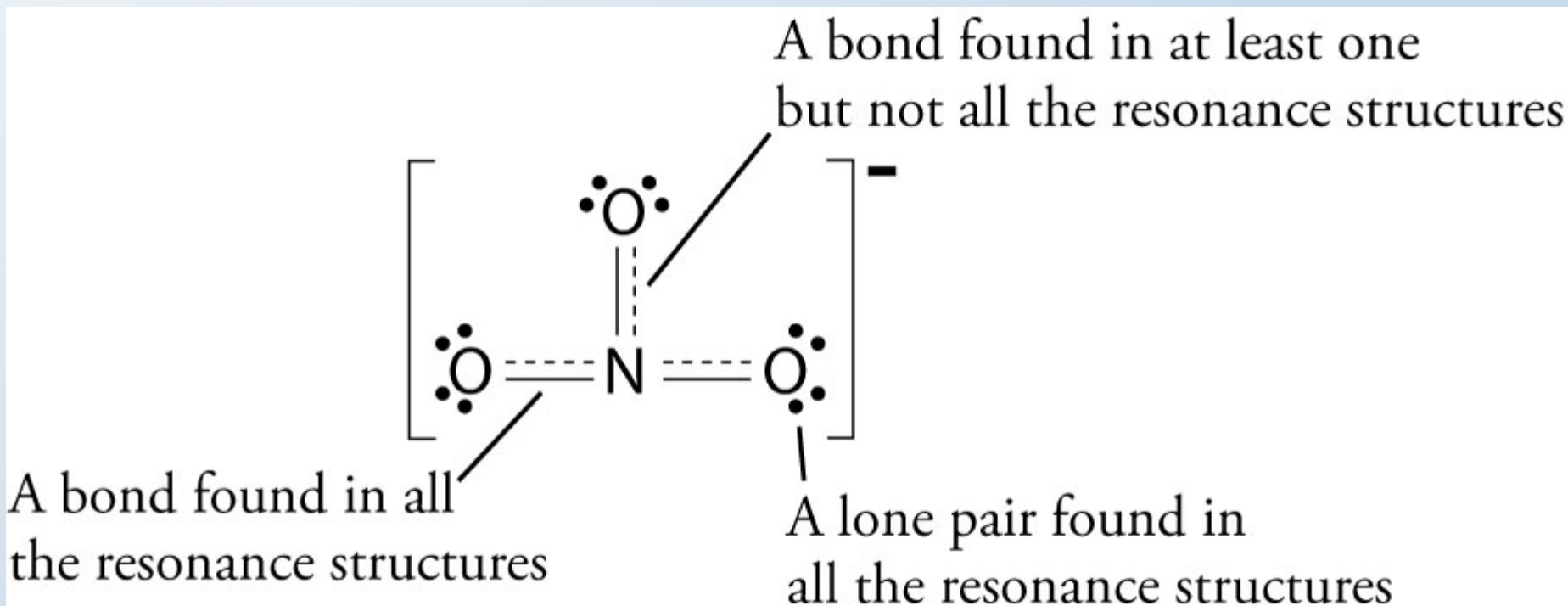
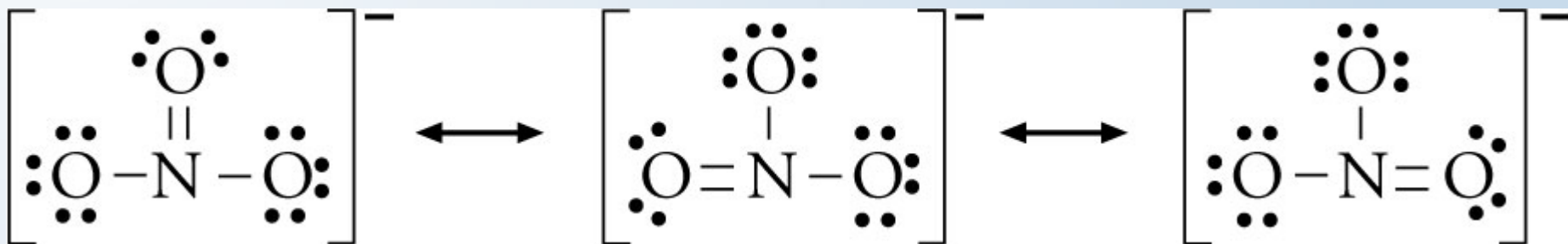


# Resonance Hybrid

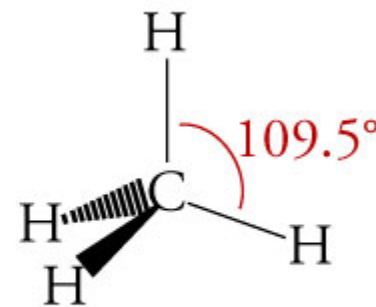
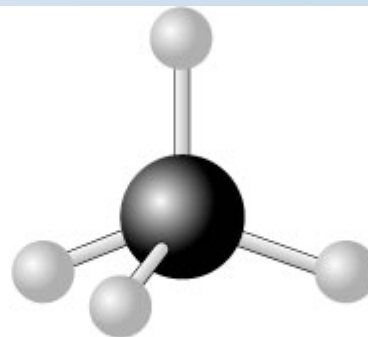
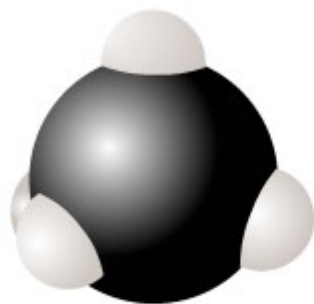
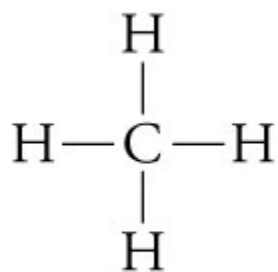
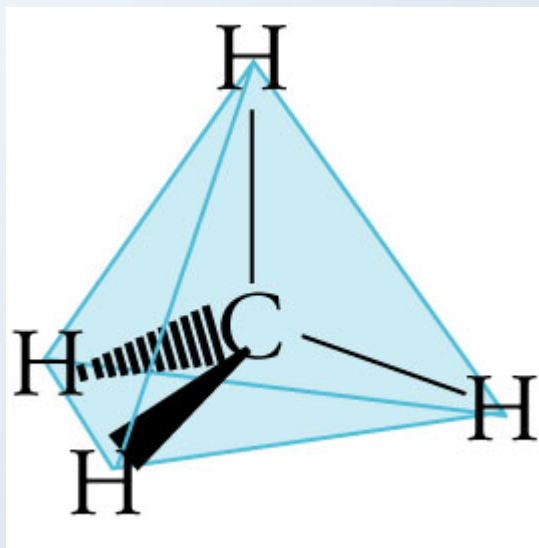
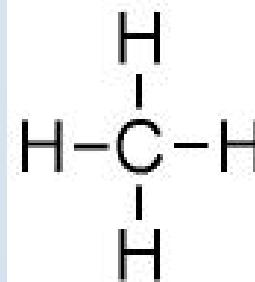
- To blend the resonance structures into a single resonance hybrid:
  - Step 1: Draw the skeletal structure, using solid lines for the bonds that are found in all of the resonance structures.
  - Step 2: Where there is sometimes a bond and sometimes not, draw a dotted line.
  - Step 3: Draw only those lone pairs that are found on every one of the resonance structures. (Leave off the lone pairs that are on one or more resonance structure but not on all of them.)



# Nitrate Resonance



# Methane, CH<sub>4</sub>



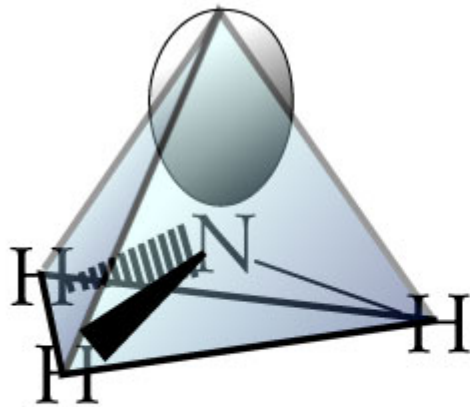
Lewis structure

Space-filling model

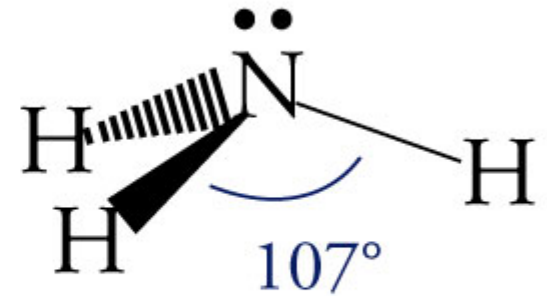
Ball-and-stick model

Geometric Sketch

# Ammonia, $\text{NH}_3$ , Geometry

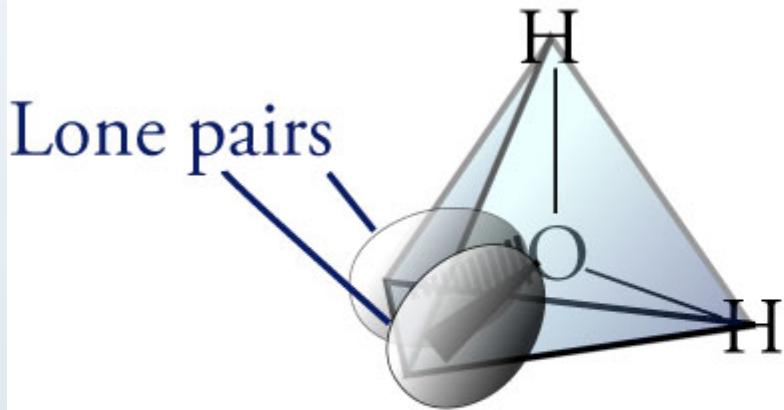


Electron group geometry  
(tetrahedral)

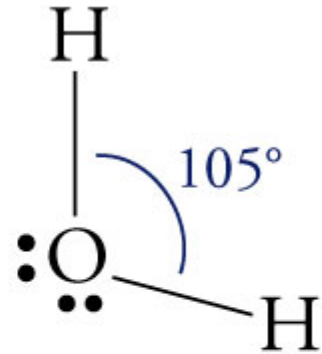


Molecular geometry  
(trigonal pyramid)

# Water, $\text{H}_2\text{O}$ , Geometry

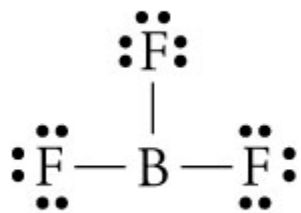


Electron group geometry  
(tetrahedral)

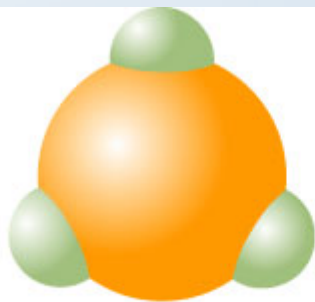


Molecular geometry  
(bent)

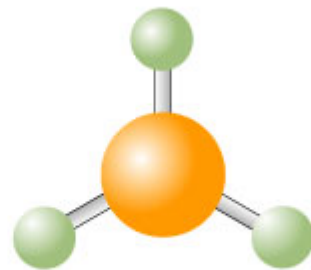
# Trigonal Planar Geometry - $\text{BF}_3$



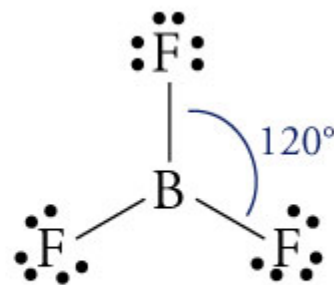
Lewis structure



Space-filling model



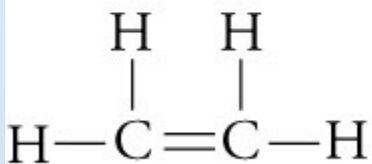
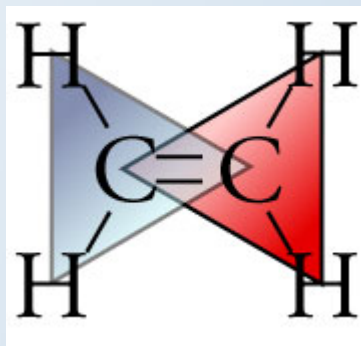
Ball-and-stick model



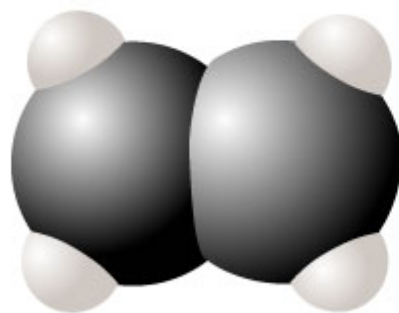
Geometric Sketch



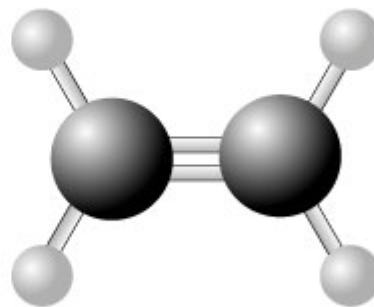
# Ethene (ethylene)



Lewis structure

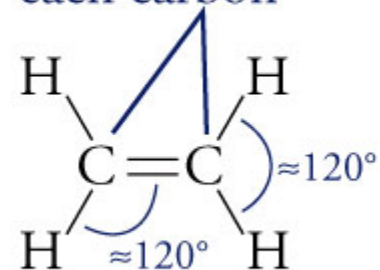


Space-filling model



Ball-and-stick model

Trigonal planar  
geometry around  
each carbon

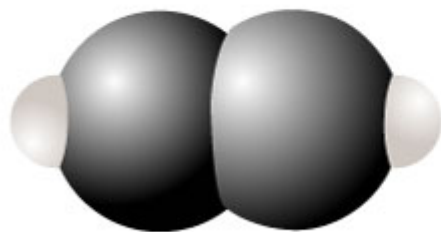


Geometric Sketch

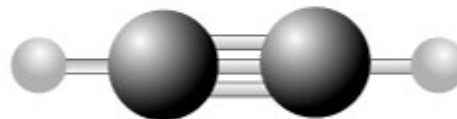
# Ethyne (acetylene) , $C_2H_2$



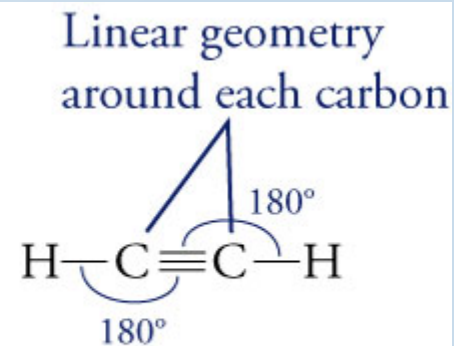
Lewis structure



Space-filling model



Ball-and-stick model

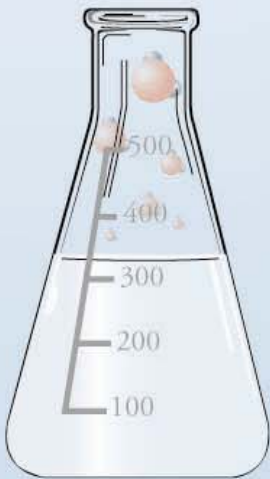


Geometric Sketch



# Steps for Molecular Geometry

- **Step 1:** To determine the name of the electron group geometry around each atom that is attached to two or more atoms, count the number of electron groups around each atom and apply the guidelines found on Table 12.3.
- **Step 2:** Use one or more of the geometric sketches shown on Table 12.3 for the geometric sketch of your molecule.



A decorative border on the left side of the slide consists of several water molecules (H<sub>2</sub>O) arranged in a vertical line. Each molecule is represented by a large red sphere (oxygen) and two smaller black spheres (hydrogen) bonded to it.

## Steps for Molecular Geometry (cont.)

- **Step 3:** To determine the name of the molecular geometry around each atom that has two or more atoms attached to it, count the number of bond groups and lone pairs, and then apply the guidelines found on Table 12.3.

