Chapter 10
Chemical Bonding

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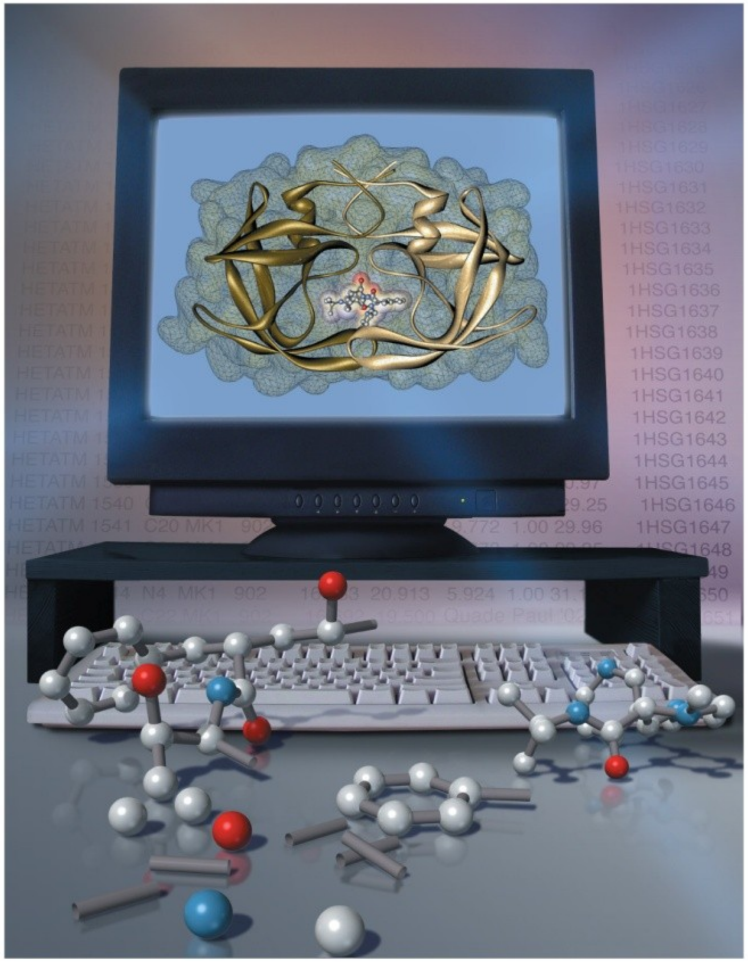
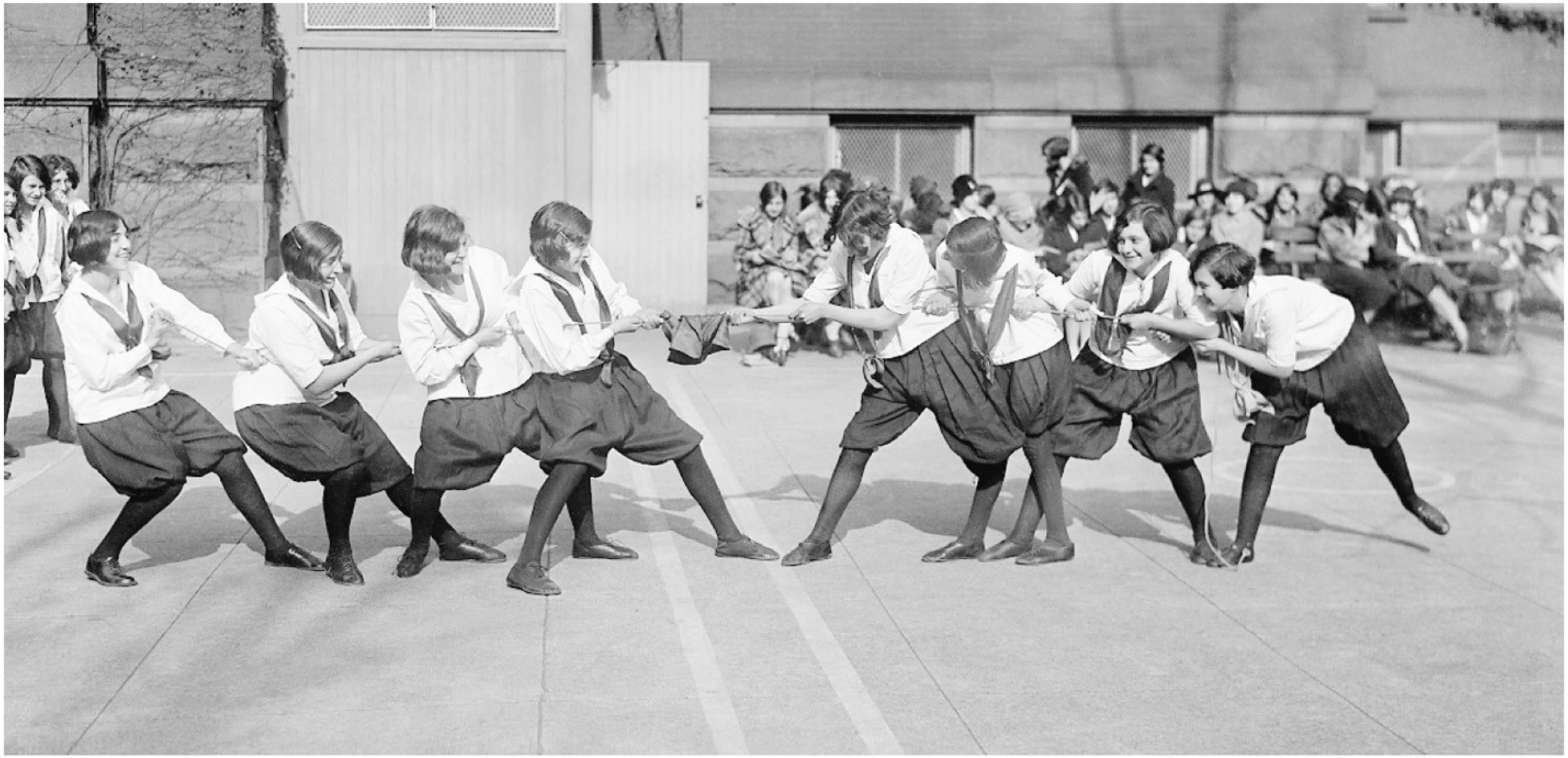
Chem 118
Introductory Chemistry



Map: Introductory Chemistry (Tro) https://chem.libretexts.org/@go/page/45050 (accessed Mar 25, 2022).



The two teams are joined together because both are holding onto the same rope. In a similar way, two atoms are bonded together when both hold onto the same electrons. A covalent bond is a bond formed when atoms share electrons.

Bonding Theories

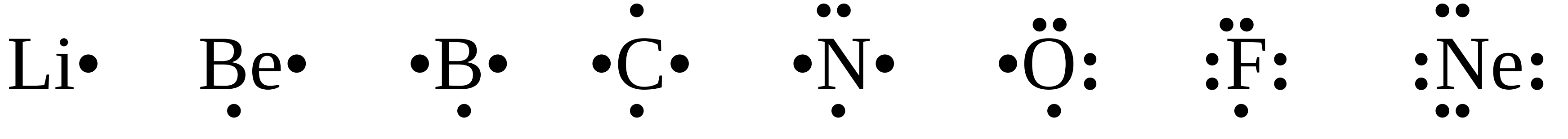
- Central theme in chemistry: How and Why atoms attach together
- This will help us understand how to:
  1. Predict the shapes of molecules.
  2. Predict properties of substances.
  3. Design and build molecules with particular sets of chemical and physical properties.

Lewis Bonding Theory

- Built on the idea of valence electrons.
- Atoms share there valence electrons.
- Gives atoms stablity.

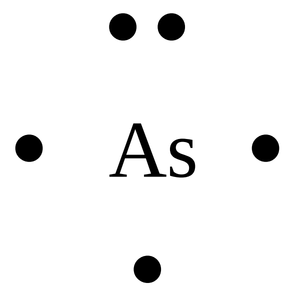
Lewis Symbols of Atoms

- Uses symbol of element to represent nucleus and inner electrons.
- Uses dots around the symbol to represent valence electrons.
  - ✓ Puts one electron on each side first, then pair.
- **Remember that elements in the same group have the same number of valence electrons; therefore, their Lewis dot symbols will look alike.**



Practice—Write the Lewis Symbol for Arsenic, Continued.

- As is in group 15, therefore it has 5 valence electrons.



Lewis Bonding Theory

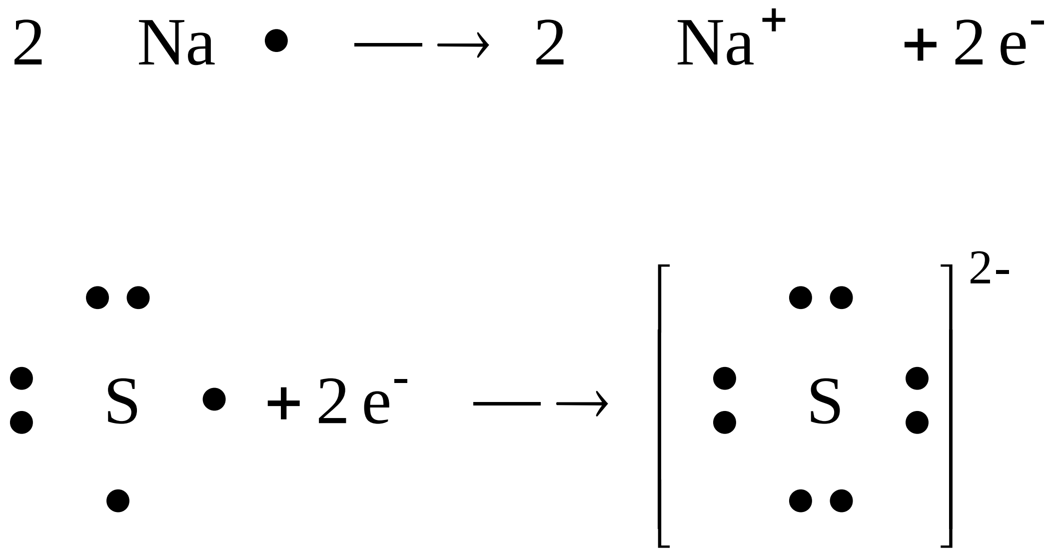
- Atoms ONLY come together for a single reason:
- to produce a more stable electron configuration.
- Atoms bond together by either transferring or sharing electrons.
- A lot of atoms like to have 8 electrons in their outer shell.
  - ✓ Octet rule.
  - ✓ There are some exceptions to this rule—the key to remember is to try to get an electron configuration like a noble gas.
    - ➢ Li and Be try to achieve the He electron arrangement.

Lewis Symbols of Ions

- Cations have Lewis symbols without valence electrons.
  - ✓ Lost in the cation formation.
  - ✓ They now have a full “outer” shell that was the previous second highest energy shell.
- Anions have Lewis symbols with 8 valence electrons.
  - ✓ Electrons gained in the formation of the anion.

Li• Li^+ :F: $[:F:]^-$

Practice—Show How the Electrons Are Transferred and the Bond Is Formed When Na Reacts with S, Continued.



Ionic Bonds

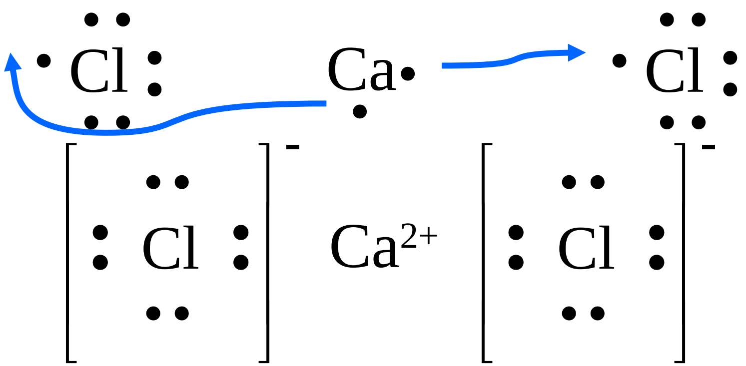
- Metal to nonmetal.
- Metal loses electrons to form cation.
- Nonmetal gains electrons to form anion.
- Ionic bond results from + to − attraction.
  - ✓Larger charge = stronger attraction.
  - ✓Smaller ion = stronger attraction.
- Lewis theory allows us to predict the correct formulas of ionic compounds.

Example 10.3—Using Lewis Theory to Predict Chemical Formulas of Ionic Compounds

Predict the formula of the compound that forms between calcium and chlorine.

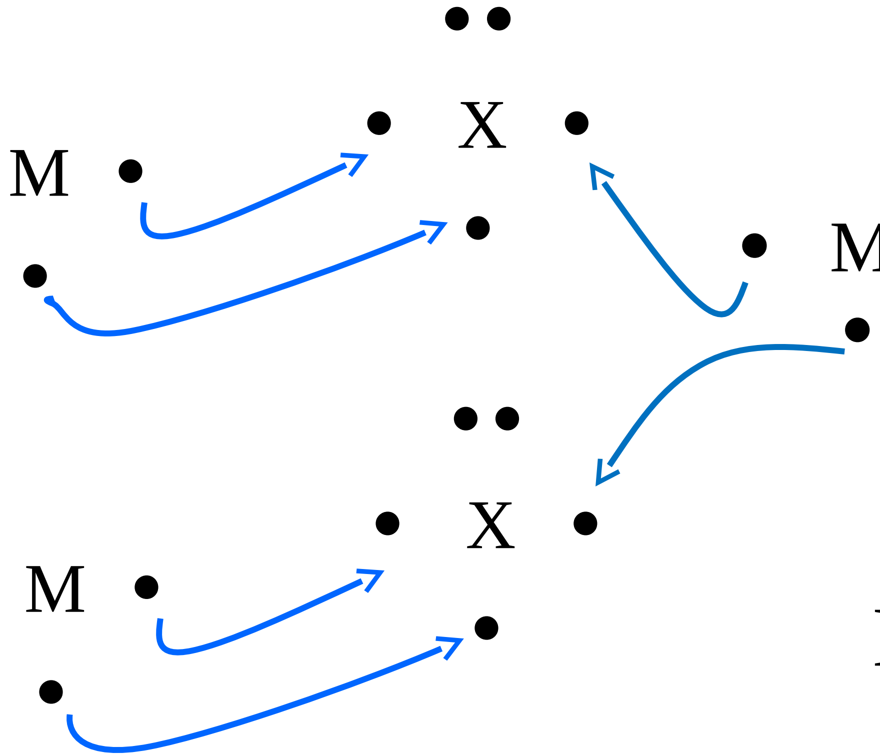
Draw the Lewis dot symbols of the elements.

Transfer all the valance electrons from the metal to the nonmetal, adding more of each atom as you go, until all electrons are lost from the metal atoms and all nonmetal atoms have 8 electrons.



$CaCl_2$

Practice—Use Lewis Symbols to Predict the Formula of an Ionic Compound Made from Reacting a Metal, M, that Has 2 Valence Electrons with a Nonmetal, X, that Has 5 Valence Electrons, Continued.



M_3X_2

Covalent Bonds

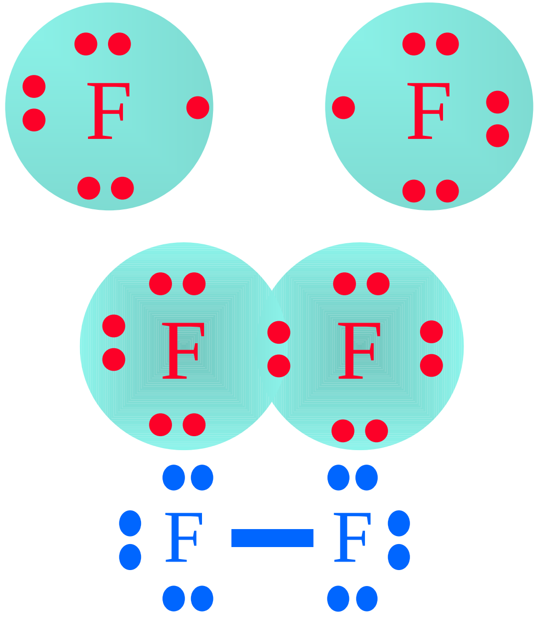
- Often found between two nonmetals.
- Typical of molecular species.
- Atoms bonded together to form molecules.
  - ✓ Strong attraction.
- **Atoms share pairs of electrons** to attain octets.
- Molecules generally weakly attracted to each other.
  - ✓ Observed physical properties of molecular substance due to these attractions.

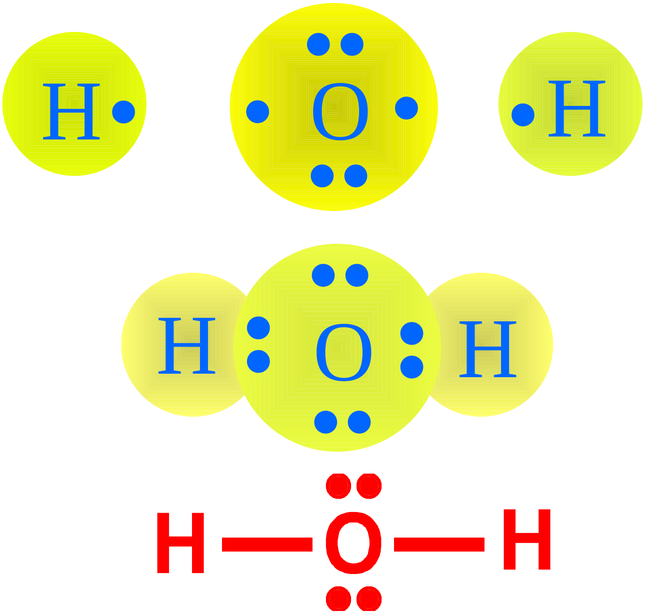
Using Lewis Atomic Structures to Predict Bonding Between Nonmetal Atoms

- Nonmetal atoms often bond to achieve an octet of valence electrons by sharing electrons.
  - ✓ Though there are many exceptions to the Octet rule.
- In Lewis theory, atoms share electrons to complete their octet.
- This may involve sharing electrons with multiple atoms or sharing multiple pairs of electrons with the same atom.

Single Covalent Bonds

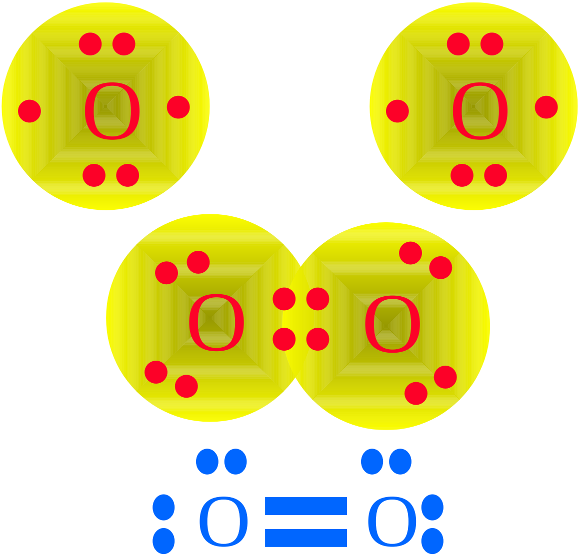
- Two atoms share one pair of electrons.
  - ✓ 2 electrons.
- One atom may have more than one single bond.





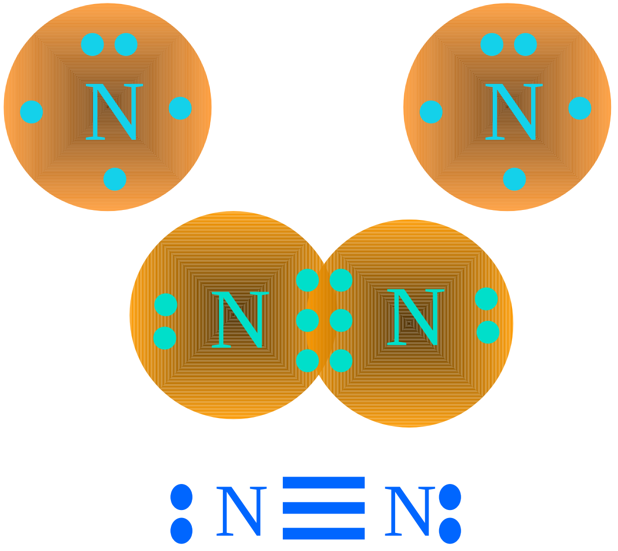
Double Covalent Bond

- Two atoms sharing two pairs of electrons.
  - ✓ 4 electrons.
- Shorter and stronger than single bond.



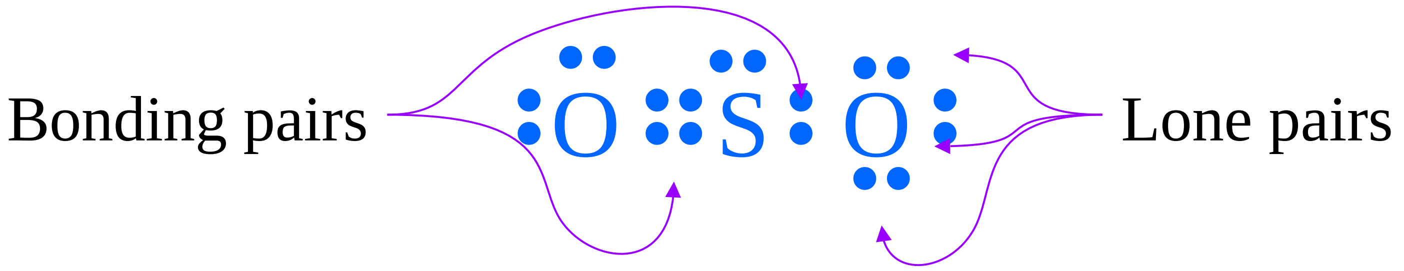
Triple Covalent Bond

- Two atoms sharing 3 pairs of electrons.
  - ✓6 electrons.
- Shorter and stronger than single or double bond.



Bonding and Lone Pair Electrons

- Electrons that are shared by atoms are called **bonding pairs**.
- Electrons that are not shared by atoms but belong to a particular atom are called **lone pairs**.
  - ✓ Also known as **nonbonding pairs**.



Multiplicity and Bond Properties

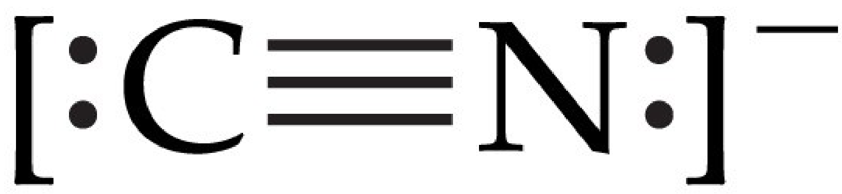
- The more electrons two atoms share, the stronger they are bonded together.
- This explains the observation that triple bonds are stronger than similar double bonds, which are stronger than single bonds.
  - ✓ C≡N is stronger than C=N, C=N is stronger than C─N.
- This explains the observation that triple bonds are shorter than similar double bonds, which are shorter than single bonds.
  - ✓ C≡N is shorter than C=N, C=N is shorter than C─N

Trends in Bond Length and Energy

Bond	**Length (pm)**	**Energy (kJ/mol)**
C-C	154	346
C=C	134	602
C≡C	120	835
C-N	147	305
C=N	128	615
C≡N	116	887
C-O	143	358
C=O	123	799

Polyatomic Ions

- The polyatomic ions are attracted to opposite ions by ionic bonds.
  - ✓Form crystal lattices.
- Atoms in the polyatomic ion are held together by covalent bonds.

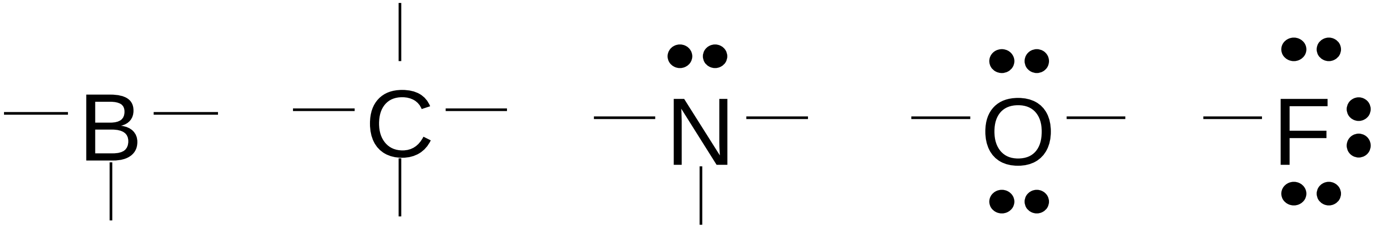


Lewis Formulas of Molecules

- Shows patterns of valence electron distribution in the molecules.
- Allows us to predict shapes of molecules.
- Allows us to predict properties of molecules and how they will interact together.

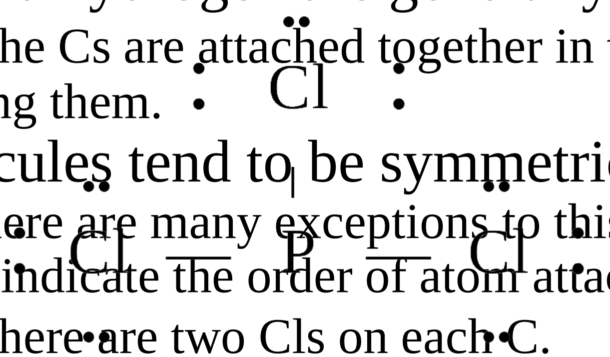
Lewis Structures

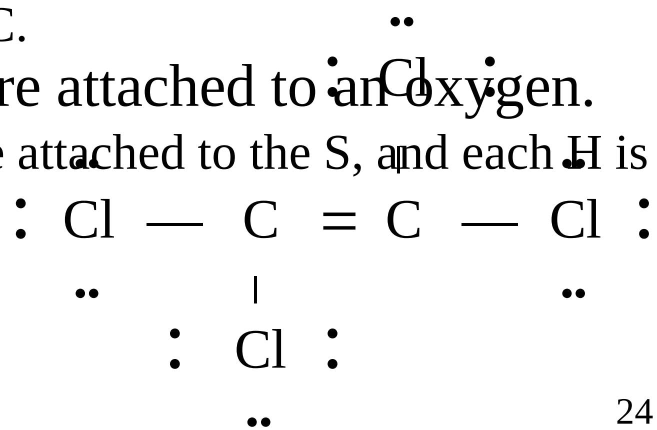
- Some common bonding patterns.
  - ✓ C = 4 bonds & 0 lone pairs.
    - ➢ 4 bonds = 4 single, or 2 double, or single + triple, or 2 single + double.
  - ✓ N = 3 bonds & 1 lone pair.
  - ✓ O = 2 bonds & 2 lone pairs.
  - ✓ H and halogen = 1 bond.
  - ✓ Be = 2 bonds & 0 lone pairs.
  - ✓ B = 3 bonds & 0 lone pairs.



Writing Lewis Structures for Covalent Molecules

1. Attach the atoms together in a skeletal structure.
   - ✓ Most metallic element is generally central.
     - ➢ In PCl_3, the P is central because it is further left on the periodic table and therefore more metallic.
   - ✓ Halogens and hydrogen are generally terminal.
     - ➢ In C_2Cl_4, the Cs are attached together in the center and the Cls are surrounding them.
   - ✓ Many molecules tend to be symmetrical.
     - ➢ Though there are many exceptions to this, chemical formulas are often written to indicate the order of atom attachment.
     - ➢ In C_2Cl_4, there are two Cls on each C.
   - ✓ In oxyacids, the acid hydrogens are attached to an oxygen.
     - ➢ In H_2SO_4, the S is central, the Os are attached to the S, and each H is attached to a different O.





Writing Lewis Structures for Covalent Molecules, Continued

2. Calculate the total number of valence electrons available for bonding.
   - ✓ Use group number of periodic table to find number of valence electrons for each atom.
   - ✓ If you have a cation, subtract 1 electron for each + charge.
   - ✓ If you have an anion, add 1 electron for each − charge.
   - ✓ In PCl_3, P has 5 e^- and each Cl has 7 e^- for a total of 26 e^-.
   - ✓ In ClO_3^-, Cl has 7 e^- and each O has 6 e^- for a total of 25 e^-.
   - ✓ Add 1 e^- for the negative charge to get a grand total of 26 e^-

Example HNO_3

1. Write skeletal structure.
   - ✓ Since this is an oxyacid, H on outside attached to one of the Os; N is central.

		O	
H	O	N	O

2. Count valence electrons.

N = 5

H = 1

$O_3 = 3 \cdot 6 = 18$

Total = 24 e^-

Example HNO_3, Continued

3. Attach atoms with pairs of electrons and subtract from the total.

```
                O
                |
H  —  O  —  N  —  O
```

$N = 5$

$H = 1$

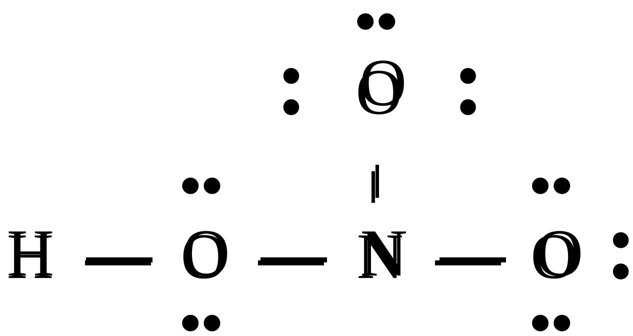
$O_3 = 3 \cdot 6 = 18$

Total = 24 e^-

	Electrons
Start	24
Used	8
Left	16

Example HNO_3, Continued

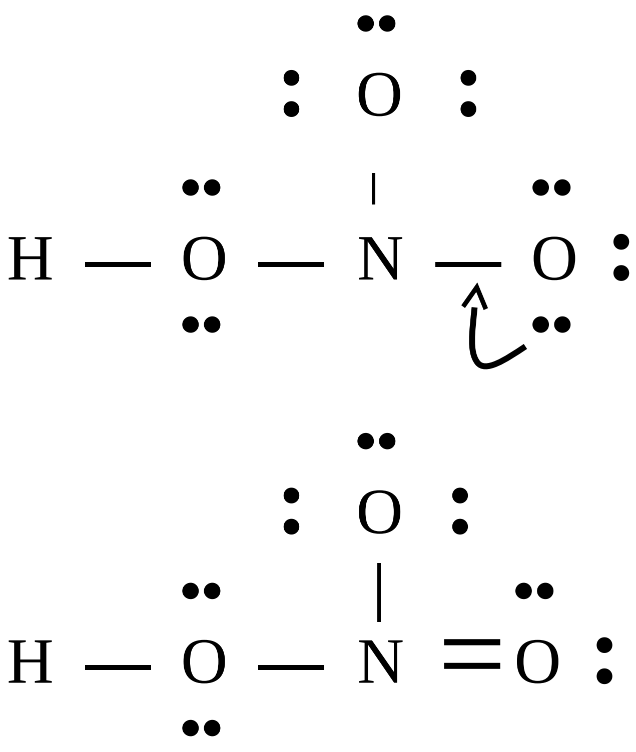
4. Complete octets, outside-in.
   - ✓ H is already complete with 2.
     - ➢ 1 bond.
   - ✓ Keep going until all atoms have an octet or you run out of electrons.



		Electrons		Electrons
N = 5				
H = 1	Start	24	Start	16
$O_3 = 3 \cdot 6 = 18$	Used	8	Used	16
Total = 24 e^-	Left	16	Left	0

Example HNO_3, Continued

5. If central atom does not have octet, bring in electron pairs from outside atoms to share.
   - ✓ Follow common bonding patterns if possible.



Example 10.4—Writing Lewis Structures for Covalent Compounds

Example 10.4:

- Write the Lewis structure of CO_2.

Example:
Write the Lewis structure of CO_2.

Information:

Given: CO_2

Find: Lewis structure

Solution Map: formula → skeletal → electron distribution → Lewis

- Apply the solution map.
  - ✓ Write skeletal structure.
    - ➢ Least metallic atom central.
    - ➢ H terminal.
    - ➢ Symmetry.

O — C — O

Example:
Write the Lewis structure of CO_2.

Information:

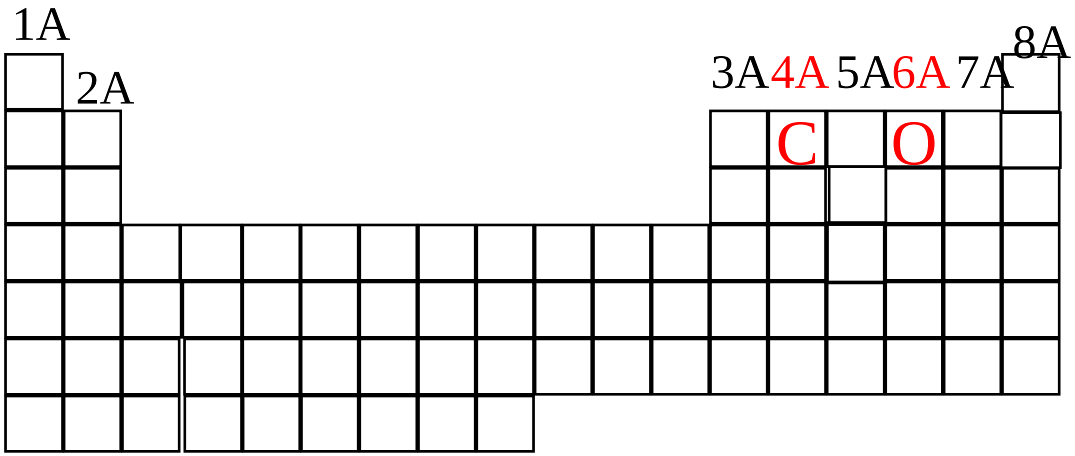
Given: CO_2

Find: Lewis structure

Solution Map: formula → skeletal → electron distribution → Lewis

- Apply the solution map.
  - ✓ Count and distribute the valence electrons.
    - ➢ Count valence electrons.

O — C — O



$$C = 4$$
$$O = 2 \cdot 6$$
$$\text{Total} \quad CO_2 = 16$$

Example:
Write the Lewis structure of CO_2.

Information:

Given: CO_2

Find: Lewis structure

Solution Map: formula → skeletal → electron distribution → Lewis

- Apply the solution map.
  - ✓ Count and distribute the valence electrons.
    - ➢ Attach atoms.

O — C — O

$$\begin{aligned} C &= 4 \\ O &= 2 \cdot 6 \\ \hline \text{Total} \quad CO_2 &= 16 \end{aligned}$$

$$\begin{aligned} \text{Start} &= 16\ e^- \\ \text{Used} &= 4\ e^- \\ \hline \text{Left} &= 12\ e^- \end{aligned}$$

Example:
Write the Lewis structure of CO_2.

Information:

Given: CO_2

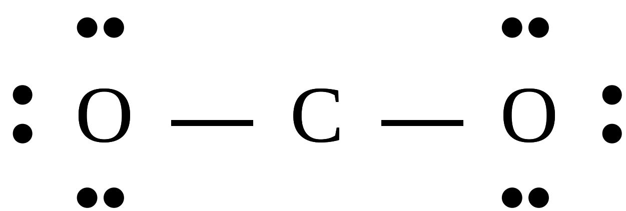
Find: Lewis structure

Solution Map: formula → skeletal → electron distribution → Lewis

- Apply the solution map.
  - ✓ Count and distribute the valence electrons.
    - ➢ Complete octets.
      - – Outside atoms first.

$$C = 4$$
$$O = 2 \cdot 6$$
$$\text{Total} \quad CO_2 = 16$$

$$\text{Start} = 16\ e^-$$
$$\text{Used} = 4\ e^-$$
$$\text{Left} = 12\ e^-$$

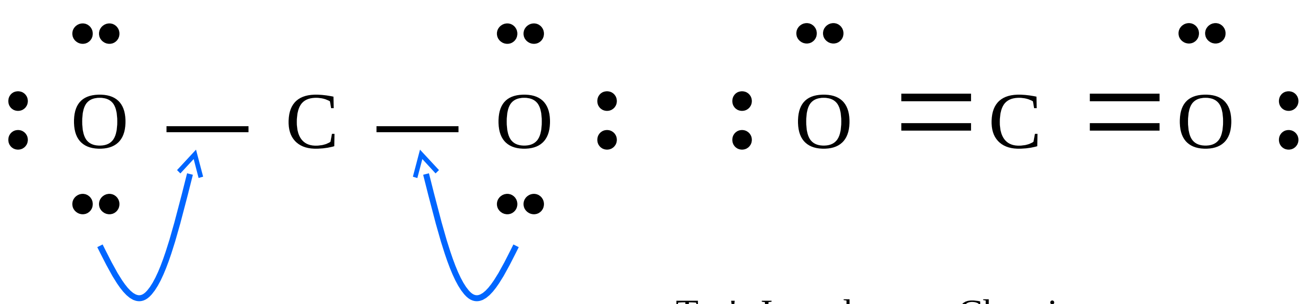


$$\text{Start} = 12\ e^-$$
$$\text{Used} = 12\ e^-$$
$$\text{Left} = 0\ e^-$$

Example:
Write the Lewis structure of CO_2.

Information:
Given: CO_2
Find: Lewis structure
Solution Map: formula → skeletal → electron distribution → Lewis

- Apply the solution map.
  - ✓ Count and distribute the valence electrons.
    - ➢ Complete octets.
      - – If not enough electrons to complete octet of central atom, bring in pairs of electrons from attached atom to make multiple bonds.



Start = 12 e^-
Used = 12 e^-
Left = 0 e^-

Example:
Write the Lewis structure of CO_2.

Information:

Given: CO_2

Find: Lewis structure

Solution Map: formula → skeletal → electron distribution → Lewis

- Check:

Start

$C = 4\ e^-$

$O = 2 \cdot 6\ e^-$

Total $CO_2 = 16\ e^-$

End

Bonding $= 4 \cdot 2\ e^-$

Lone pairs $= 4 \cdot 2\ e^-$

Total $CO_2 = 16\ e^-$

The skeletal structure is symmetrical.
All the electrons are accounted for.

Writing Lewis Structures for Polyatomic Ions

- The procedure is the same, the only difference is in counting the valence electrons.
- For polyatomic cations, take away one electron from the total for each positive charge.
- For polyatomic anions, add one electron to the total for each negative charge.

Example NO_3^-

1. Write skeletal structure.
   - ✓ N is central because it is the most metallic.

O

O N O

2. Count valence electrons.

N = 5

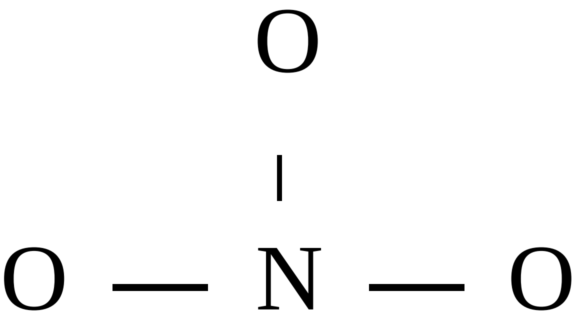
O_3 = 3·6 = 18

(-) = 1

Total = 24 e^-

Example NO_3^-, Continued

3. Attach atoms with pairs of electrons and subtract from the total.



$N = 5$

$O_3 = 3 \cdot 6 = 18$

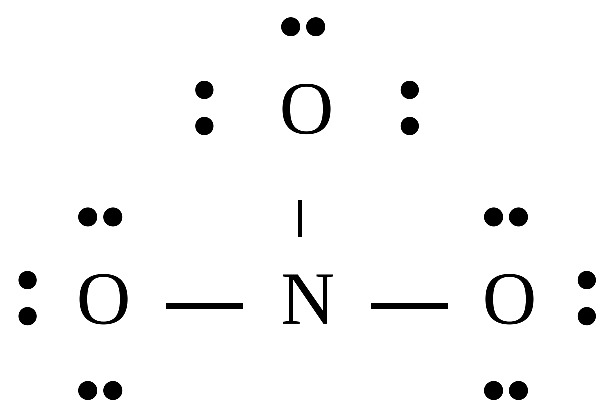
$(-) = 1$

Total = 24 e^-

	Electrons
Start	24
Used	6
Left	18

Example NO_3^-, Continued

3. Complete octets, outside-in.
   - ✓ Keep going until all atoms have an octet or you run out of electrons.



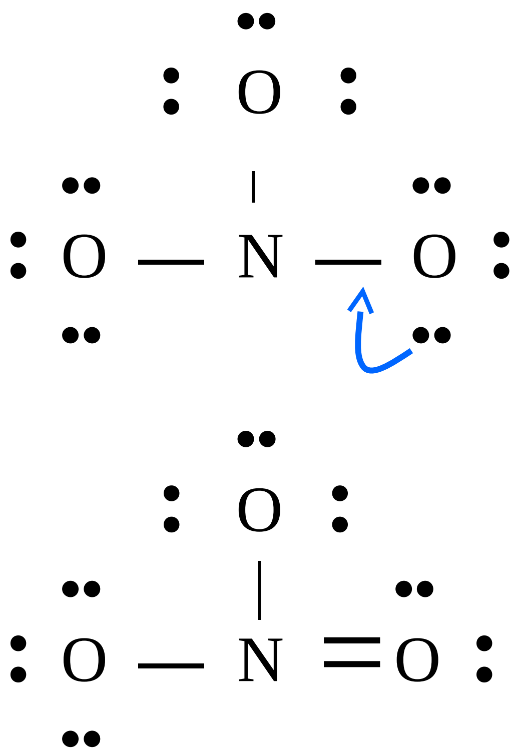
N = 5
$O_3 = 3 \cdot 6 = 18$
(-) = 1
Total = 24 e^-

	Electrons
Start	24
Used	6
Left	18

	Electrons
Start	18
Used	18
Left	0

Example NO_3^-, Continued

5. If central atom does not have octet, bring in electron pairs from outside atoms to share.
   - ✓ Follow common bonding patterns if possible.



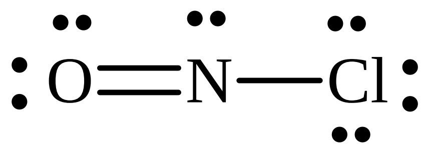
Practice—Lewis Structures

- NClO
- H_3PO_4
- H_3BO_3
- SO_3^{-2}
- NO_2^{-1}
- P_2H_4

Practice—Lewis Structures, Continued

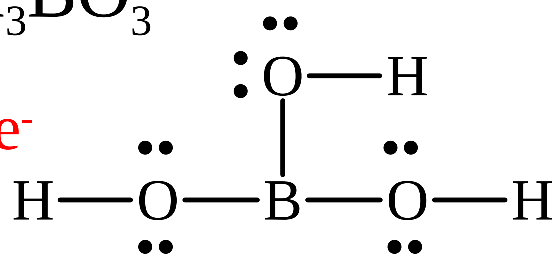
- NClO

18 e$^-$



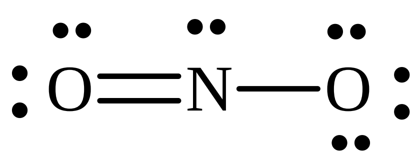
- H_3BO_3

24 e$^-$



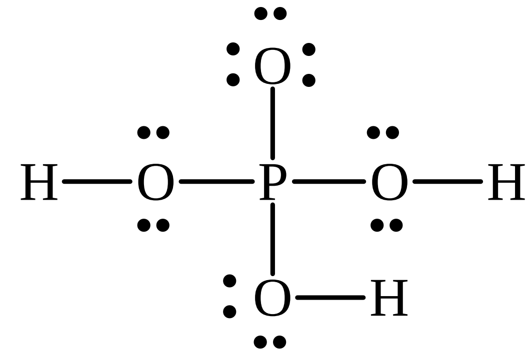
- NO_2^{-1}

18 e$^-$



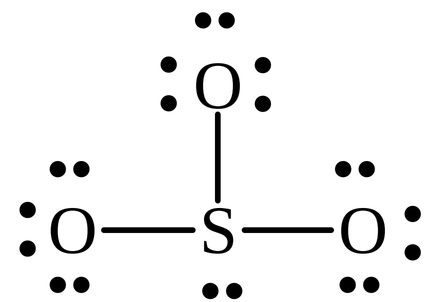
- H_3PO_4

32 e$^-$



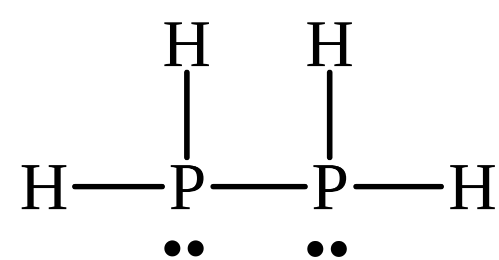
- SO_3^{-2}

26 e$^-$



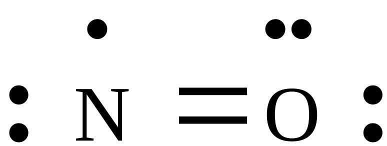
- P_2H_4

14 e$^-$



Exceptions to the Octet Rule

- H and Li, lose one electron to form cation.
  - ✓ Li now has electron configuration like He.
  - ✓ H can also share or gain one electron to have configuration like He.
- Be shares two electrons to form two single bonds.
- B shares three electrons to form three single bonds.
- Expanded octets for elements in Period 3 or below.
  - ✓ Using empty valence *d* orbitals.
- Some molecules have odd numbers of electrons.
  - ✓ NO

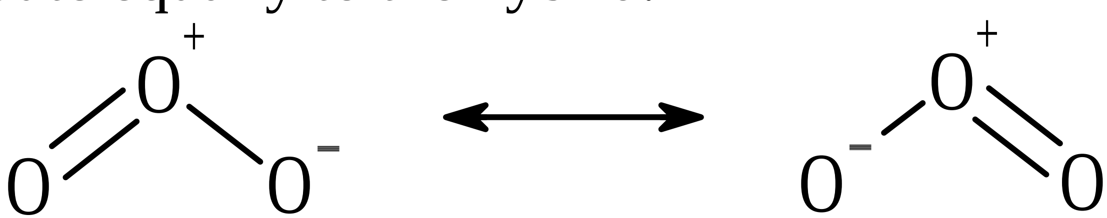


Resonance

- We can often draw more than one valid Lewis structure for a molecule or ion.
- In other words, no one Lewis structure can adequately describe the actual structure of the molecule.
- The actual molecule will have some characteristics of all the valid Lewis structures we can draw.

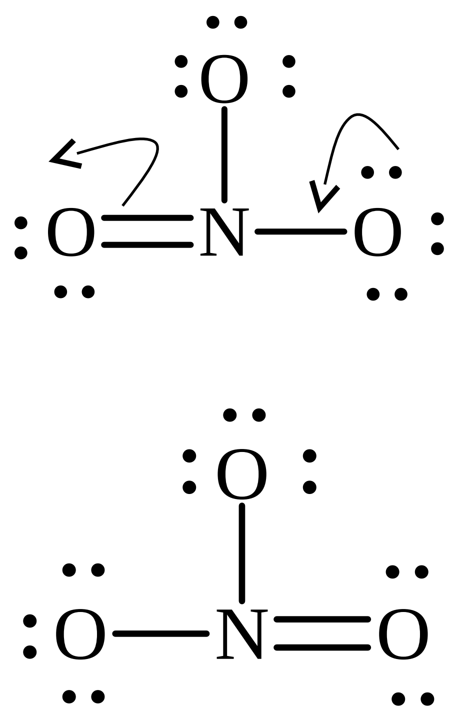
Resonance, Continued

- Lewis structures often do not accurately represent the electron distribution in a molecule.
  - ✓ Lewis structures imply that O_3 has a single (147 pm) and double (121 pm) bond, but actual bond length is between (128 pm).
- Real molecule is a *hybrid* of all possible Lewis structures.
- Resonance stabilizes the molecule.
  - ✓ Maximum stabilization comes when resonance forms contribute equally to the hybrid.



Drawing Resonance Structures

1. Draw first Lewis structure that maximizes octets.
2. Move electron pairs from outside atoms to share with central atoms.
3. If central atoms, 2nd row, only move in electrons, you can move out electron pairs from multiple bonds.



Practice—Draw Lewis Resonance Structures for CNO^{-}

(C Is Central with N and O Attached)

Practice—Draw Lewis Resonance Structures for CNO^-

(C Is Central with N and O Attached), Continued

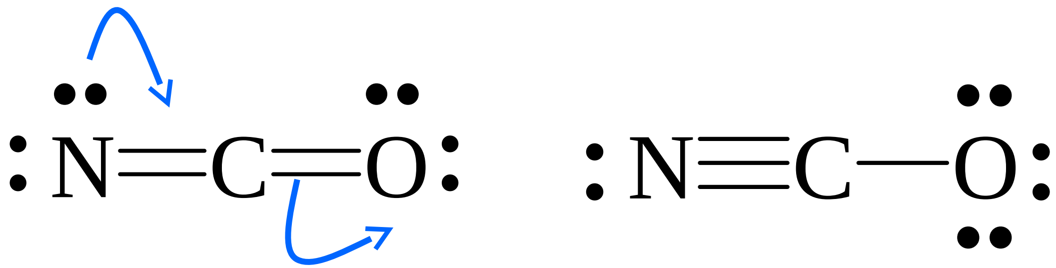
C = 4

N = 5

O = 6

(-) = 1

Total = 16 e^-



Molecular Geometry

- Molecules are three-dimensional objects.
- We often describe the shape of a molecule with terms that relate to geometric figures.
- These geometric figures have characteristic “corners” that indicate the positions of the surrounding atoms with the central atom in the center of the figure.
- The geometric figures also have characteristic angles that we call **bond angles**.

Electron Pairs

Linear

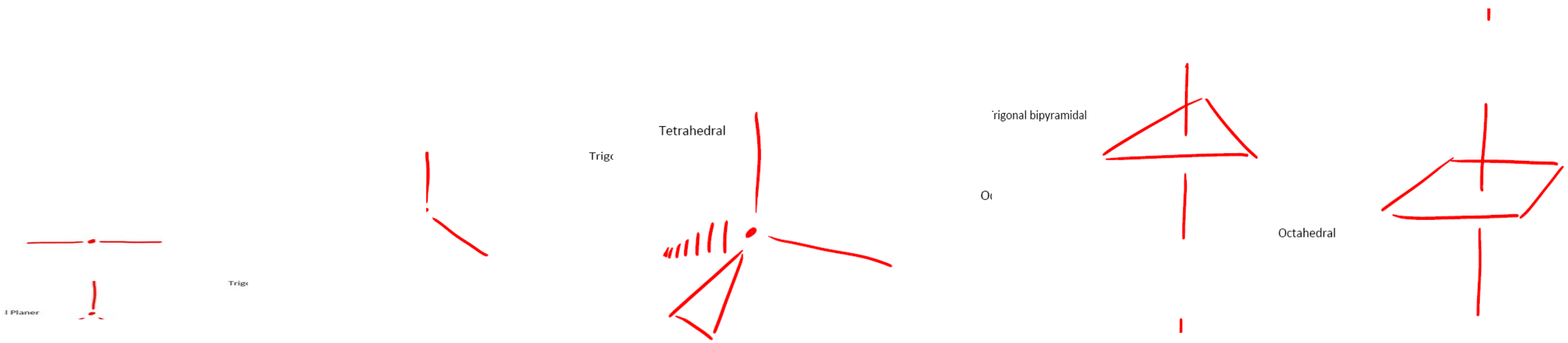
Trigonal planar

Tetrahedral

Trigonal bipyramidal

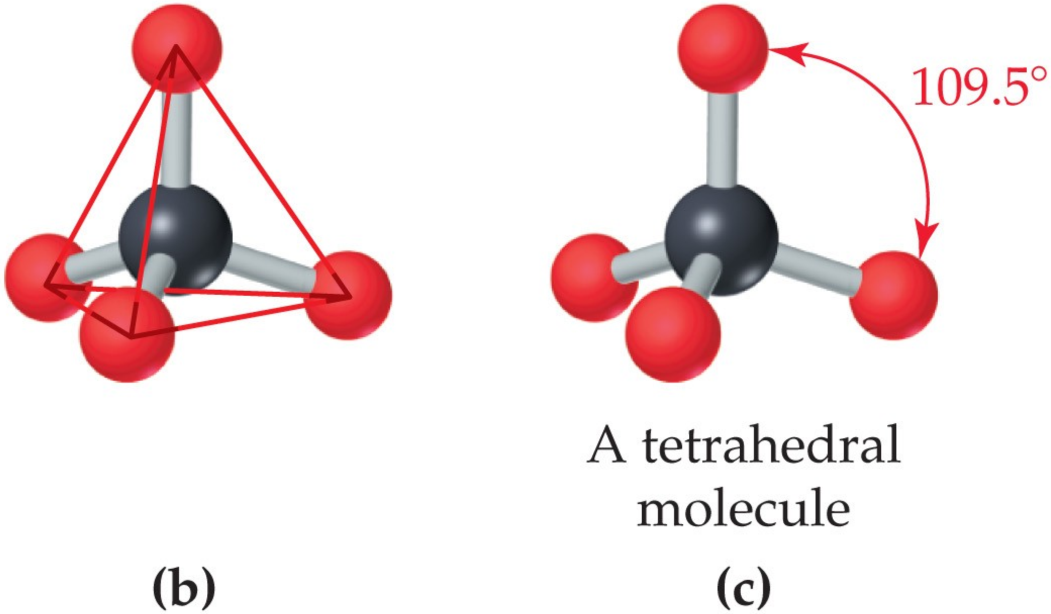
Octahedral

Practice drawing these shapes below



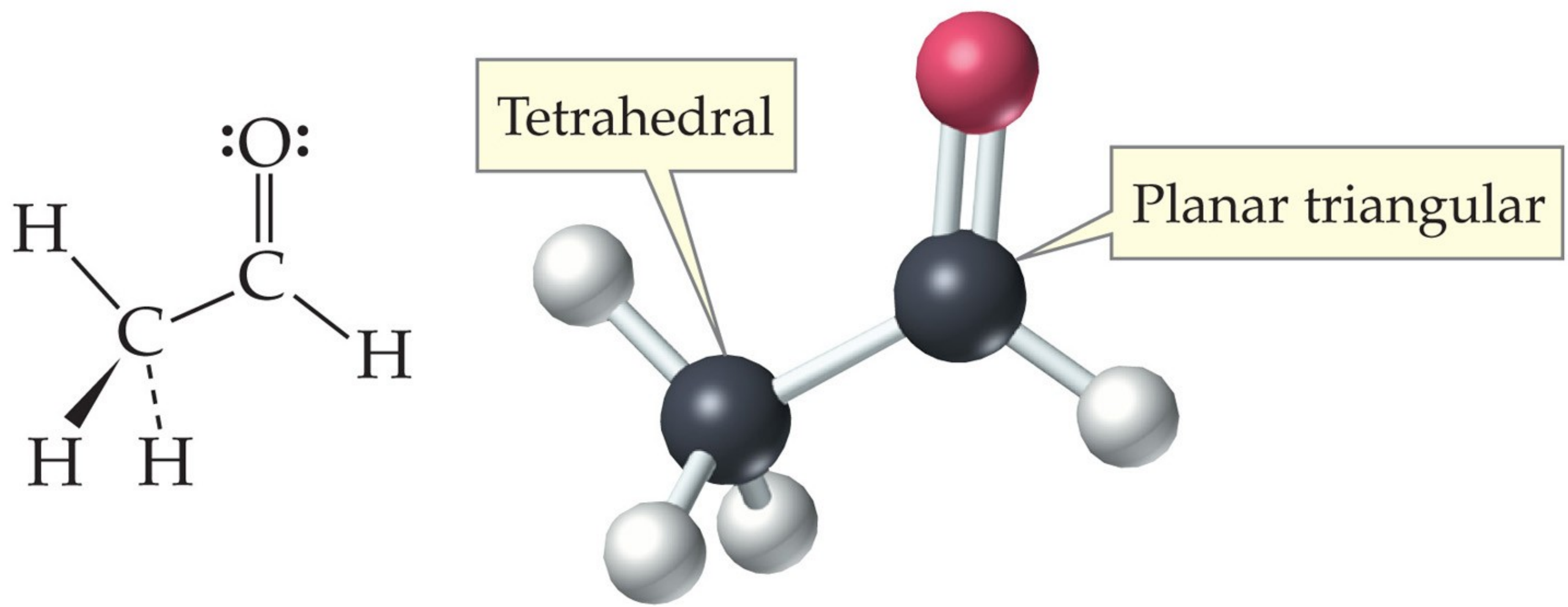
Linear **TP** **Tetra** **TBP** **Octa**

- Linear molecules have bond angles of 180°.
- Planar triangular molecules have bond angles of 120°.
- Tetrahedral molecules have bond angles of 109.5°.



Students always forget that tetrahedral is not flat! Please do not forget.

Predict the molecular geometry of acetaldehyde. One carbon on acetaldehyde is connected to four other atoms through bonds, resulting in a tetrahedral arrangement. The other carbon has only three bonds, so the arrangement around it is planar triangular. Therefore the overall structure is as shown.

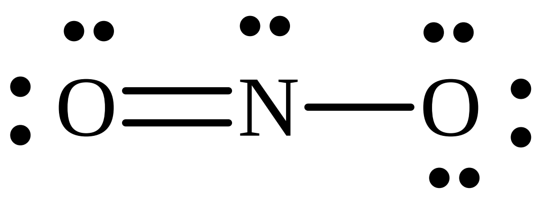


VSEPR Theory

- Electron groups around the central atom will be most stable when they are as far apart as possible. We call this **valence shell electron pair repulsion theory**.
  - ✓ Since electrons are negatively charged, they should be most stable when they are separated.
- The resulting geometric arrangement will allow us to predict the shapes and bond angles in the molecule.

Electron Groups

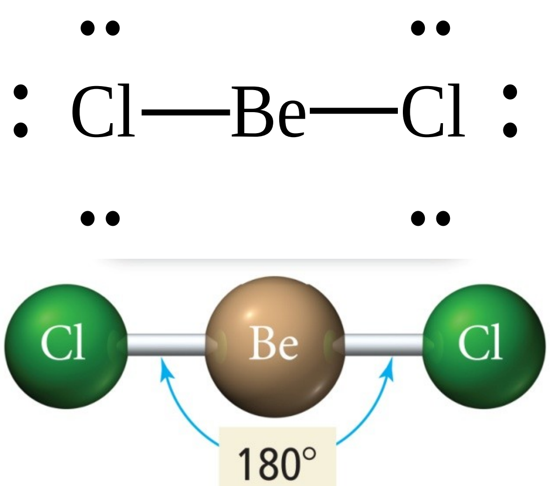
- The Lewis structure predicts the arrangement of valence electrons around the central atom(s).
- Each lone pair of electrons constitutes one electron group on a central atom.
- Each bond constitutes one electron group on a central atom.
  - ✓ Regardless of whether it is single, double, or triple.

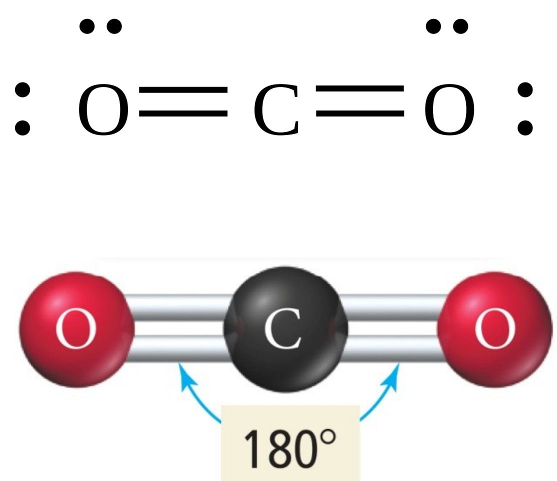


There are 3 electron groups on N.
1 lone pair.
1 single bond.
1 double bond.

Linear Geometry

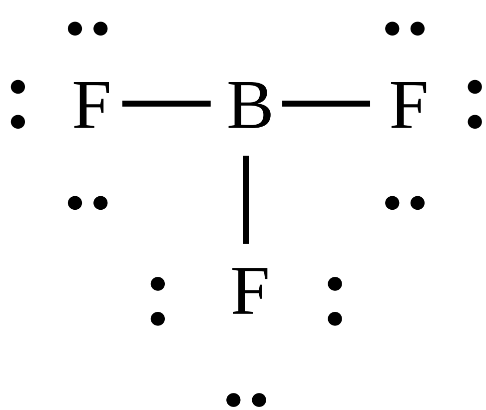
- When there are two electron groups around the central atom, they will occupy positions opposite each other around the central atom.
- This results in the molecule taking a **linear geometry**.
- The bond angle is 180°.

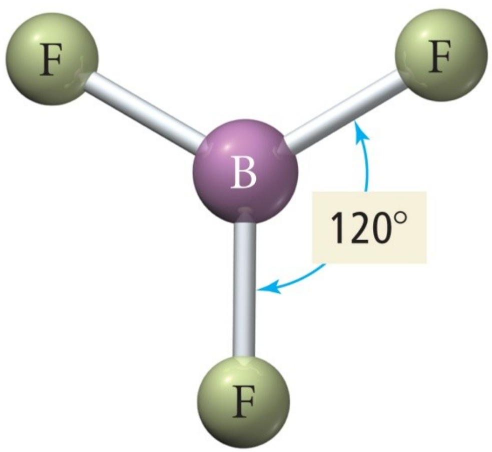




Trigonal Geometry

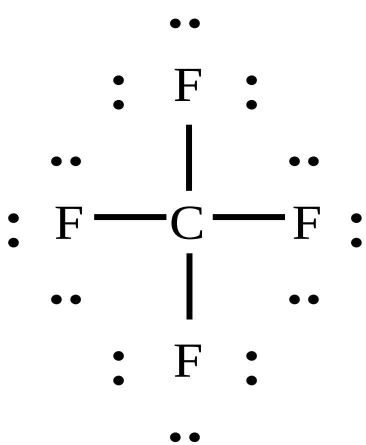
- When there are three electron groups around the central atom, they will occupy positions in the shape of a triangle around the central atom.
- This results in the molecule taking a **trigonal planar geometry**.
- The bond angle is 120°.

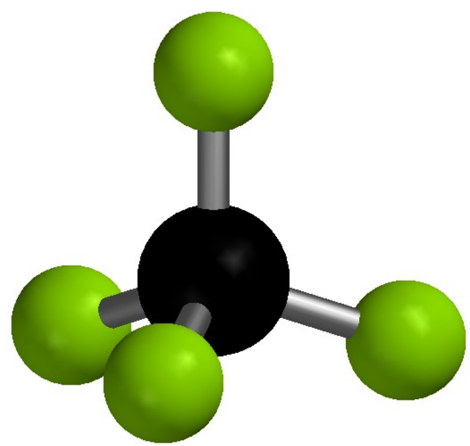
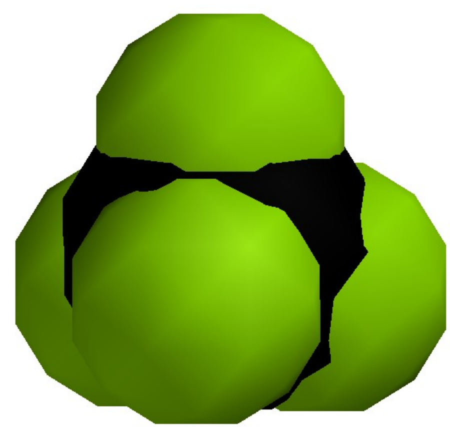




Tetrahedral Geometry

- When there are four electron groups around the central atom, they will occupy positions in the shape of a tetrahedron around the central atom.
- This results in the molecule taking a **tetrahedral geometry**.
- The bond angle is 109.5°.

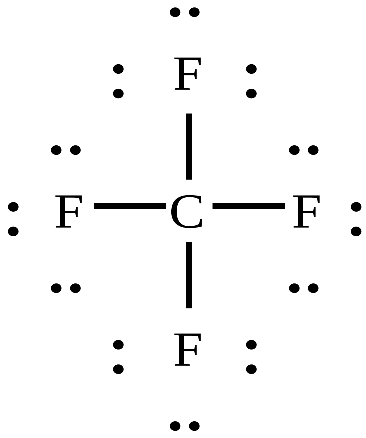


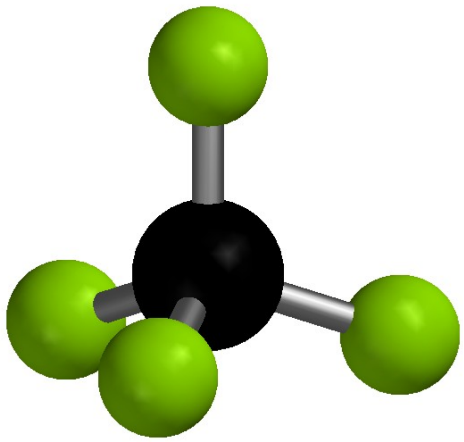



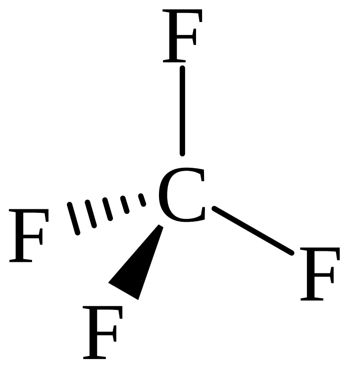
Sketching a Molecule

- Because molecules are three-dimensional objects, our drawings should indicate their three-dimensional quality
- By convention:
  - ✓ A filled wedge indicates that the attached atom is coming out of the paper toward you.
  - ✓ A dashed wedge indicates that the attached atom is going behind the paper away from you.

Sketching a Molecule, Continued





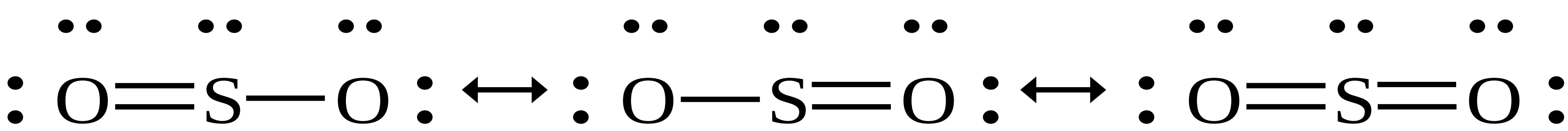


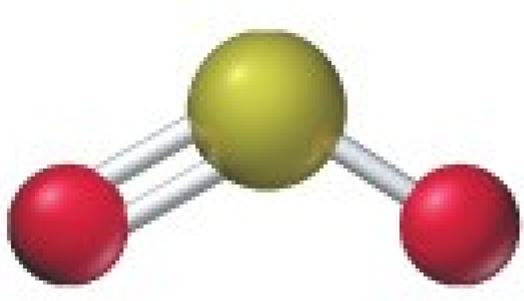
Derivative Shapes

- The molecule's shape will be one of basic molecular geometries if all the electron groups are bonds and all the bonds are equivalent.
- Molecules with lone pairs or different kinds of surrounding atoms will have distorted bond angles and different bond lengths, but the shape will be a derivative of one of the basic shapes.

Derivative of Trigonal Geometry

- When there are three electron groups around the central atom, and one of them is a lone pair, the resulting shape of the molecule is called a **bent shape**.
- The bond angle is < 120°.

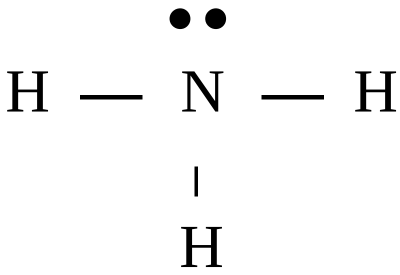


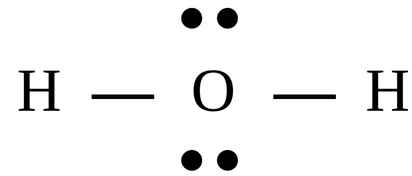


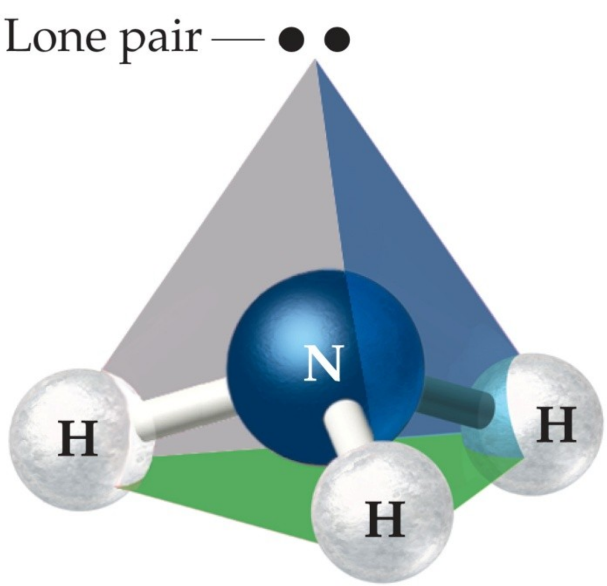
Derivatives of Tetrahedral Geometry

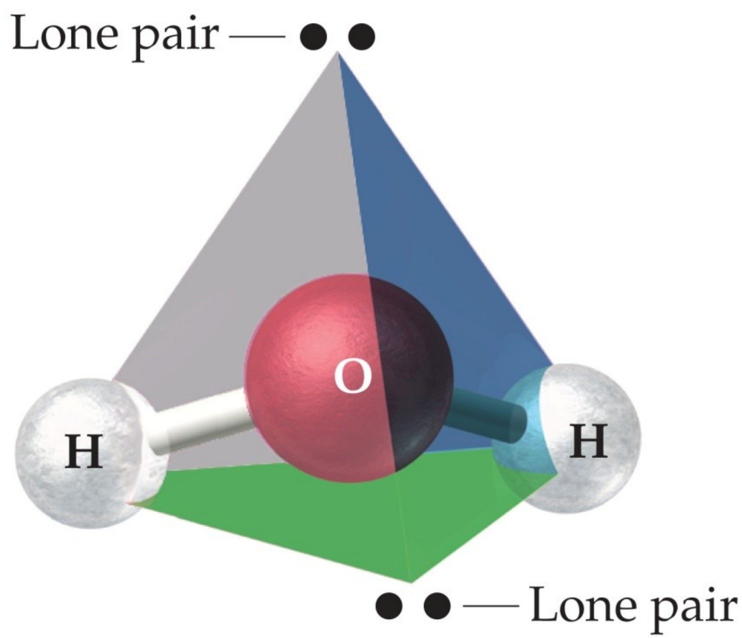
- When there are four electron groups around the central atom, and one is a lone pair, the result is called a **Trigional pyramidal shape**.
  - ✓ Because it is a triangular-base pyramid with the central atom at the apex.
- When there are four electron groups around the central atom, and two are lone pairs, the result is called a **tetrahedral–bent shape**.
  - ✓ It is planar.
  - ✓ It looks similar to the trigonal planar bent shape, except the angles are smaller.
- For both shapes, the bond angle is < 109.5°.

Tetrahedral Derivatives



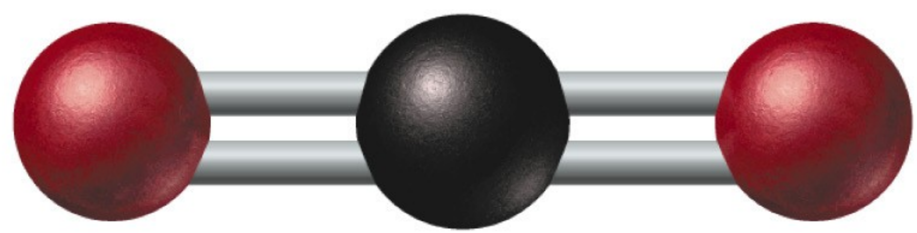






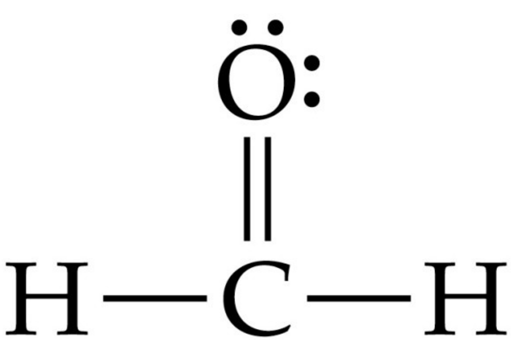
Molecular Geometry: Linear

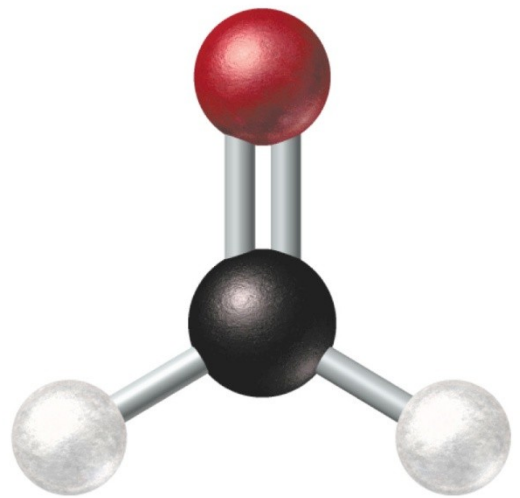
- Electron groups rround central atom = 2.
- Bonding groups = 2.
- Lone pairs = 0.
- Electron geometry = linear.
- Angle between electron groups = 180°.



Molecular Geometry: Trigonal Planar

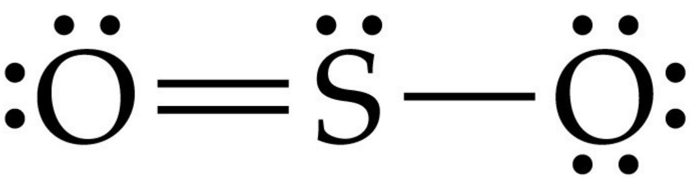
- Electron groups around central atom = 3.
- Bonding groups = 3.
- Lone pairs = 0.
- Electron geometry = trigonal planar.
- Angle between electron groups = 120°.

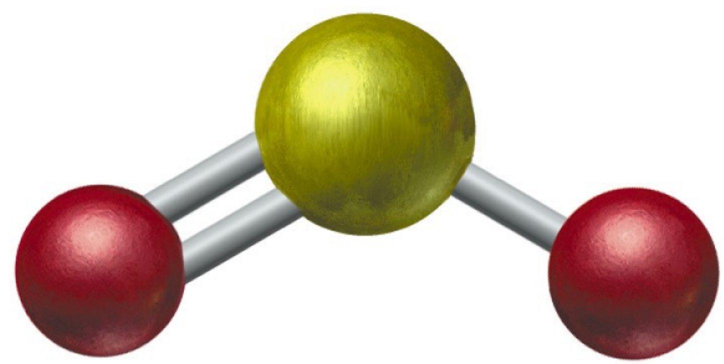




Molecular Geometry: Bent

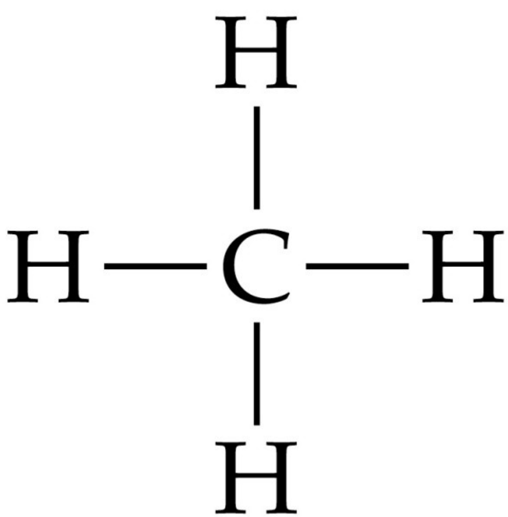
- Electron groups around central atom = 3.
- Bonding groups = 2.
- Lone pairs = 1.
- Electron geometry = trigonal planar.
- Angle between electron groups = 120°.

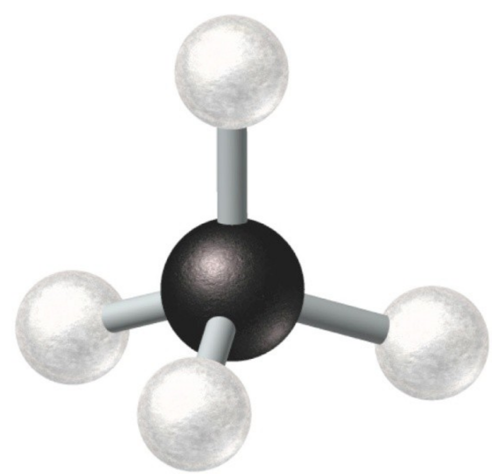




Molecular Geometry: Tetrahedral

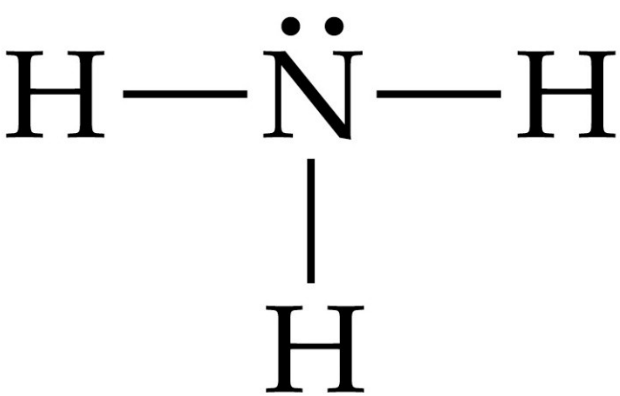
- Electron groups around central atom = 4.
- Bonding groups = 4.
- Lone pairs = 0.
- Electron geometry = tetrahedral.
- Angle between electron groups = 109.5°.

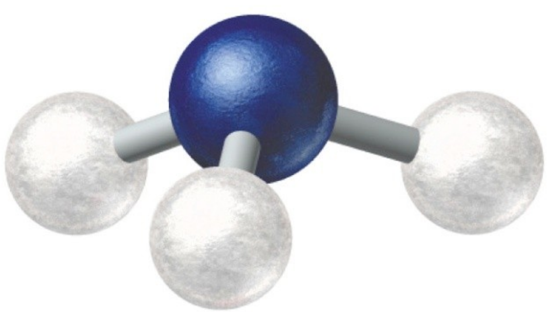




Molecular Geometry: Trigonal Pyramid

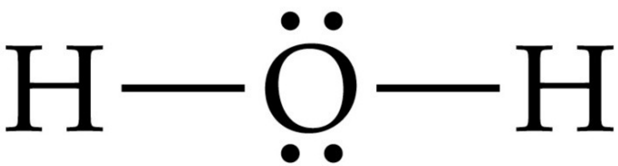
- Electron groups around central atom = 4.
- Bonding groups = 3.
- Lone pairs = 1.
- Electron geometry = tetrahedral.
- Angle between electron groups = 109.5°.

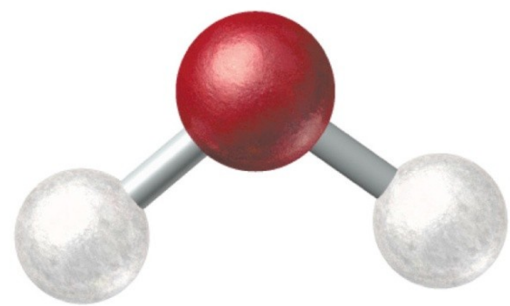




Molecular Geometry: Bent

- Electron groups around central atom = 4.
- Bonding groups = 2.
- Lone pairs = 2.
- Electron geometry = tetrahedral.
- Angle between electron groups = 109.5°.





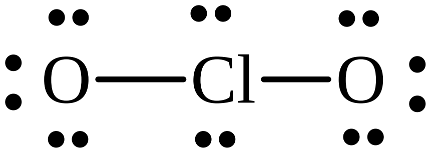
Electron areas	**Molecular Geometry**	**Bond angles**	**Example**
2 bonding. 0 lone pairs	Linear	180°	C_2H_2, SCN^{-1}
3 bonding, 0 lone pairs	Trigonal planar	120°	C_2H_4, BF_3, CO_3^{-2}
2 bonding, 1 lone pair	Trigonal planar bent	< 120°	CH_2O, $SnCl_2$
4 bonding, 0 lones pairs	Tetrahedral	109.5°	CH_4, $POCl_3$
3 bonding, 1 lone pair	Pyramidal (trigonal)	~107°	$:NH_3$
2 bonding, 2 lone pairs	Tetrahedral bent	~105°	H_2O, OF_2

Predicting the Molecular Shapes Around Central Atoms

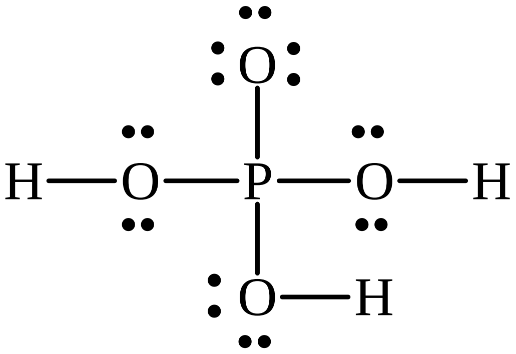
1. Draw the Lewis structure.
2. Determine the number of electron groups around the central atom.
3. Classify each electron group as bonding or lone pair, and count each type.
   - ✓ Remember: Multiple bonds count as one group.
4. Use the previous slide's table to determine the shape and bond angles.

Practice—Predict the Molecular Geometry Around the Central Atom

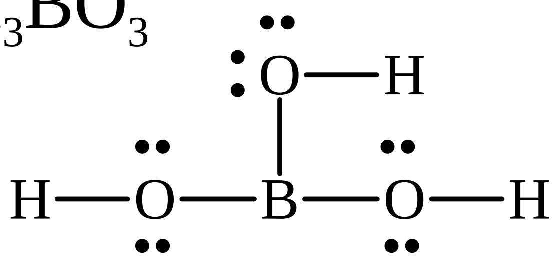
- ClO_2^-



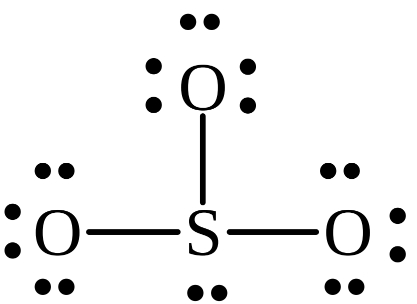
- H_3PO_4



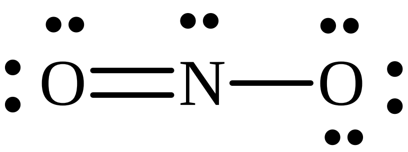
- H_3BO_3



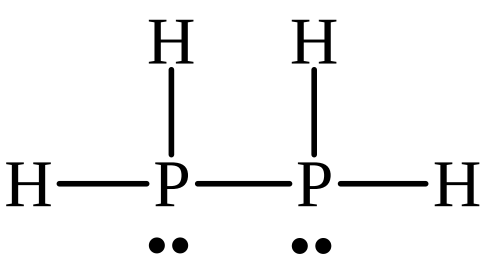
- SO_3^{2-}



- NO_2^{-1}



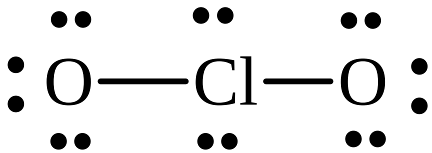
- P_2H_4



Practice—Predict the Molecular Geometry Around the Central Atom, Continued

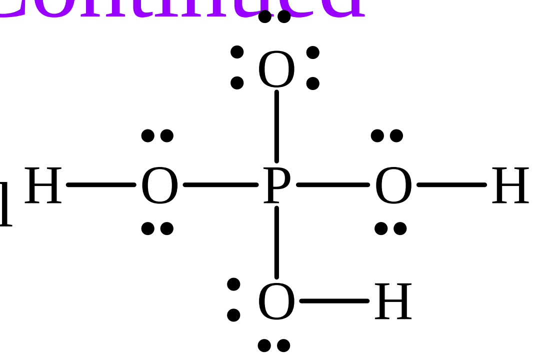
- ClO_2^-

Tetr.
bent



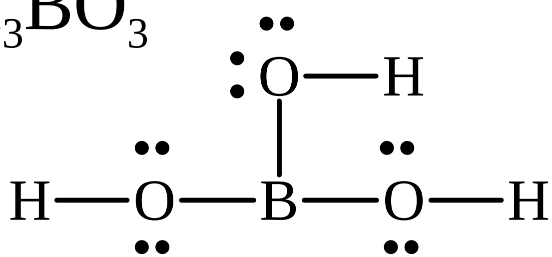
- H_3PO_4

Tetrahedral



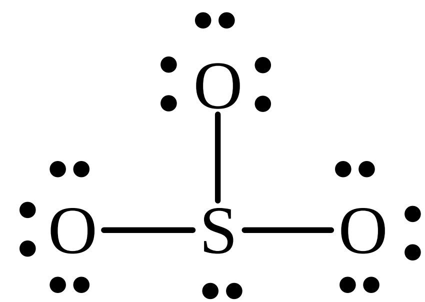
- H_3BO_3

Trigonal



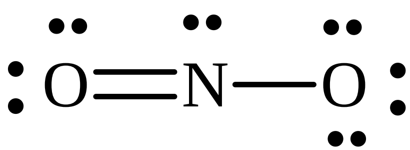
- SO_3^{2-}

Trig.
pyramidal



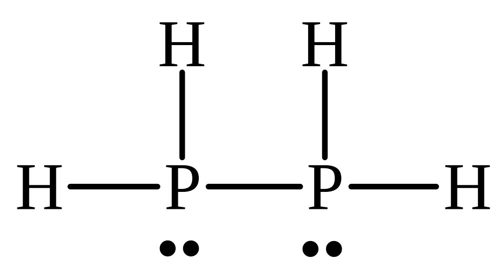
- NO_2^{-1}

Trig.
bent



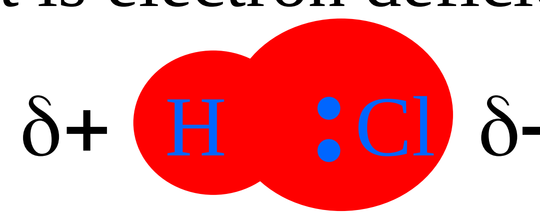
- P_2H_4

Trig.
pyramidal



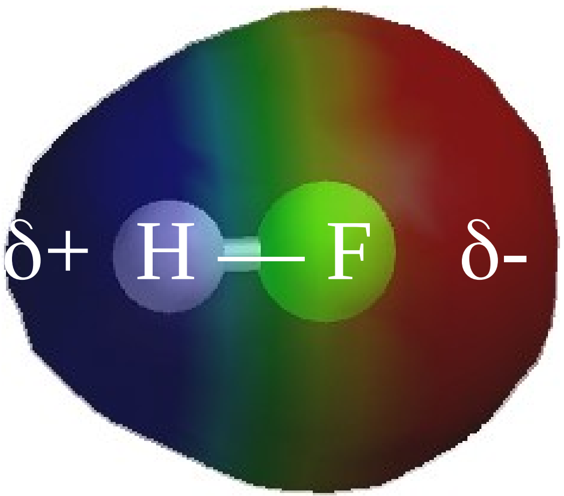
Bond Polarity

- Bonding between unlike atoms results in unequal sharing of the electrons.
  - ✓One atom pulls the electrons in the bond closer to its side.
  - ✓One end of the bond has larger electron density than the other.
- The result is **bond polarity**.
  - ✓The end with the larger electron density gets a partial negative charge and the end that is electron deficient gets a partial positive charge.



Electronegativity

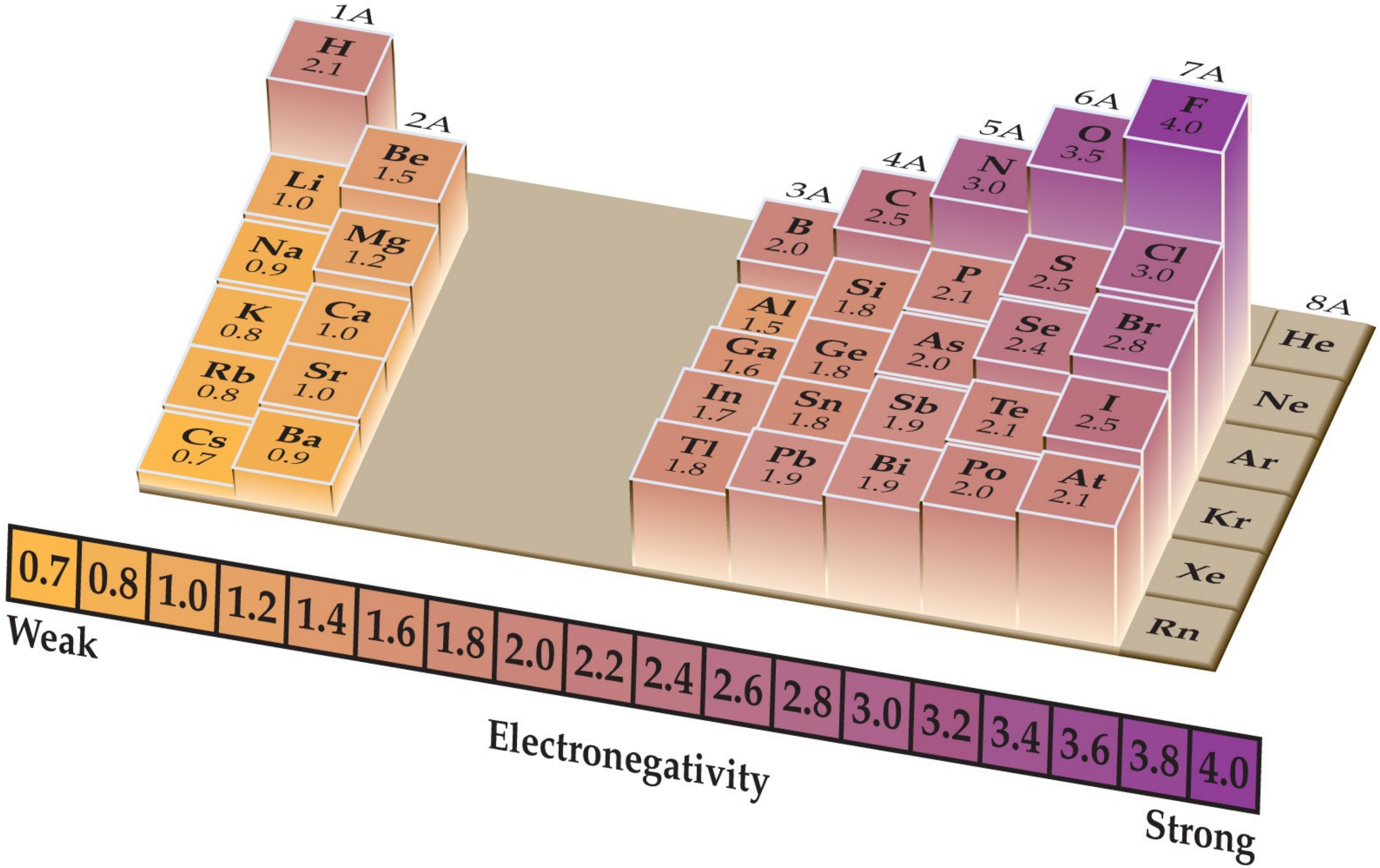
- Measure of the pull an atom has on bonding electrons.
- Increases across the period (left to right).
- Decreases down the group (top to bottom).
- The larger the difference in electronegativity, the more polar the bond.
  - ✓ Negative end toward more electronegative atom.



Electronegativity, Continued

2.1																	
1.0	1.5											2.0	2.5	3.0	3.5	4.0	
0.9	1.2											1.5	1.8	2.1	2.5	3.0	
0.8	1.0	1.3	1.5	1.6	1.6	1.5	1.8	1.8	1.8	1.9	1.6	1.6	1.8	2.0	2.4	2.8	
0.8	1.0	1.2	1.4	1.6	1.8	1.9	2.2	2.2	2.2	1.9	1.7	1.7	1.8	1.9	2.1	2.5	
0.7	0.9	1.1	1.3	1.5	1.7	1.9	2.2	2.2	2.2	2.4	1.9	1.8	1.8	1.9	2.0	2.2	
0.7	0.9	1.1															

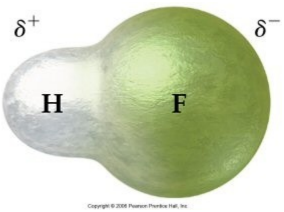
Electronegativity, Continued



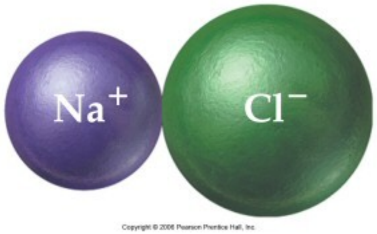
Bond Polarity

3.0-3.0
= 0.0

4.0-2.1
= 1.9



3.0-0.9
= 2.1



Covalent **Ionic**

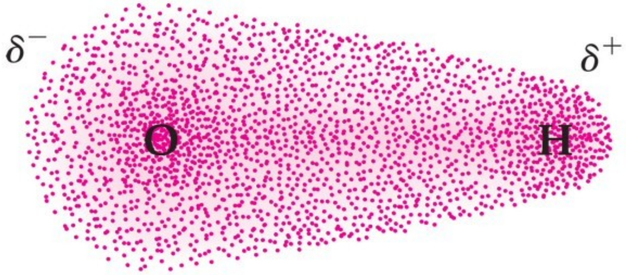
Pure Polar

0 0.4 2.0 4.0

Electronegativity difference

Dipole Moments

- A dipole is a material with positively and negatively charged ends.
- Polar bonds or molecules have one end slightly positive, δ^+, and the other slightly negative, δ^-.
  - ✓ Not "full" charges, come from nonsymmetrical electron distribution.
- Dipole moment, μ, is a measure of the size of the polarity.
  - ✓ Measured in debyes, D.



For Each of the Following Bonds, Determine Whether the Bond Is Ionic or Covalent. If Covalent, Determine if It Is Polar or Pure. If Polar, Indicate the Direction of the Dipole.

- Pb-O
- P-S
- Mg-Cl
- H-O

For Each of the Following Bonds, Determine Whether the Bond Is Ionic or Covalent. If Covalent, Determine if It Is Polar or Pure. If Polar, Indicate the Direction of the Dipole, Continued.

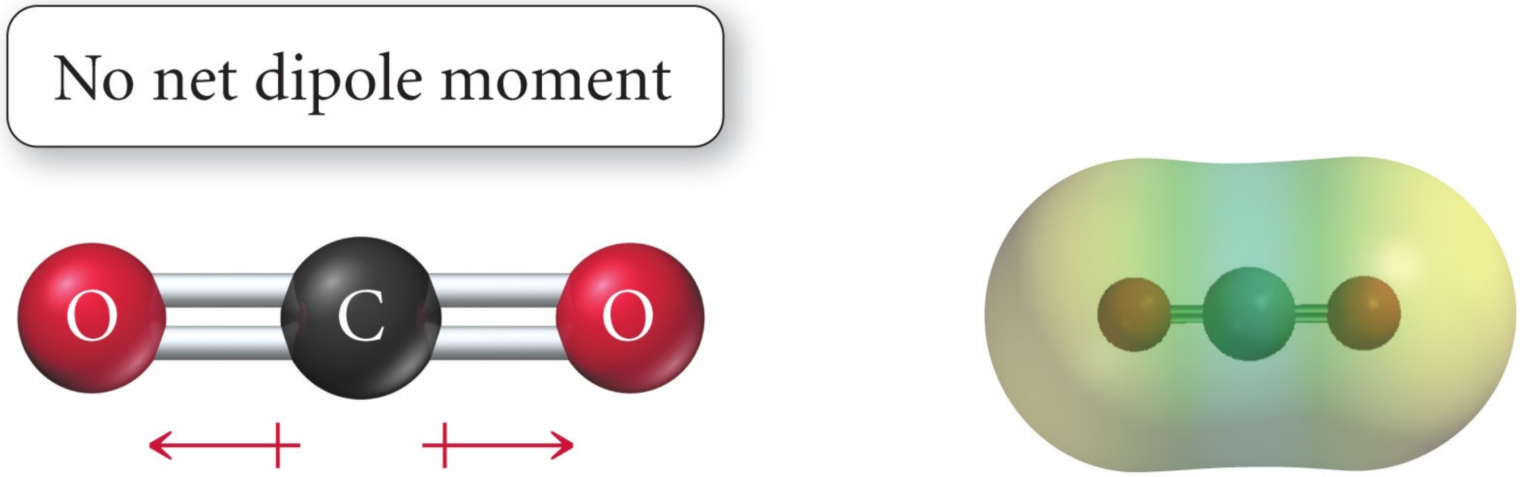
- Pb-O (3.5 - 1.9) = 1.6 ∴ polar covalent.
- P-S (2.5 - 2.1) = 0.4 ∴ pure covalent.
- Mg-Cl (3.0 - 1.2) = 1.8 ∴ ionic.

- H-O (3.5 - 2.1) = 1.4 ∴ polar covalent.

Polarity of Molecules

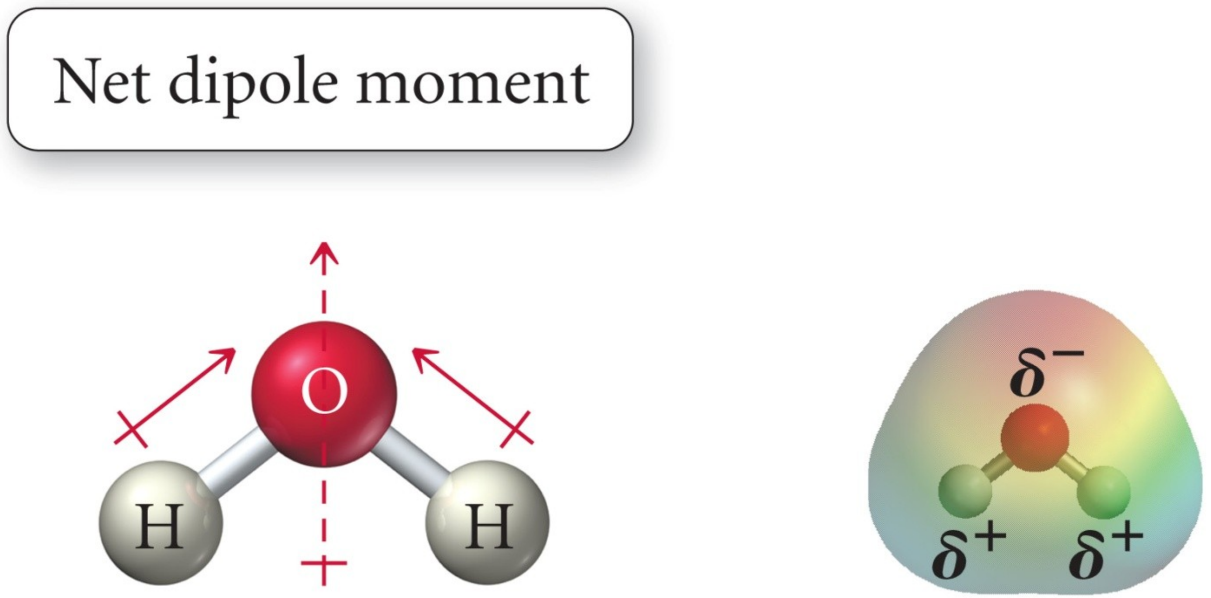
- In order for a molecule to be polar it must:
  1. Have polar bonds.
     - Electronegativity difference—theory.
     - Bond dipole moments—measured.
  2. Have an unsymmetrical shape.
     - Vector addition.
- Polarity effects the intermolecular forces of attraction.

Molecule Polarity



The O—C bond is polar. The bonding electrons are pulled equally toward both O ends of the molecule. The net result is a nonpolar molecule.

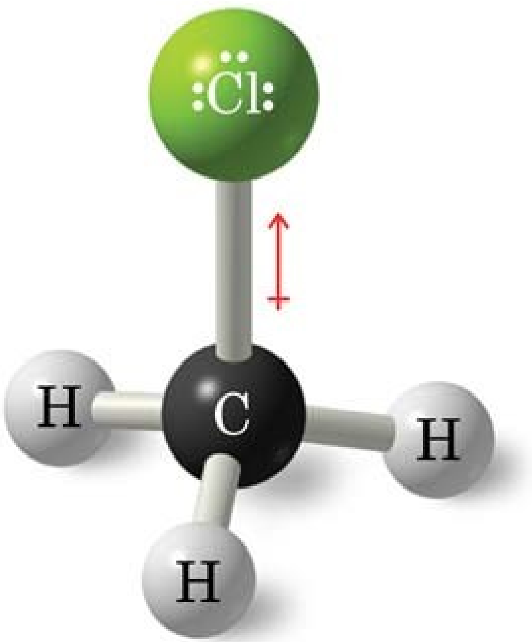
Molecule Polarity, Continued



The H—O bond is polar. Both sets of bonding electrons are pulled toward the O end of the molecule. The net result is a polar molecule.

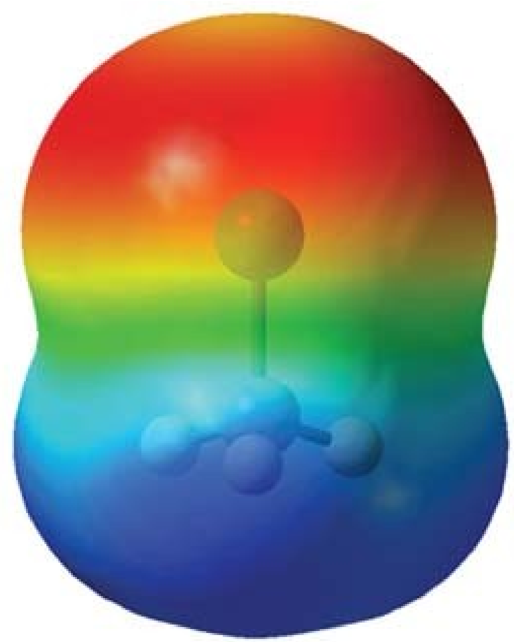
Polar Covalent Bonds and Electronegativity

This is a picture (EPM) of a chloromethane. The red area is a high concentration of electrons, and blue means low concentration of electrons.



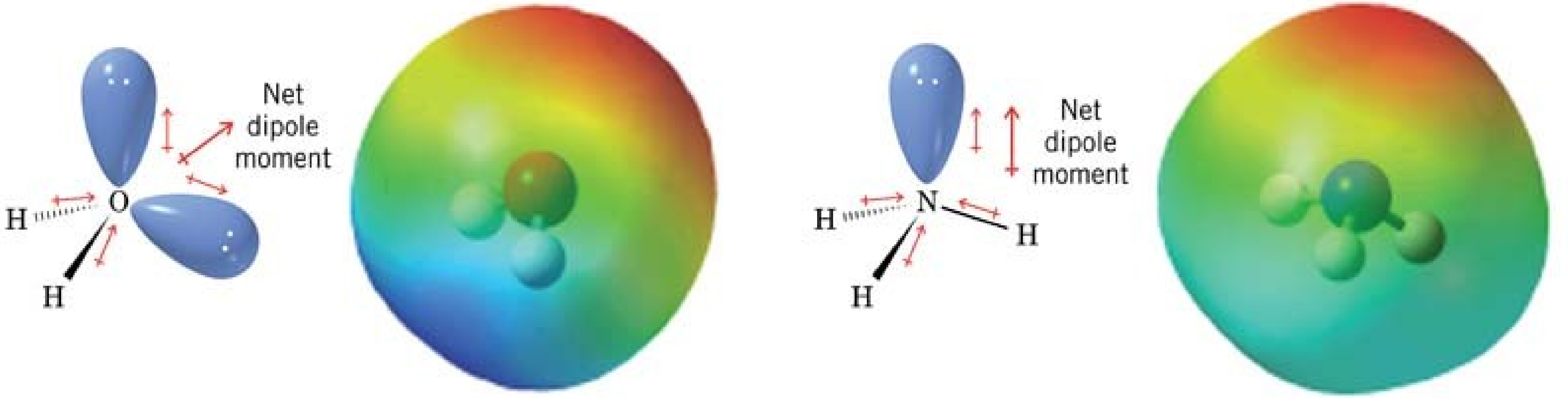
F

$\mu = 1.87$ D



Polar Covalent Bonds and Electronegativity

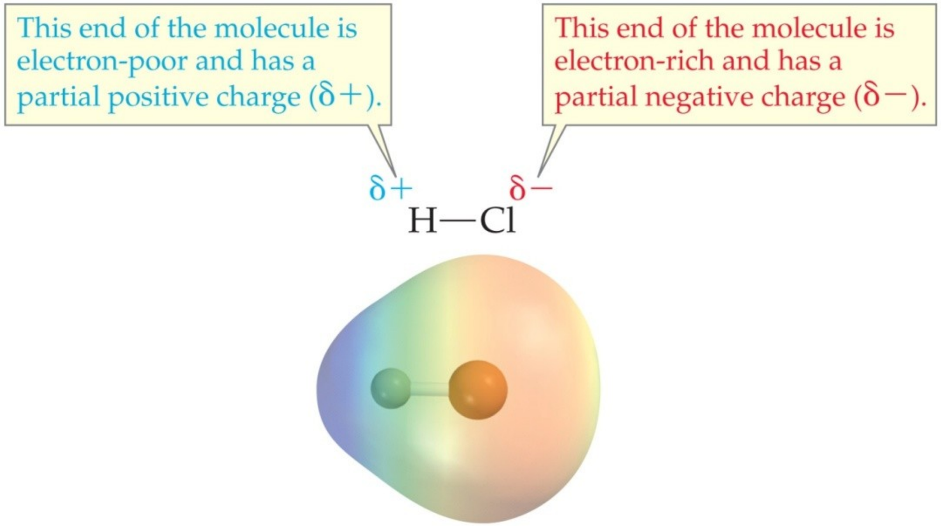
- What causes the electrons to congregate onto different atoms?



Water

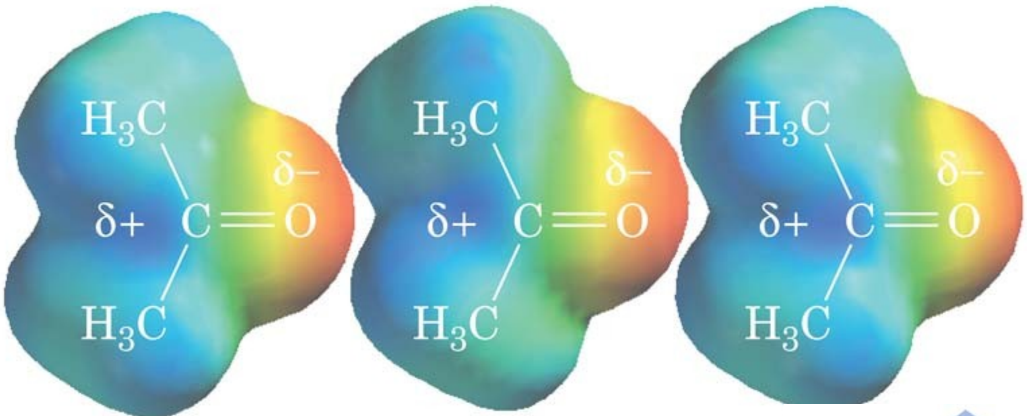
Ammonia

Atoms like chlorine want electrons more then hydrogen. As chlorine pulls electrons away from they hydrogen atom, a charge builds up on the chlorine. Instead of sharing electrons-equally-the molecule begins to look, like an ionic compound. The chemistry term is the molecule has a dipole moment and is said to be polar; just like earth has a North and South Pole.



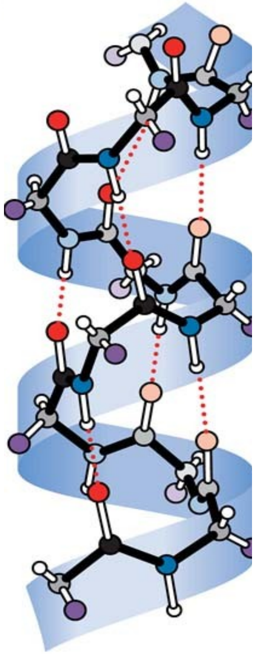
- Molecules can attract each other. This may not seem like much, but this is how DNA is held together. This helps scientists

Ultimately

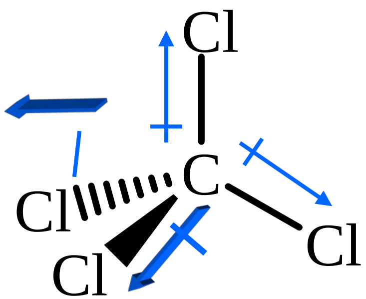


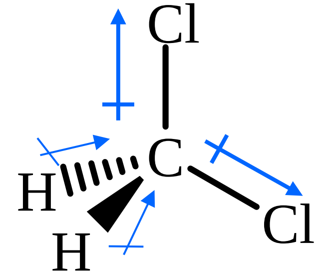
explain the differences in melting point, boiling point, as well as other physical properties.

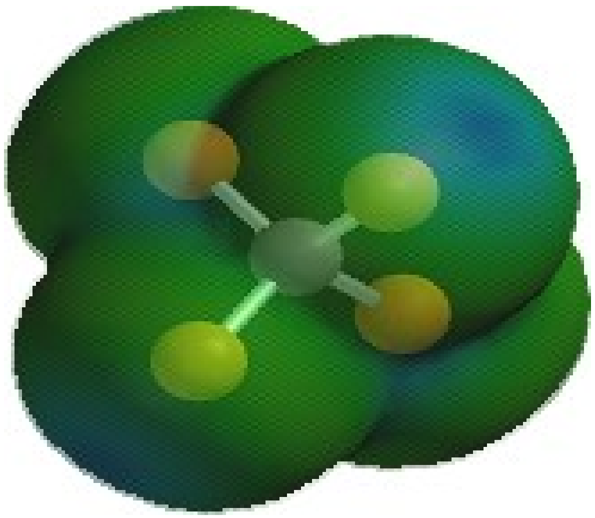
So how do we predict this behavior?



Dipole Moment

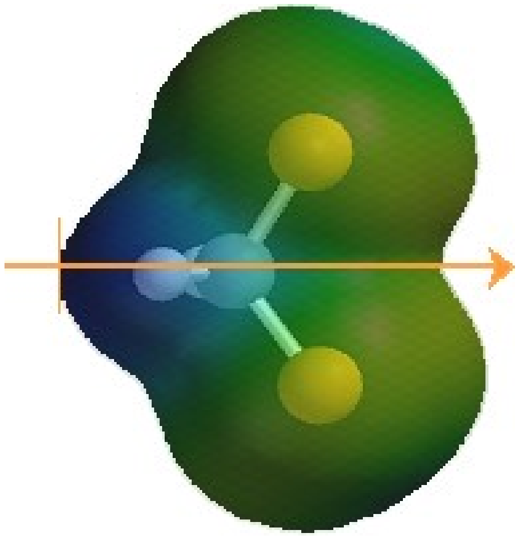






CCl_4

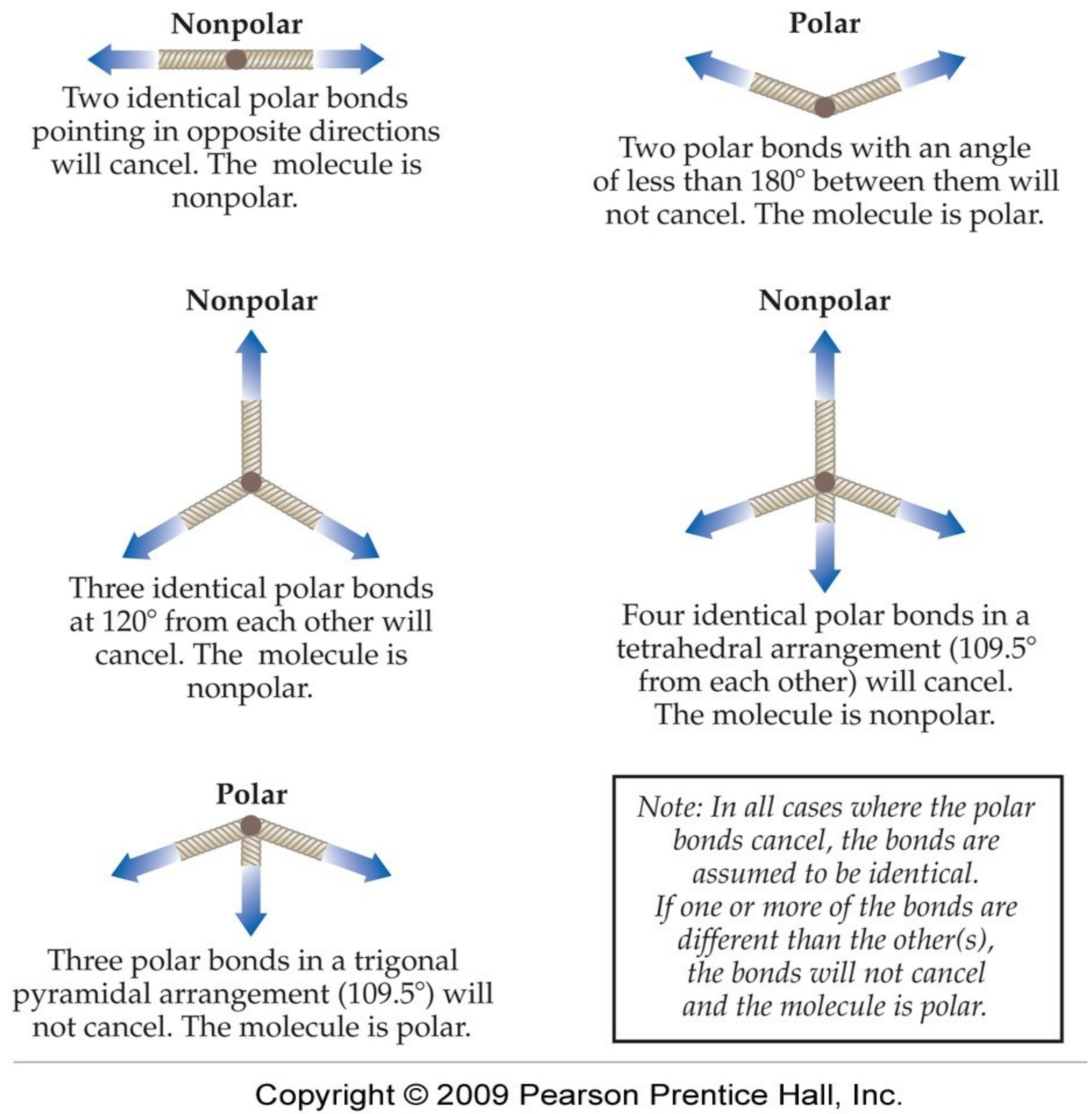
$\mu = 0.0$ D



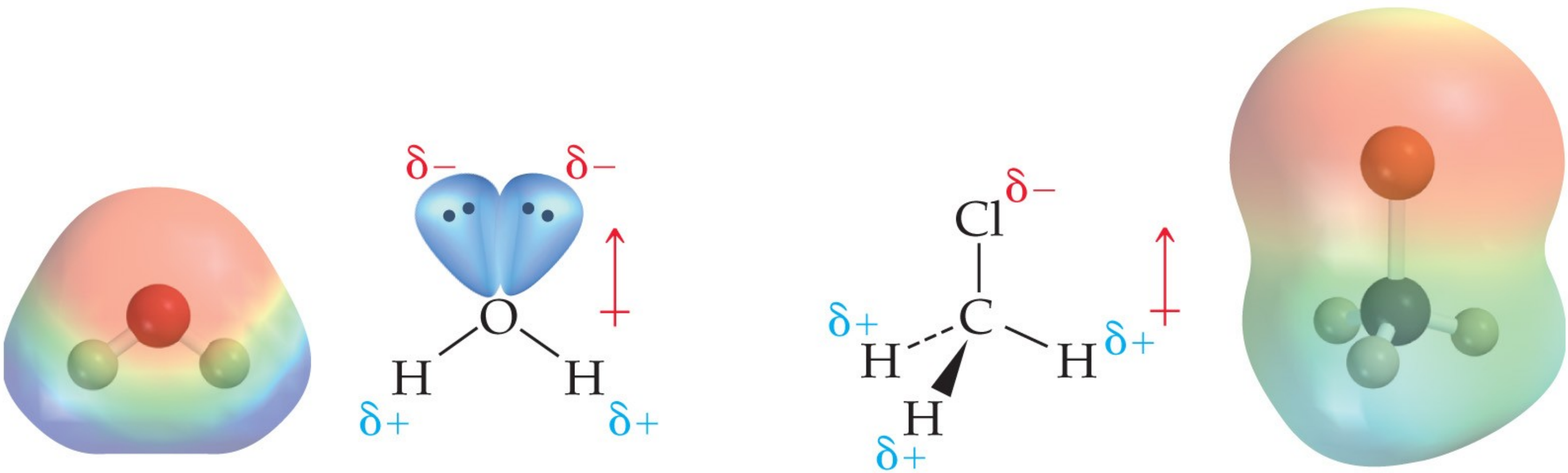
CH_2Cl_2

$\mu = 2.0$ D

Adding Dipole Moments

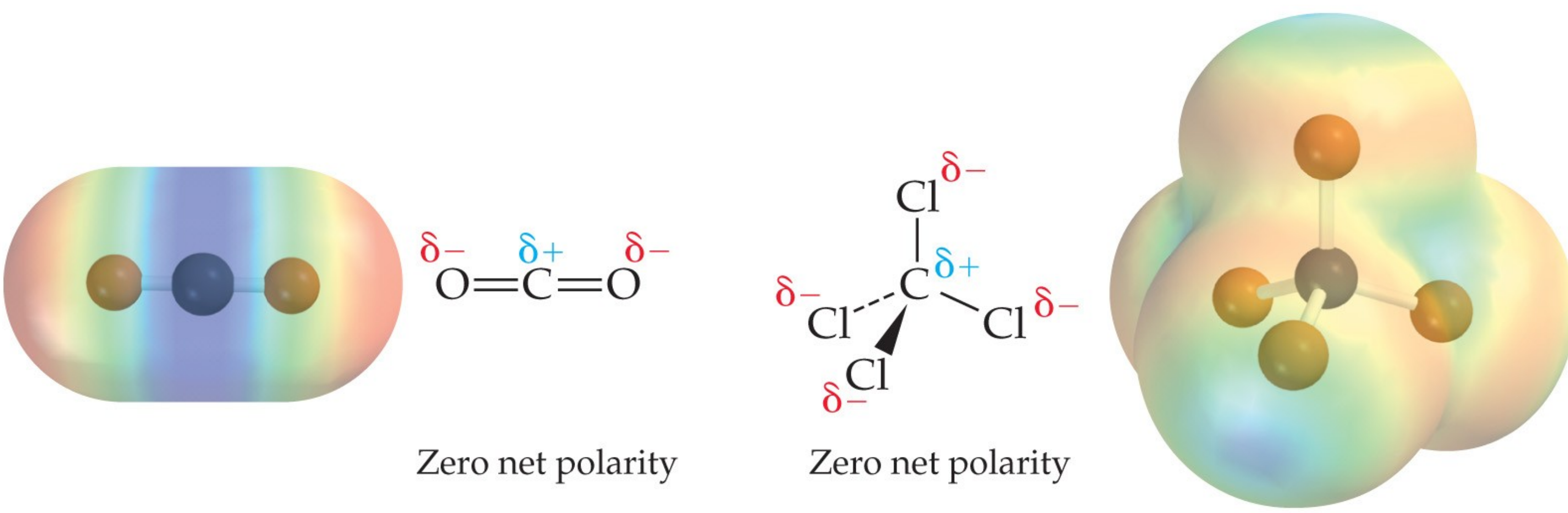


- Dipoles or polarity can be represented by an arrow pointing to the negative end of the molecule with a cross at the positive end resembling a + sign.



Water, H_2O Chloromethane, CH_3Cl

- Just because a molecule has polar covalent bonds does not mean that the molecule is polar overall.
- Carbon dioxide and tetrachloromethane molecules have no net polarity because their symmetrical shapes cause the individual bond polarities to cancel each other out.
- Notice: these do Not have a North/South Pole



Example 10.11—Determining if a Molecule Is Polar

Example 10.11:

- Determine if NH_3 is polar.

Example:
Determine if NH_3 is polar.

- Write down the given quantity and its units.

Given: NH_3

Example:
Determine if NH_3 is polar.

Information:

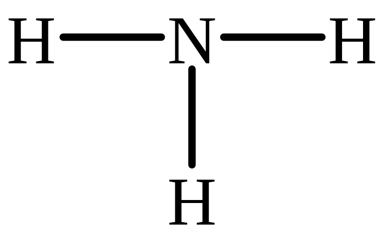
Given: NH_3

Find: If polar

Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Draw the Lewis structure.
    - ➢ Write skeletal structure.



Example:
Determine if NH_3 is polar.

Information:

Given: NH_3

Find: If polar

Solution Map:

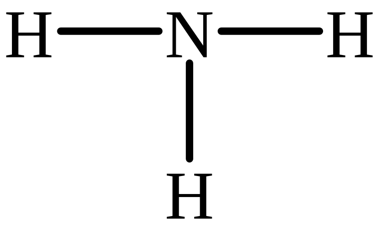
formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Draw the Lewis structure.
    - ➢ Count valence electrons.

N = 5

H = 3 · 1

Total NH_3 = 8



Example:
Determine if NH_3 is polar.

Information:

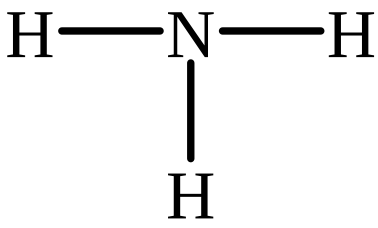
Given: NH_3

Find: If polar

Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Draw the Lewis structure.
    - ➢ Attach atoms.



N = 5
H = 3 · 1
Total NH_3 = 8

Start 8 e^-
Used 6 e^-
Left 2 e^-

Example:
Determine if NH_3 is polar.

Information:

Given: NH_3

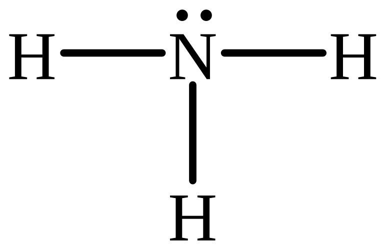
Find: If polar

Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Draw the Lewis structure.
    - ➢ Complete octets.

N = 5
H = 3 · 1
Total NH_3 = 8



Start 2 e^-
Used 2 e^-
Left 0 e^-

Example:

Determine if NH_3 is polar.

Information:

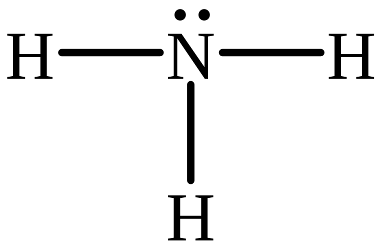
Given: NH_3

Find: If polar

Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Determine if bonds are polar.



Electronegativity

N = 3.0

H = 2.1

3.0 – 2.1 = 0.9

∴ polar covalent

Example:
Determine if NH_3 is polar.

Information:

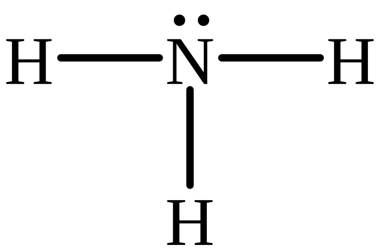
Given: NH_3

Find: If polar

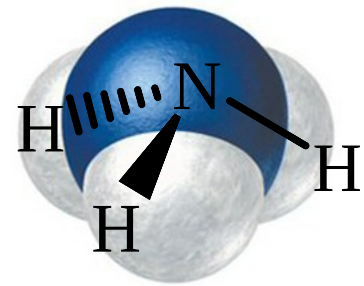
Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Determine shape of molecule.



4 areas of electrons around N;
3 bonding areas
1 lone pair



Shape = trigonal pyramid

Example:
Determine if NH_3 is polar.

Information:

Given: NH_3

Find: If polar

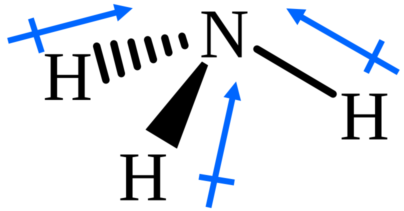
Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Apply the solution map.
  - ✓ Determine molecular polarity.

Bonds = polar

Shape = trigonal pyramid



Molecule = polar

Example:
Determine if NH_3 is polar.

Information:

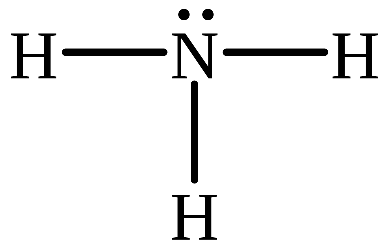
Given: NH_3

Find: If polar

Solution Map:

formula → Lewis → polarity and shape → molecule polarity

- Check:



$N = 5$

$H = 3 \cdot 1$

Total $NH_3 = 8$

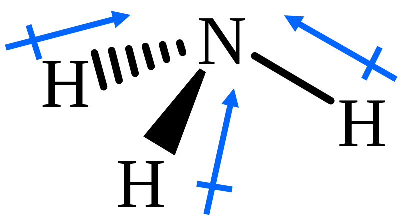
Bonding = $3 \cdot 2\ e^-$

Lone pairs = $1 \cdot 2\ e^-$

Total $NH_3 = 8\ e^-$

Bonds = polar

Shape = trigonal pyramid

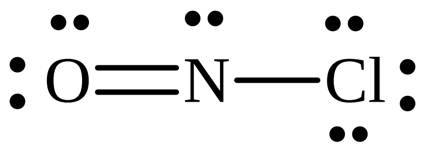


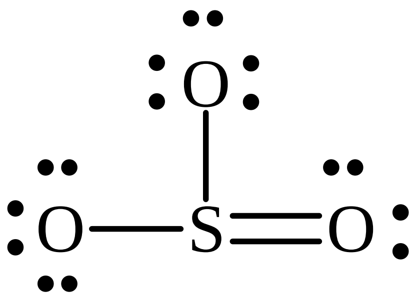
Molecule = polar

The Lewis structure is correct. The bonds are polar and the shape is unsymmetrical, so the molecule should be polar.

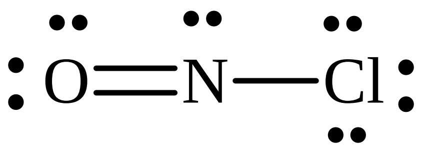
Practice—Decide Whether Each of the Following Molecules Is Polar

EN
O = 3.5
N = 3.0
Cl = 3.0
S = 2.5

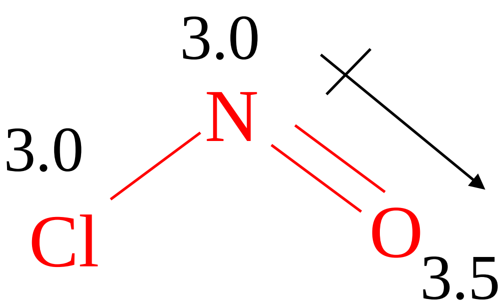




Practice—Decide Whether the Each of the Following Molecules Is Polar, Continued

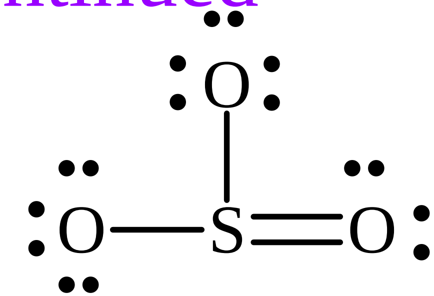


Trigonal
bent

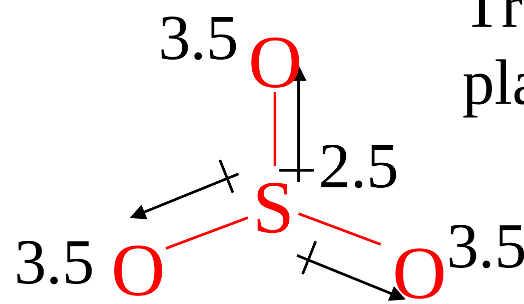


1. Polar bonds, N—O
2. Asymmetrical shape

Polar



Trigonal
planar



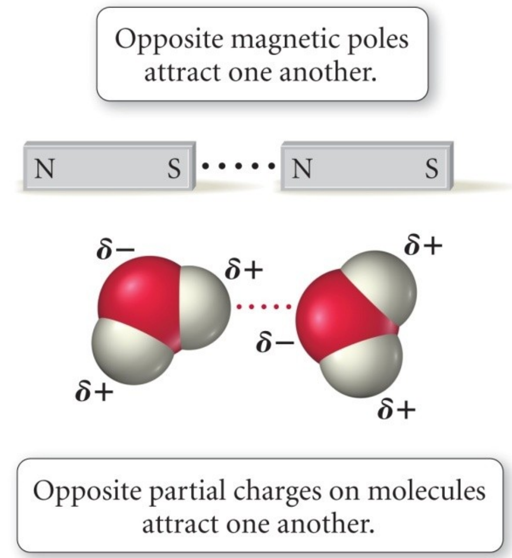
1. Polar bonds, all S—O
2. Symmetrical shape

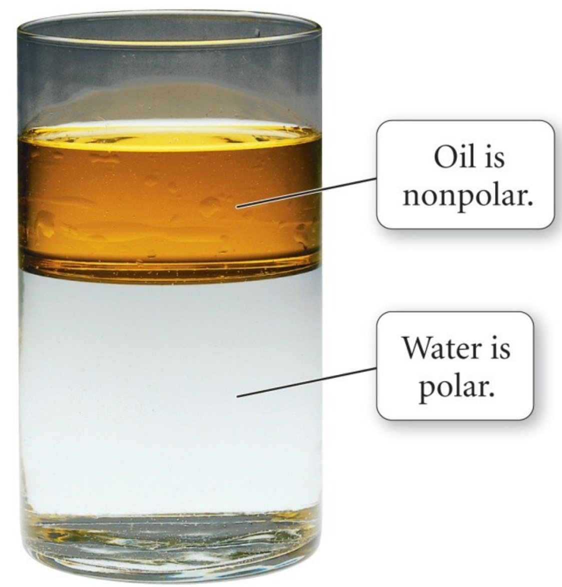
Nonpolar

Molecular Polarity Affects

Solubility in Water

- Polar molecules are attracted to other polar molecules.
- Since water is a polar molecule, other polar molecules dissolve well in water.
    - ✓ And ionic compounds as well.
- Some molecules have both polar and nonpolar parts.





- In HCl, electrons spend more time near the chlorine than the hydrogen. Although the molecule is overall neutral, the chlorine is more negative than the hydrogen, resulting in partial charges on the atoms.
- Partial charges are represented by a δ- on the more negative atom and δ+ on the more positive atom.
- The ability of an atom to attract electrons is called the atom’s **electronegativity**.