

Chapter 8 Covalent Bonding

Polar Bonds and Molecules

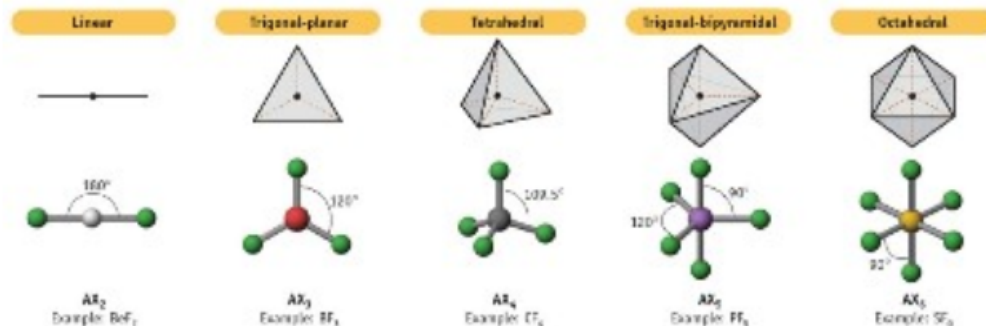
Molecular Compounds

The Nature of Covalent Bonding

Bonding Theories



COVALENT BONDING CHAPTER 8B



© 2008 Macmillan - Thomson

Topics:

1. Covalent Bonding

Objectives:

1. Define bond dissociation in relation to bond strength.
2. Recognize molecular orbitals (sigma and pi bonds) that relate to single, double and triple bonds.
3. Understand and draw molecules that have exceptions to the octet rule and demonstrate resonance (when bonds are intermediate in length).
4. Understand the theory behind molecular geometric shapes (linear, bent, trigonal planar, pyramidal, and tetrahedral).
5. Explain hybrid orbital theory in terms of molecular geometry.
6. Identify and define intermolecular attractions (Van Der Waals, Hydrogen Bonding, Network Solids) and how these affect chemical properties of the molecules.
7. Explain covalent bonding in terms of bonds (nonpolar, polar, and coordinate covalent) and molecules (nonpolar & polar).
8. Define and recognize polyatomic ions.
9. Understanding how to represent molecules, compounds and types of covalent bonds (single, double, triple) in various ways (molecular & structural formulas, Lewis structures).



Polarity

Complete the chart by filling in all requested information.

Consider how electrons are shared.

?	0	0.94	1.67	2.54	3.30
% ionic	? %	20 %	50 %	80%	90 %
% covalent	? %	80%	50 %	20 %	10 %
Examples	7 diatomic molecules?	Al_2Se_3	CaS	AlF_3	FrF CsF
Bond Polarity	? bonds	? bonds	? bonds	? Bonds ? Compounds	
Molecular Polarity	? molecule	? molecules	? molecule		

Polyatomic Ions (e.g. NH_4^+) have what type of bonds?



Polarity

Complete the chart by filling in all requested information.

Consider how electrons are shared.

END	0	0.94	1.67	2.54	3.30
% ionic	0 %	20 %	50 %	80%	90 %
% covalent	100%	80%	50 %	20 %	10 %
Examples	H ₂ N ₂ O ₂ F ₂ Cl ₂ I ₂ Br ₂	Al ₂ Se ₃	CaS	AlF ₃	FrF CsF
Bond Polarity	Non-polar bonds	Polar bonds	Polar bonds	Ionic Bonds Ionic Compounds	
Molecular Polarity	Non-polar molecules	Polar molecule	Polar molecule		

Polyatomic Ions (e.g. NH₄⁺) have coordinate covalent bonds.



How Do Scientists Predict Whether a Covalent Bond Will Form? How Are Covalent Bonds Illustrated?



IONIZATION ENERGIES AND ELECTRONEGATIVITIES

1												18							
<table border="1" style="width: 100%; height: 100%; border-collapse: collapse;"> <tr> <td style="width: 50%; text-align: center;">313</td> <td style="width: 50%;"></td> </tr> <tr> <td style="text-align: center;">H</td> <td style="text-align: center;">2.2</td> </tr> </table>	313		H	2.2	<p>← First Ionization Energy (kcal/mol of atoms)</p> <p>← Electronegativity*</p>										<table border="1" style="width: 100%; height: 100%; border-collapse: collapse;"> <tr> <td style="width: 50%;"></td> <td style="width: 50%; text-align: center;">567</td> </tr> <tr> <td style="text-align: center;">He</td> <td style="text-align: center;"></td> </tr> </table>		567	He	
313																			
H	2.2																		
	567																		
He																			
		2		13		14		15		16		17							
Li	125 1.0	Be	215 1.5	B	191 2.0	C	260 2.6	N	336 3.1	O	314 3.5	F	402 4.0	Ne	497				
Na	119 0.9	Mg	176 1.2	Al	138 1.5	Si	188 1.9	P	242 2.2	S	239 2.6	Cl	300 3.2	Ar	363				
K	100 0.8	Ca	141 1.0	Ga	138 1.6	Ge	182 1.9	As	226 2.0	Se	225 2.5	Br	273 2.9	Kr	323				
Rb	96 0.8	Sr	131 1.0	In	133 1.7	Sn	169 1.8	Sb	199 2.1	Te	208 2.3	I	241 2.7	Xe	280				
Cs	90 0.7	Ba	120 0.9	Tl	141 1.8	Pb	171 1.8	Bi	168 1.9	Po	194 2.0	At		Rn	248				
Fr	0.7	Ra	122 0.9	<p>* Arbitrary scale based on fluorine = 4.0</p>															

Bond Dissociation Energy

- Indicates the strength of a covalent bond.
- The **energy required to break the bond** between two covalently bonded atoms.
- A large bond dissociation energy corresponds to a strong covalent bond.
- The diatomic H_2 molecule has relatively high dissociation energy for a single covalent bond.

Table 8.3

Bond Dissociation Energies and Bond Lengths for Covalent Bonds

Bond	Bond dissociation energy (kJ/mol)	Bond length (pm)
H—H	435	74
C—H	393	109
C—O	356	143
C=O	736	121
C≡O	1074	113
C—C	347	154
C=C	657	133
C≡C	908	121
C—N	305	147
Cl—Cl	243	199
N—N	209	140
O—H	464	96
O—O	142	132

We will discuss **five exceptions** to the simple theory of covalent bonding:

Molecular Orbitals ... sigma and pi bonds

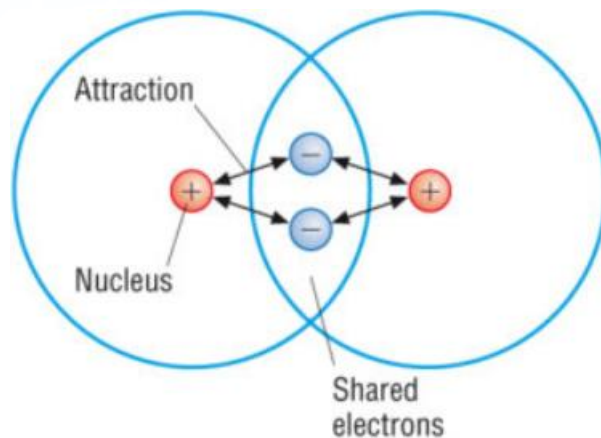
Exceptions to the octet rule

Resonance ... intermediate bond length

VSEPR Theory ... electron repulsion

Hybrid Orbitals ... sp , sp^2 , sp^3 orbitals

The model we have been using for covalent bonding assumes the orbitals are those of the **individual atoms**.



- The valence-bond model can't adequately explain the fact that some molecules contains two equivalent bonds with a bond order between that of a single bond and a double bond.
- In general, covalent bonds result from an imbalance between attraction and repulsion of the nuclei and electrons involved.

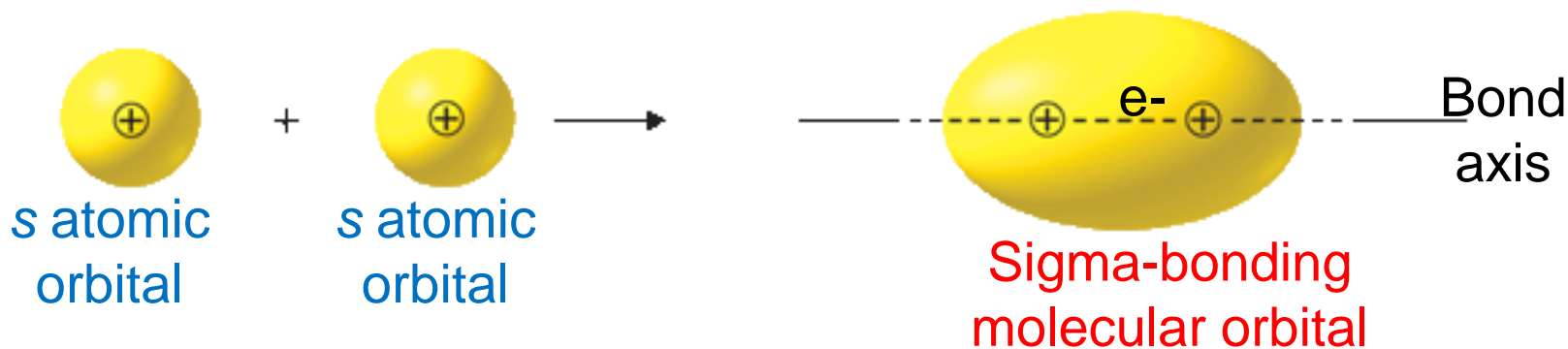
Molecular orbitals are obtained when two atoms combine and their atomic orbitals overlap to produce **molecular orbitals**, or orbitals that apply to the entire molecule.

Scientists use X-ray diffraction to map electron density within a molecule. Just as on the topographic map, the closer together the map lines are, the greater the **electron density** in the molecule.



Sigma Bonds (σ)

When two atomic orbitals (e.g. “s” in this case) combine to form a molecular orbital that is ***symmetrical around the axis*** connecting two atomic nuclei, a **sigma bond** is formed (σ).

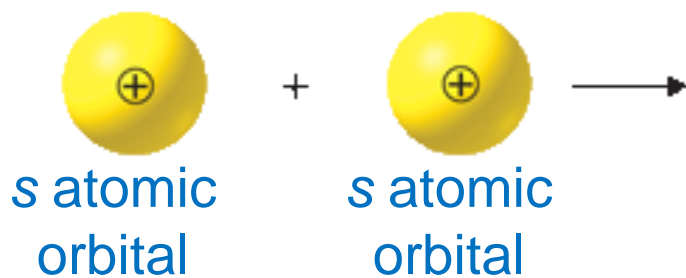


\oplus *represents the nucleus*

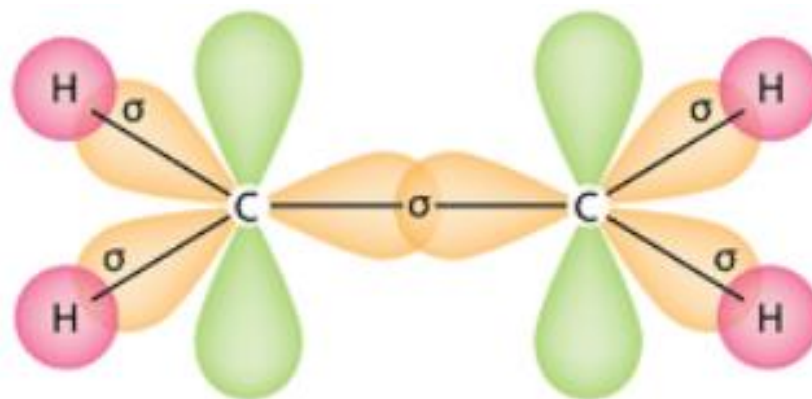
Sigma Bonds (σ)

Sigma bonds form between the nuclei of bonding atoms along the central axis.

Diatomic molecules of the same element exhibit sigma bonds (e.g. H_2 , F_2).



\oplus represents the nucleus

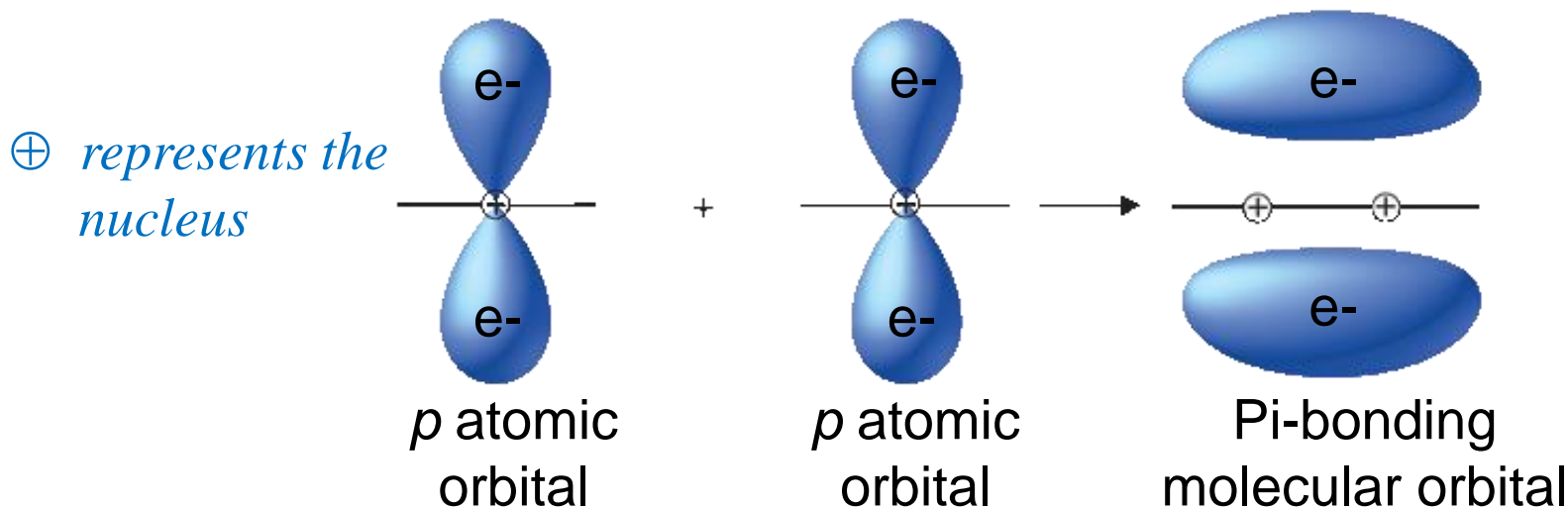


Sigma (σ) bond

Pi Bonds (π)

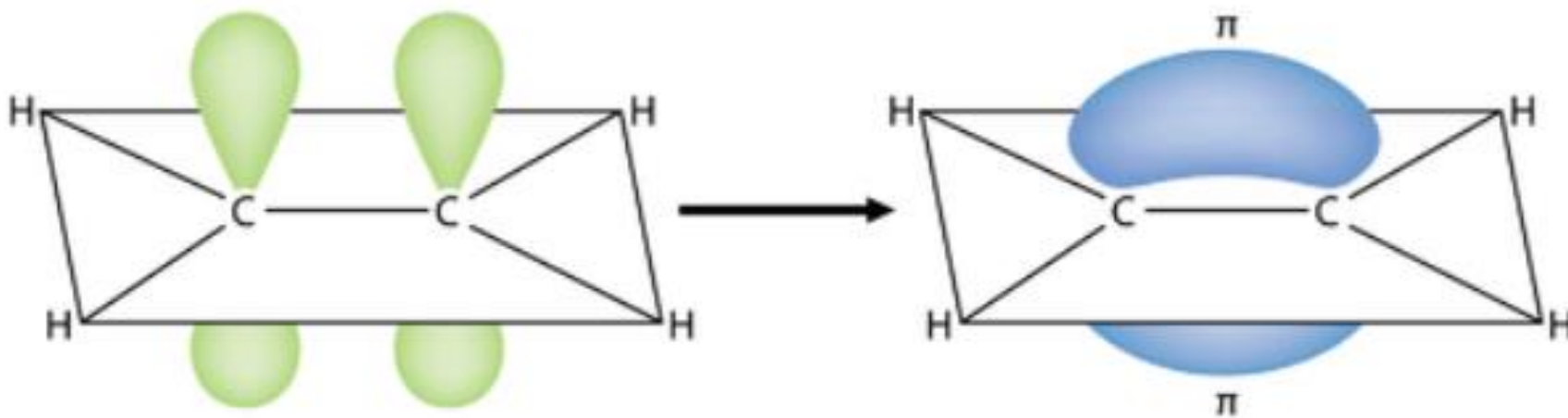
Molecular orbitals can also **overlap side to side**, **above and below the bond axis**. Pi bonds tend to be **weaker** than sigma bonds.

As shown here, the side-by-side overlap of atomic p orbitals produces what are called pi molecular orbitals.



Pi Bonds (π)

Electron density is concentrated above and below the axis plane.



Overlap of p orbitals leading to pi (π) bond

Orbital Overlap and Pi Bonds

Molecular Orbitals

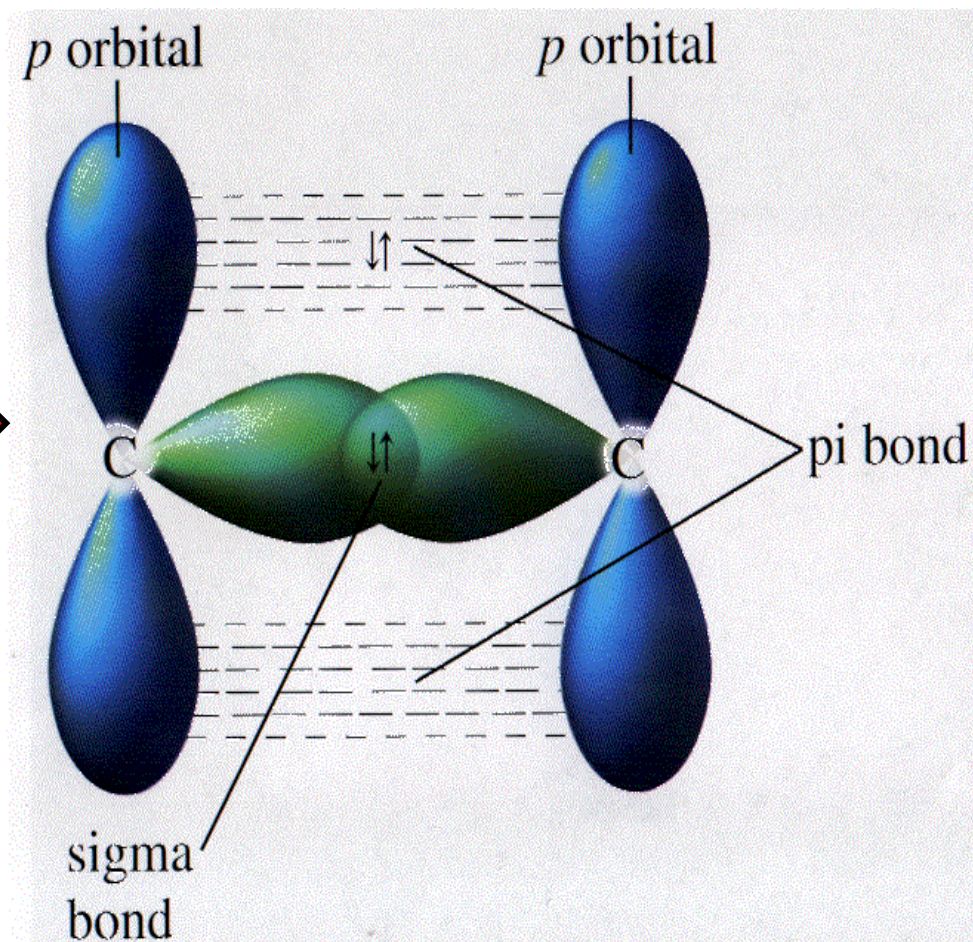
■ Double bond:

- One sigma bond (σ)
- One **pi bond** (π)

■ Triple bond



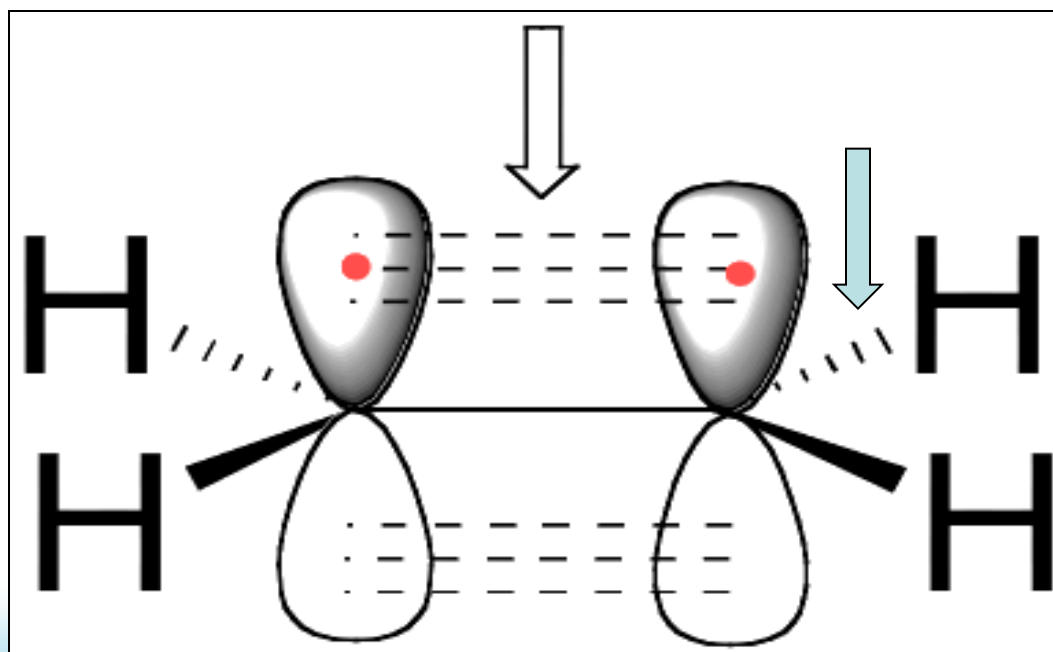
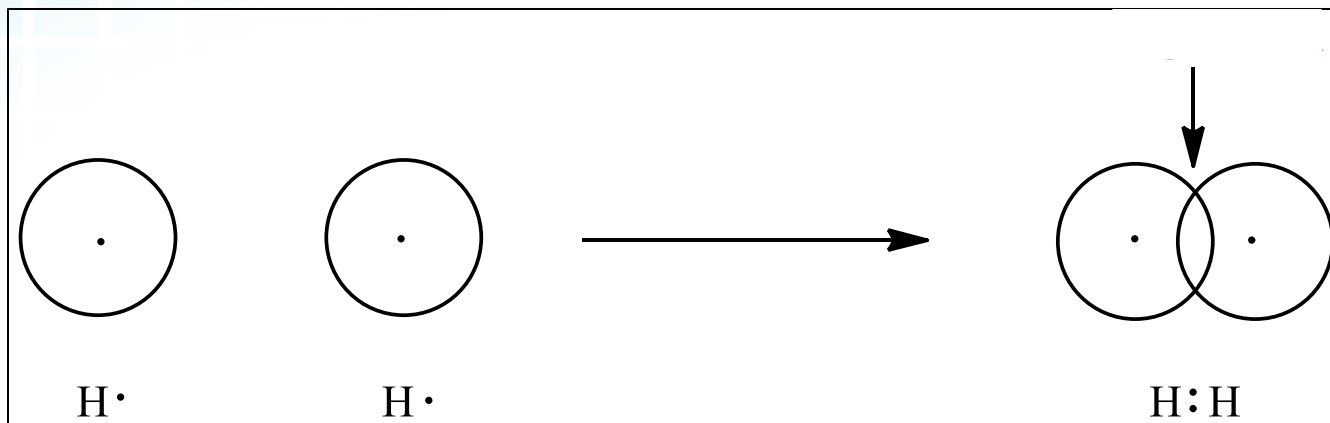
- One sigma bond (σ)
- Two pi bonds (π)
- The second pi bond oriented 90° from the first



8.1 Molecular Compounds >



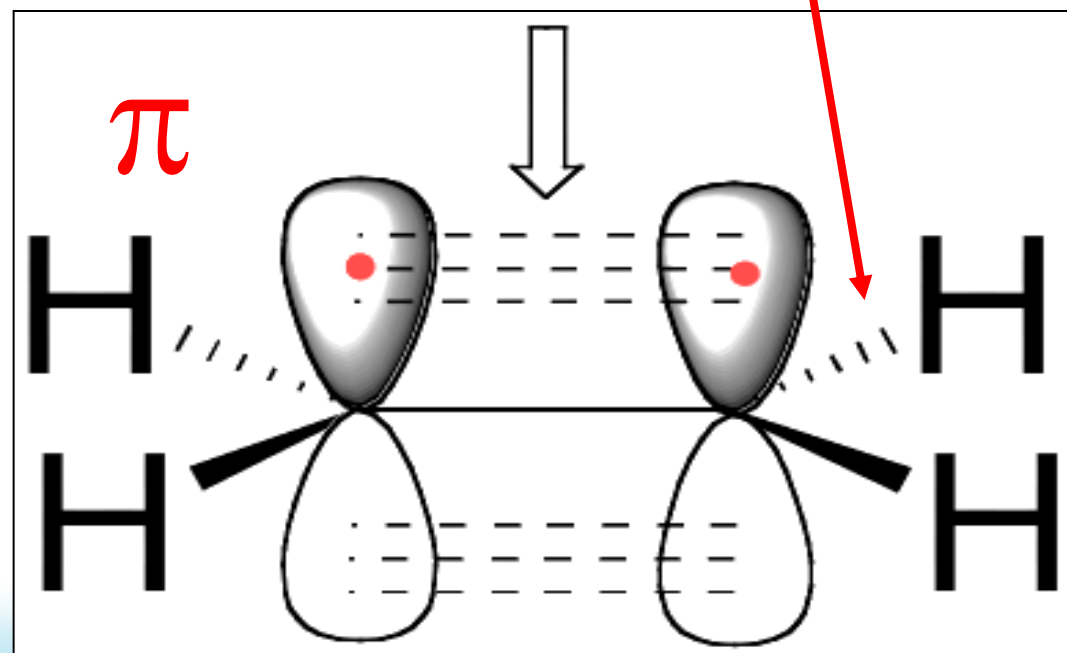
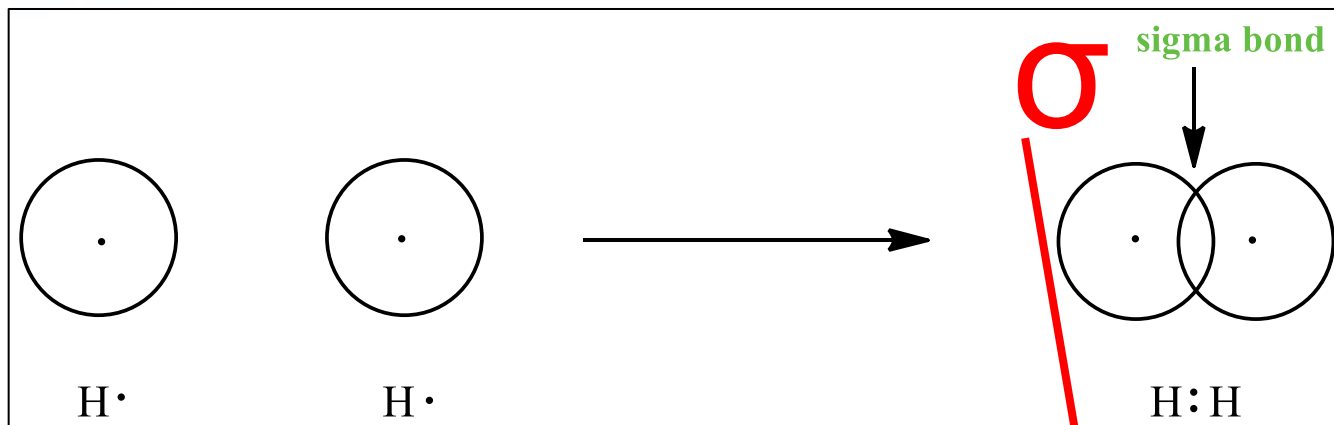
Which shows a π bond, and which a σ bond?
When would we observe each type?

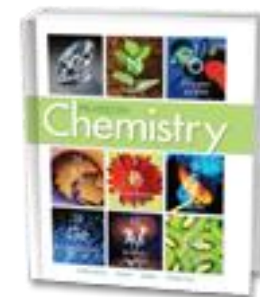


8.1 Molecular Compounds >



The top diagram shows a σ bond (when atoms exhibit symmetry around an axis) and the bottom diagram shows a π bond (when orbitals of atoms overlap side by side as in double & triple bonds).





Sigma & Pi Bonds (3:12)

<https://screencast-o-matic.com/watch/cF6hDnYoal>

Bonding and Nonbonding Electrons

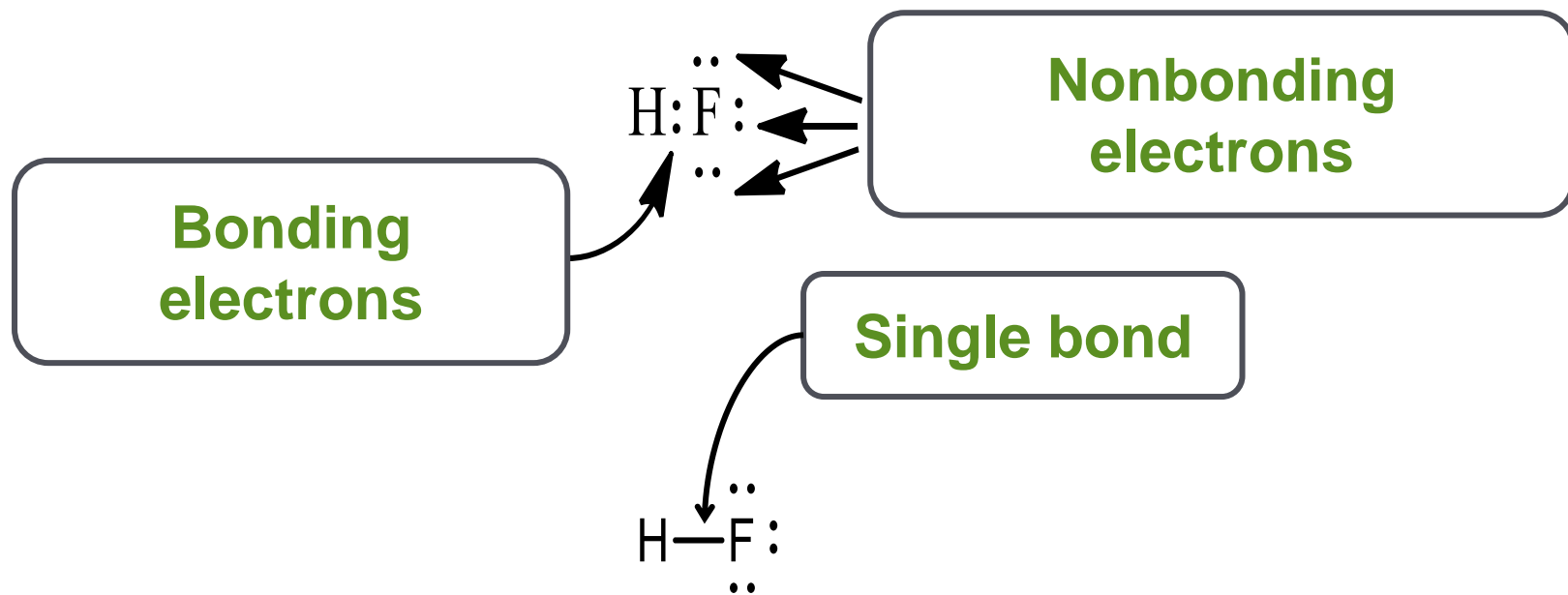
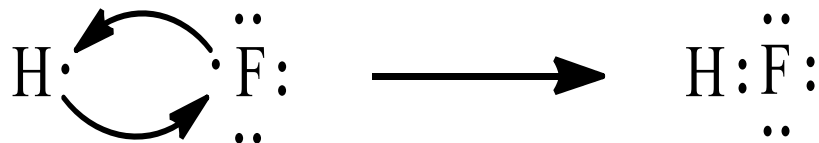


Show the Lewis Structures for the bond between hydrogen and fluorine (*before and after bonding*).

Point out the “bonding electrons”, the bond formed, and non-bonding electrons.

Bonding and Nonbonding Electrons

TRY IT



Applying the Octet Rule to PH_3

(phosphorus Hydride)



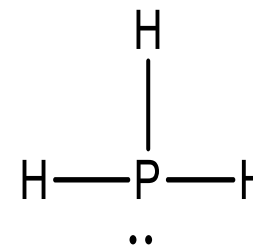
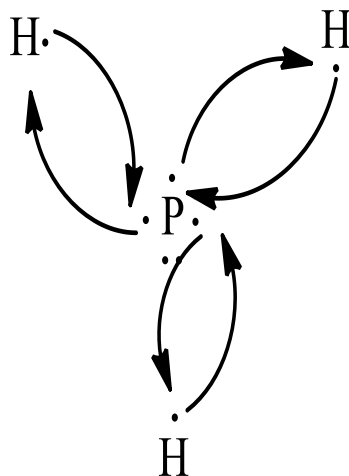
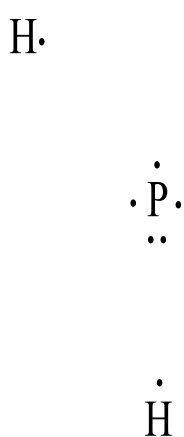
1. Draw the Lewis symbols of the atoms in the structure.
2. Form covalent bonds between the atoms.
3. The leftover electrons stay with the central atom.

octet rule: the general principle that atoms of nonmetals tend to be most stable when their valence shells are filled

Applying the Octet Rule to PH_3

(phosphorus Hydride)

TRY IT

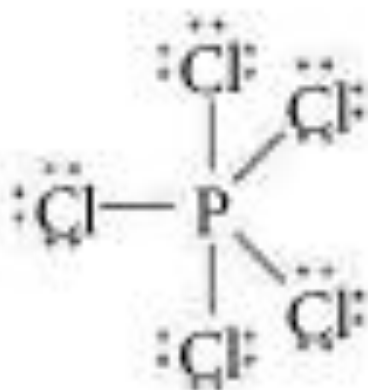


1. Draw the Lewis symbols of the atoms in the structure.
2. Form covalent bonds between the atoms.
3. The leftover electrons stay with the central atom.

octet rule: the general principle that atoms of nonmetals tend to be most stable when their valence shells are filled

Some molecules do not fulfill the octet rule (*8 valence electrons*), but are known to be stable molecules.

Phosphorus pentachloride



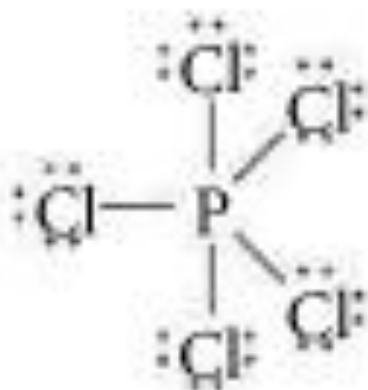
Phosphorus in PCl_5 (*Phosphorus pentachloride*) does not have an octet in its valence. What is its valence?

8.2 Nature of Covalent Bonding

Exceptions to the Octet Rule

Some molecules do not fulfill the octet rule (*8 valence electrons*), but are known to be stable molecules.

Phosphorus pentachloride



Phosphorus in PCl_5 (*Phosphorus pentachloride*) does not have an octet in its valence. What is its valence?

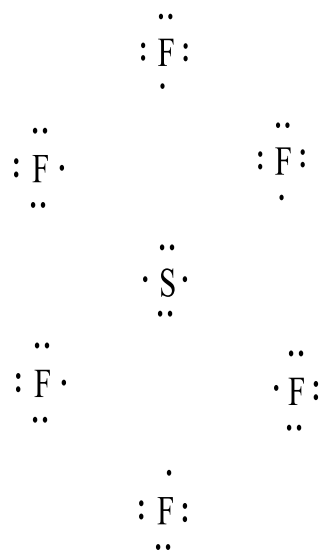
Sulfur hexafluoride (SF_6):



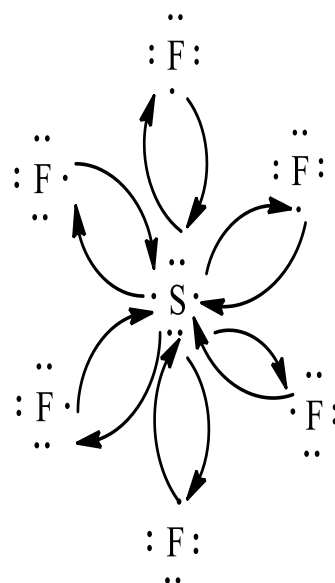
1. Draw the Lewis symbols of the atoms in the structure.
2. Form covalent bonds between the atoms.
3. Any extra electrons stay with the central atom.

Exceptions to the Octet Rule

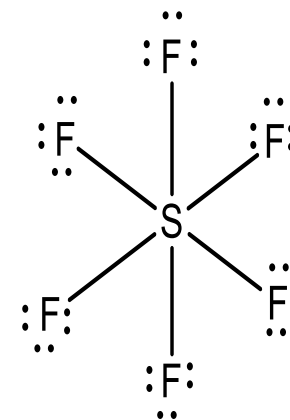
What is sulfur's valence?



1. Draw the Lewis symbols of the atoms in the structure.



2. Form covalent bonds between the atoms.



3. Any extra electrons stay with the central atom.

Sulfur's valence is 12 e-



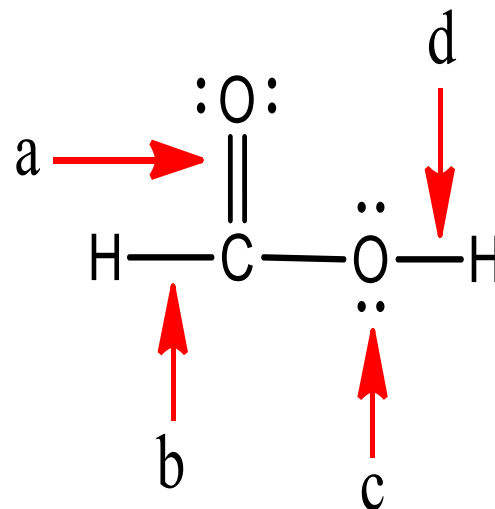
Identify Components of a Lewis Structure

Identify arrows pointing to bonding electrons.

Identify arrows pointing to nonbonding electrons.

Identify arrows pointing to structures containing sigma bonds.

Identify arrows pointing to structures containing pi bonds.





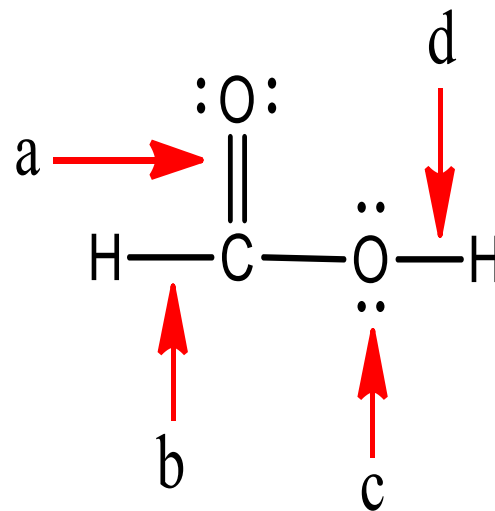
Identify Components of a Lewis Structure

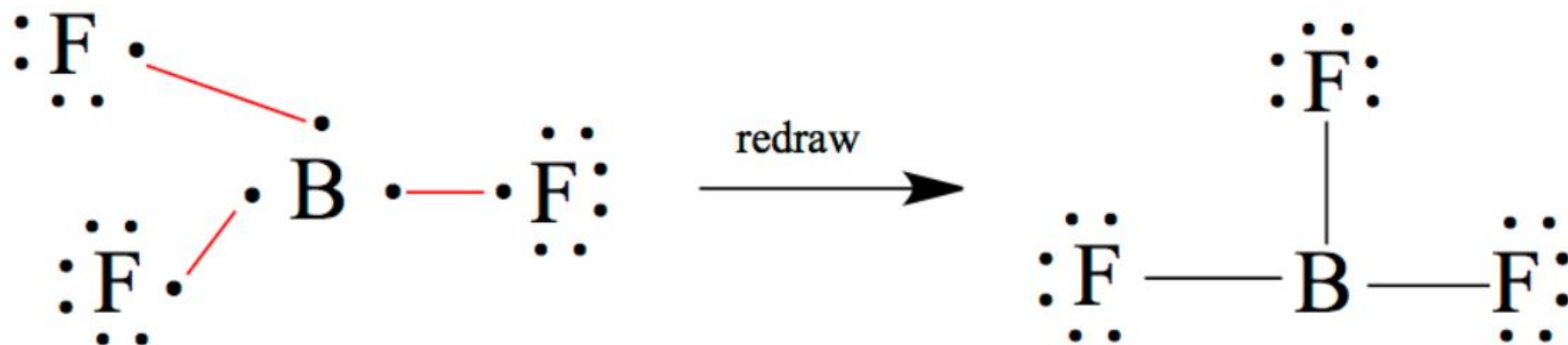
Identify arrows pointing to bonding electrons. **a, b, d**

Identify arrows pointing to nonbonding electrons. **c**

Identify arrows pointing to structures containing sigma bonds. **a, b, d**

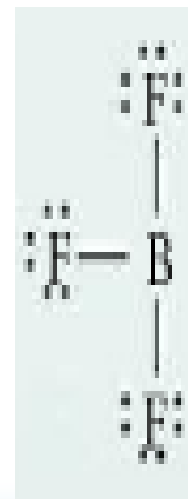
Identify arrows pointing to structures containing pi bonds. **a**

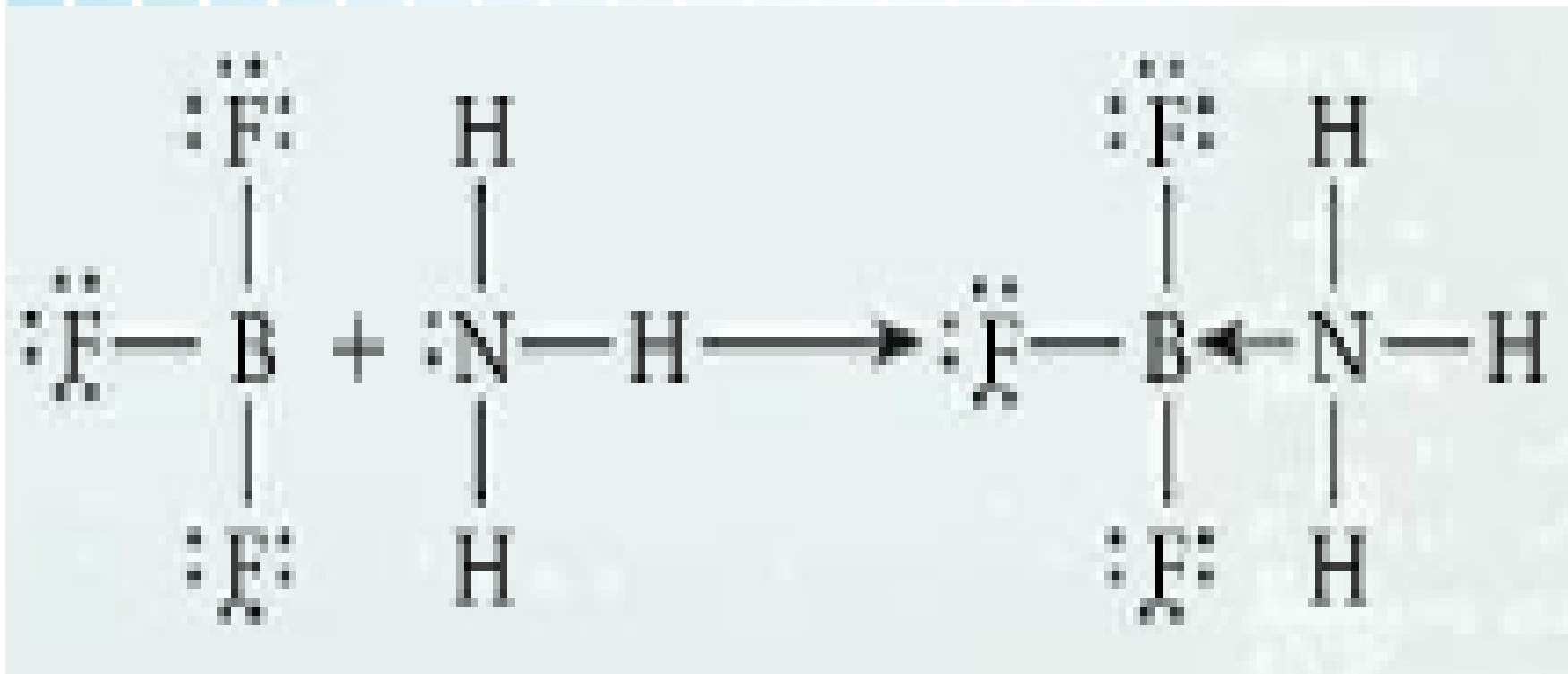




Boron in BF_3 (*Boron Trifluoride*) does not have an octet.

Boron in BF_3 has 6 valence electrons.

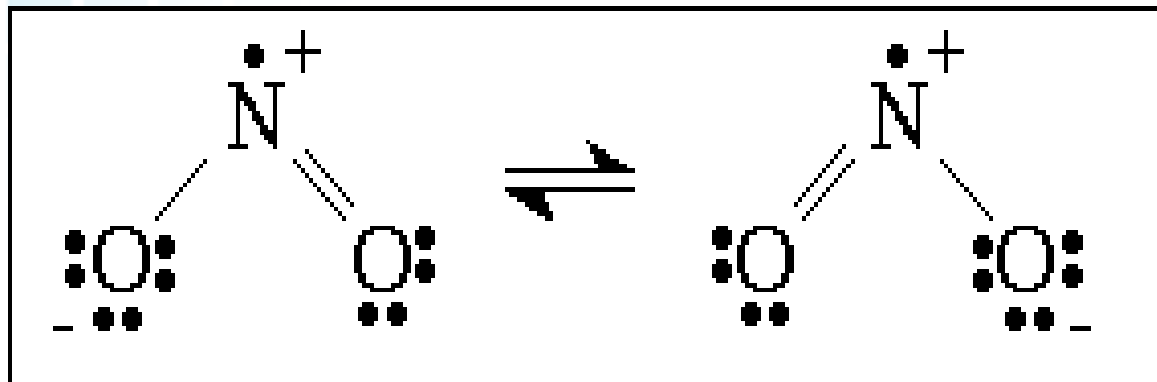




Boron in BF₃ (*Boron Trifluoride*) bonds with ammonia forming a compound with the octet rule.

Boron and nitrogen (the central atoms) have complete octets.

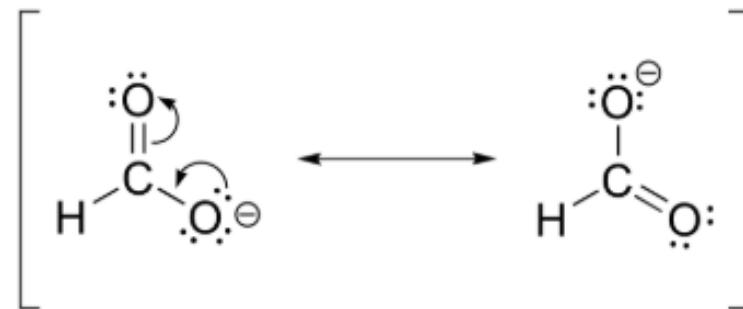
NO₂ exists as a stable molecule, but more than one Lewis Structure can be drawn.



Nitrogen dioxide molecule

The double arrows in between the molecules indicate

Resonance structures. Carboxylate Anion

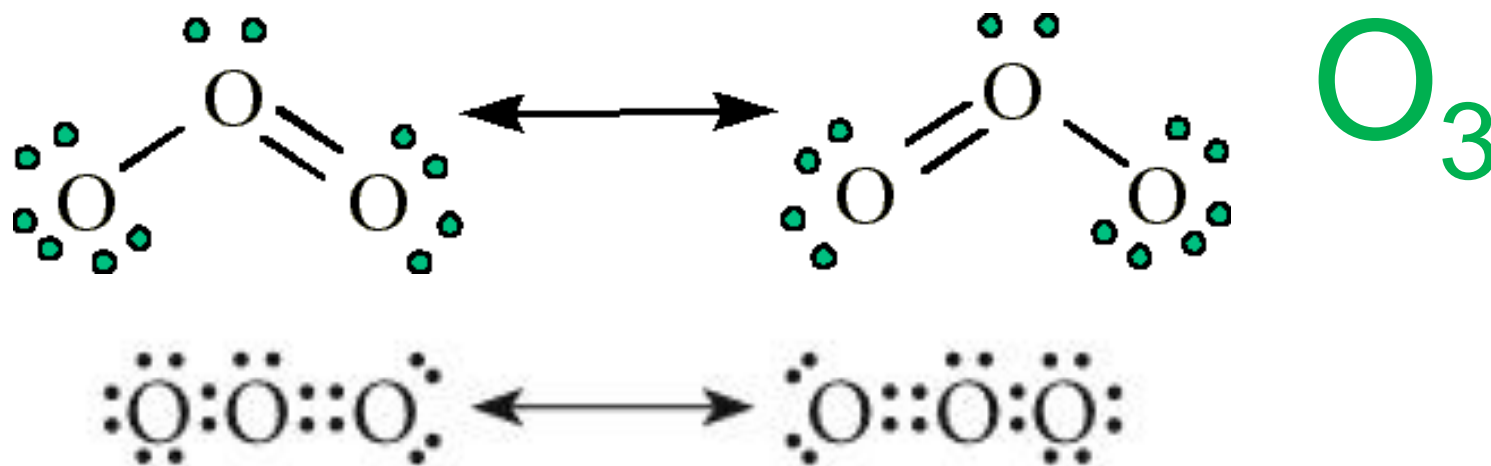


Resonance Structures

Resonance structures are used for molecules that cannot be adequately described by a single structural formula, but is a hybrid, or mixture, of two extremes.

Ozone (O_3) has two resonance structures made by shifting electron pairs without changing the positions of the oxygen atoms.

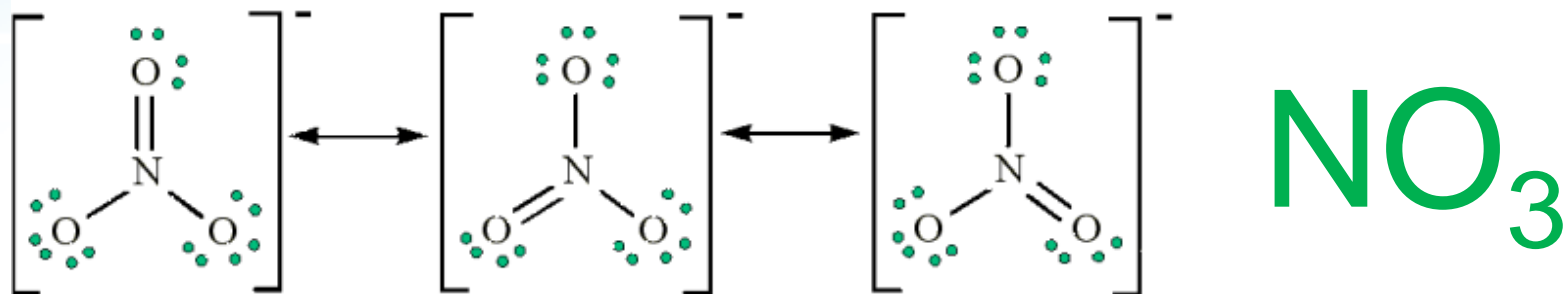
The two bonds in ozone are the same length, but are **in-between** the length of a single or double bond.



8. 2 Nature of Covalent Bonding

Resonance

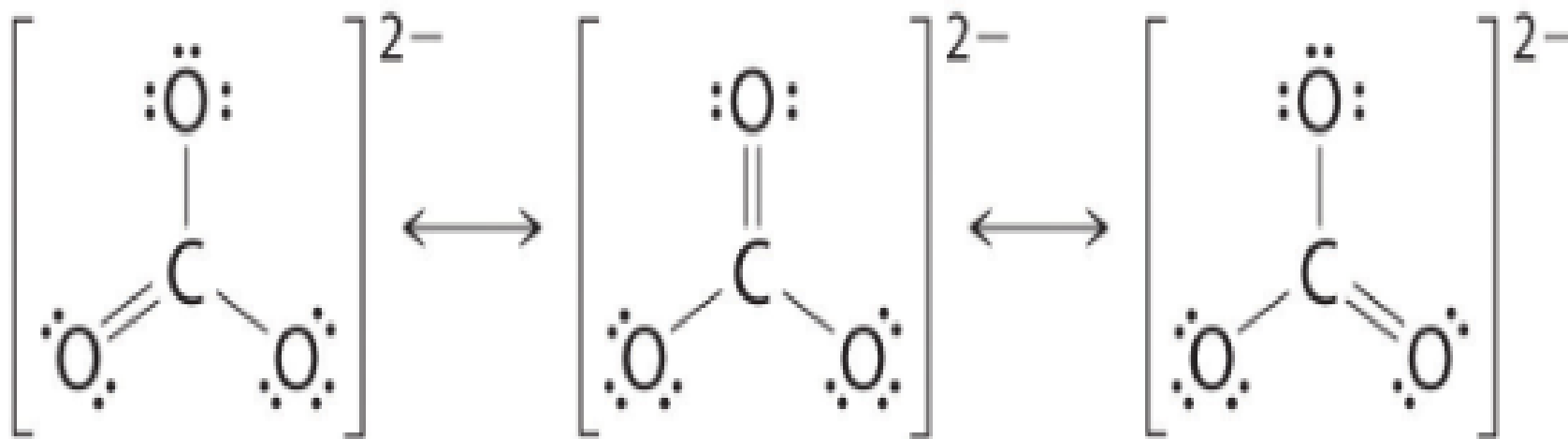
Double-headed arrows indicate that two or more structures are in resonance.



Resonance makes it possible to draw two or more valid Lewis structures with some molecules or ions, indicating **intermediate bond lengths**.

Draw a resonance structure of CO_3^{2-} :

Draw a resonance structure of CO_3^{2-} :

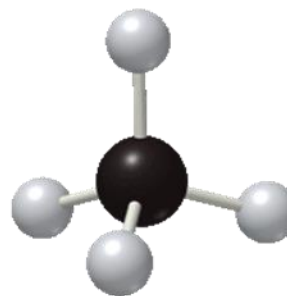
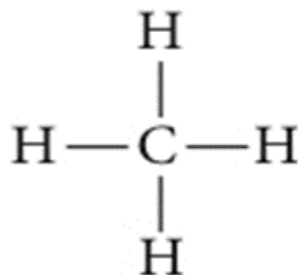
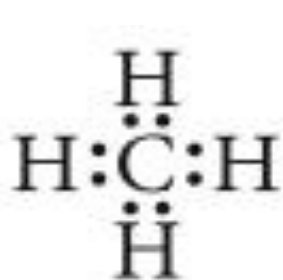


The three bonds in the carbonate polyatomic ion are the same length, but are **in-between** the **length** of a single or double bond.

Valence-Shell Electron-Pair Repulsion Theory

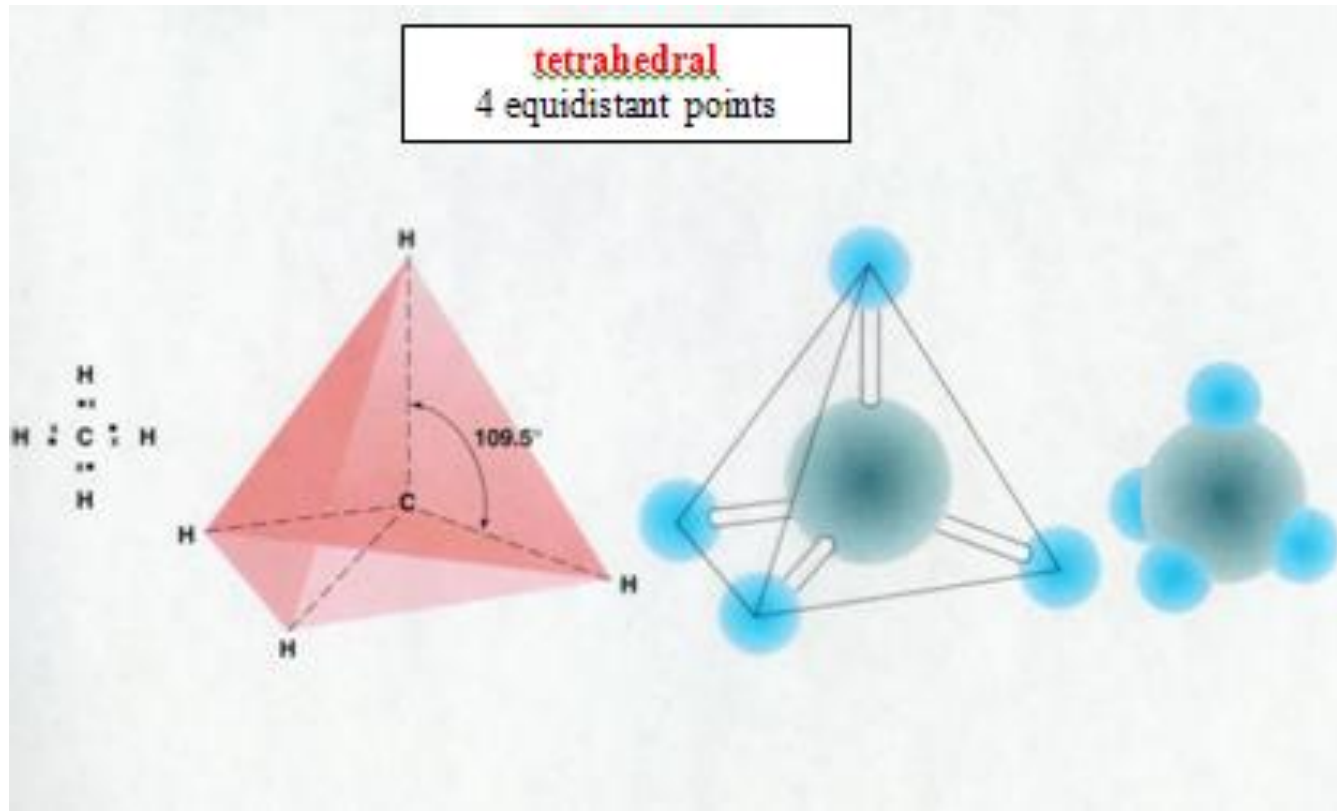
Lewis structures don't reflect the 3D shapes of molecules.

The Lewis structure and structural formula of methane (CH_4) show the molecule as if it were flat and merely two-dimensional. In reality, CH_4 molecules are 3D.

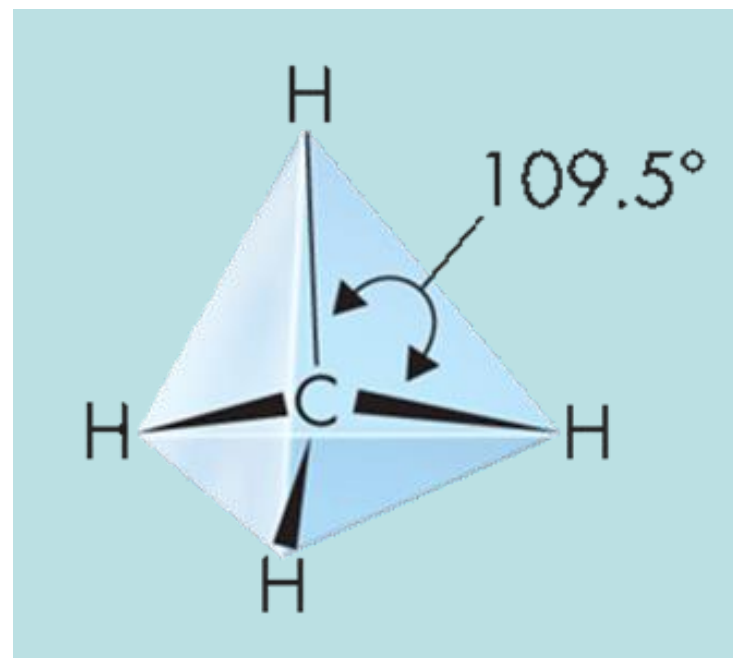
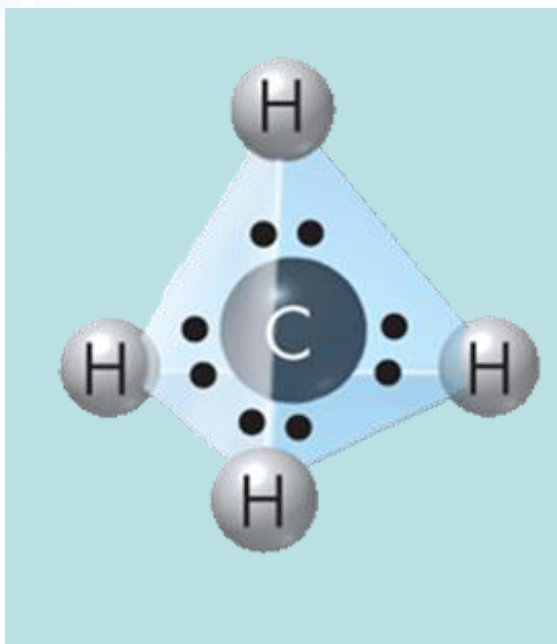


VSEPR theory states that the **repulsion** between **electron pairs** causes molecular shapes to adjust so that the valence-electron pairs stay as far apart as possible.

To form a methane molecule, the hydrogens are at the four corners of a **tetrahedron**. This shape occurs when there are four shared pairs of electrons around a central atom. **In this arrangement, all of the H-C-H angles are 109.5° , the tetrahedral angle.**



Tetrahedral molecules have 4 equidistant points around the central atom (e.g. group IVA elements).

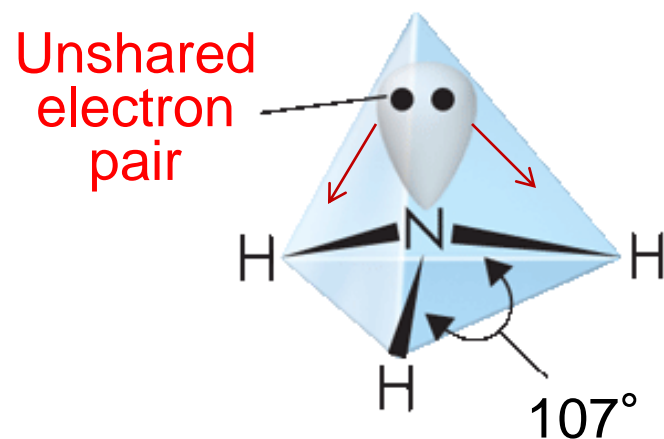


CCl_4 , Phosphate ion, Sulfate ion, CF_4 are all tetrahedral

Unshared pairs of electrons are also important in predicting the shapes of molecules.

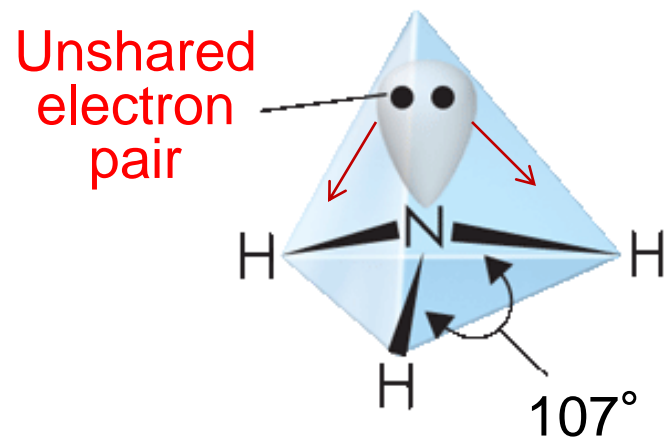
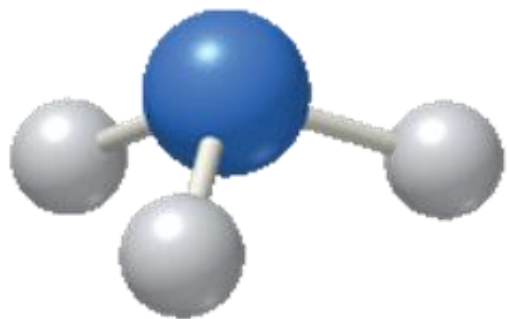
E.g. The nitrogen in ammonia (NH_3) is surrounded by four pairs of valence electrons. However, one of the valence-electron pairs is an unshared pair.

The unshared pair strongly repels the bonding pairs, forming a “pyramidal” shape.



Pyramidal molecules have 4 points around the central atom (e.g. group VA elements).

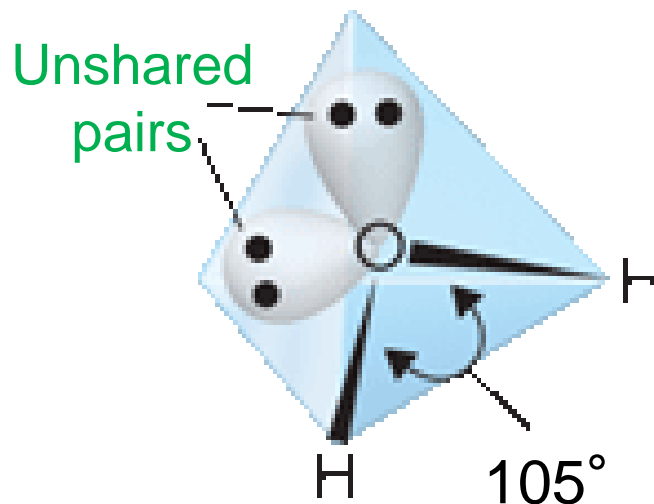
The bond angle is only 107° , rather than the tetrahedral angle of 109.5° .



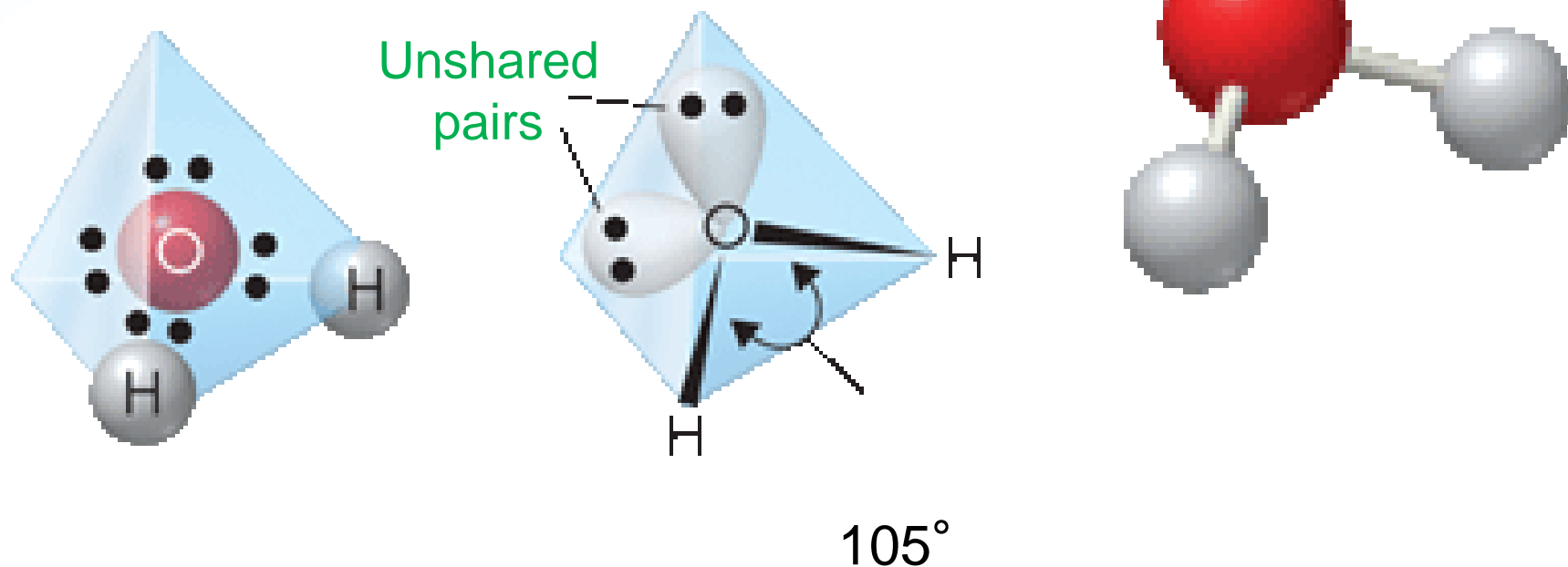
Group 5A atoms form pyramidal bonds: PCl_3 .

In a water molecule (H_2O) the two unshared pairs of electrons **repel each other** and cause the water molecule to take on a “**bent**” geometry.

The H-O-H bond angle is compressed to about 105° .

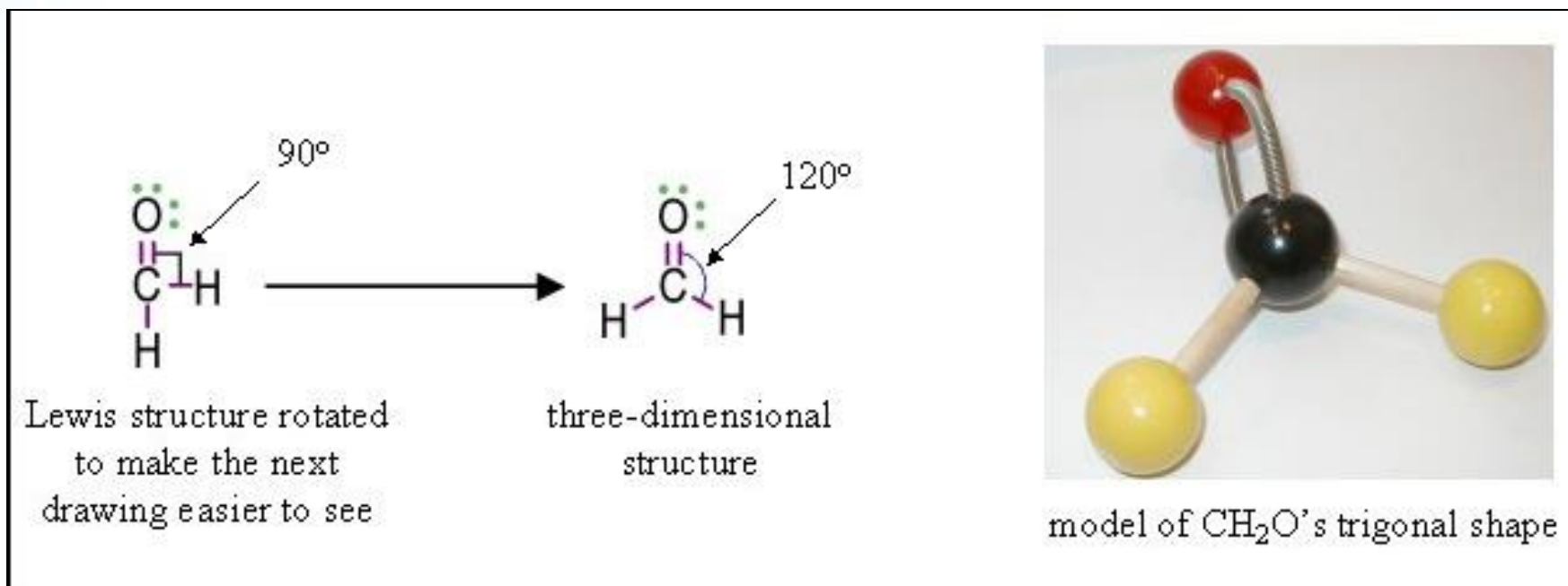


Bent molecules have 2 angled points around the central atom due to repulsion of the unbonded electron pairs (e.g. group VIA elements).

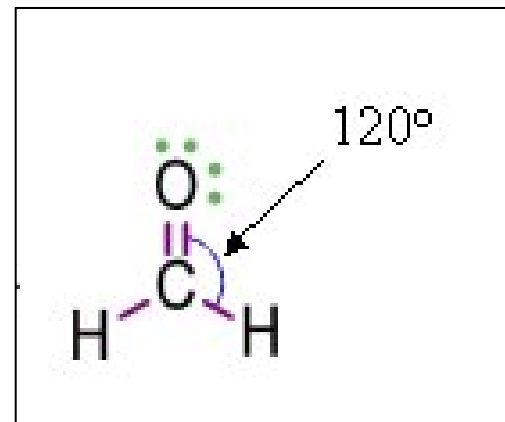
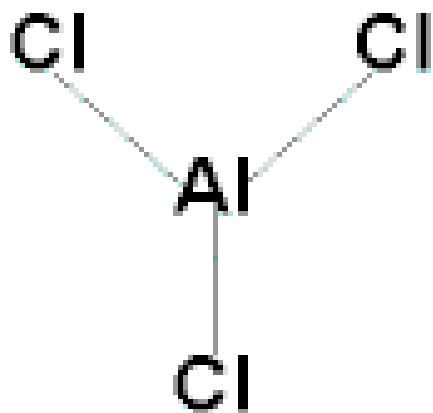
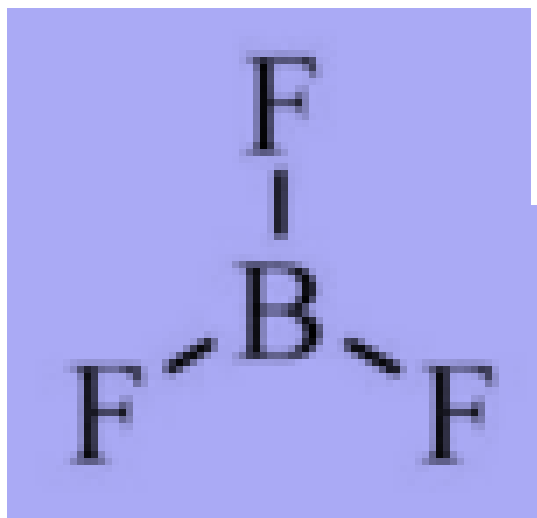


H₂S, SO₂, SCl₂, and NO₂ are all “bent”

Trigonal Planar molecules have a central atom with only three shared pairs of electrons around it, forming a bond angle of 120° (*equilateral triangle*).



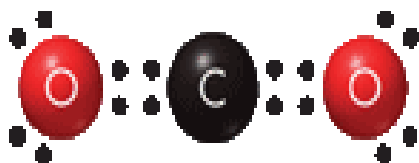
Trigonal Planar molecules have 3 equidistant points around the central atom (*e.g. group IIIA elements*).



Linear

The bond angle of linear molecules is 180° . The 7 diatomic molecules of the same element are linear.

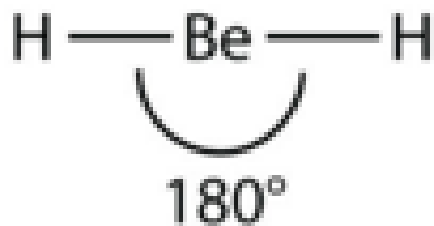
Carbon dioxide (CO_2)



180°



No unshared
electron pairs
on carbon

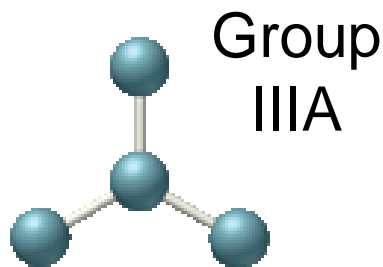




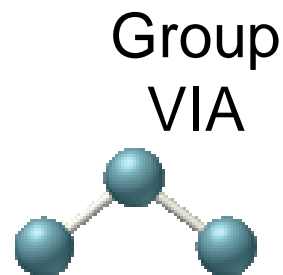
Some common molecular shapes. *The trigonal bipyramidal and the octahedral shapes involve bonds with electrons in d orbitals.*



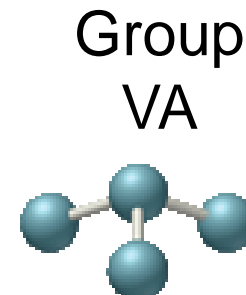
Linear



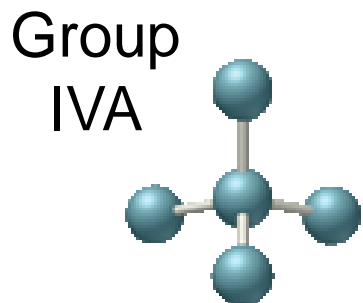
Trigonal planar



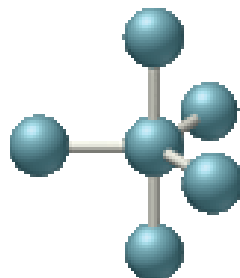
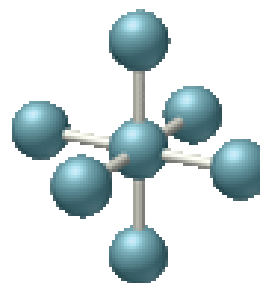
Bent



Pyramidal

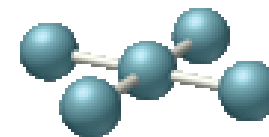


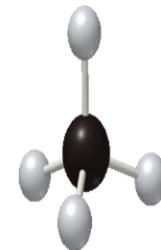
Tetrahedral

Trigonal
bipyramidal

Octahedral

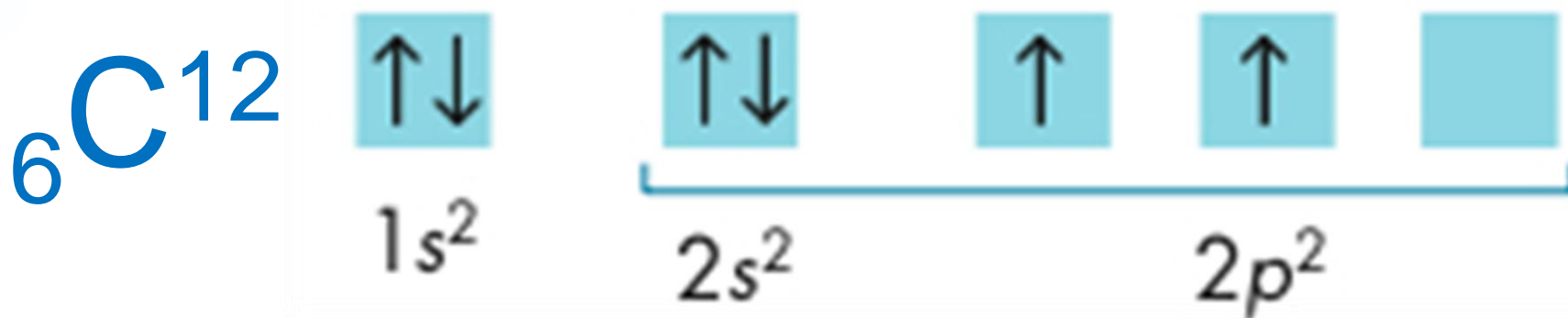
enrichment

Square
planar



The VSEPR theory works well for molecular shapes, but it does not work for describing the types of bonds formed.

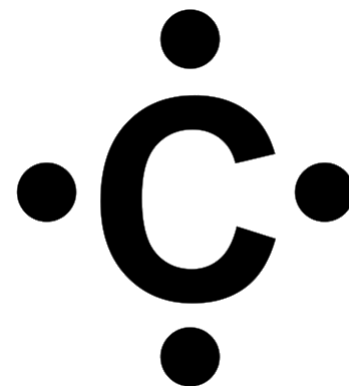
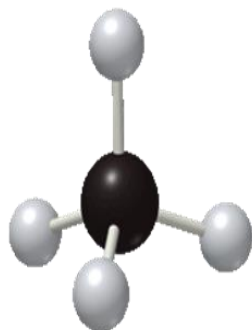
E.g. Methane is **tetrahedral** in shape, but you would not predict this pattern based on carbon's electron configuration.



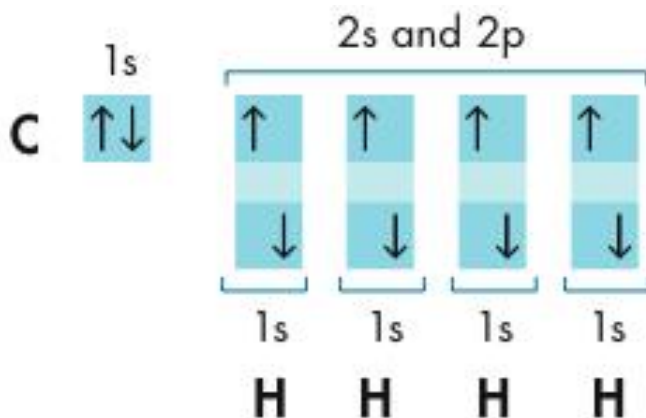
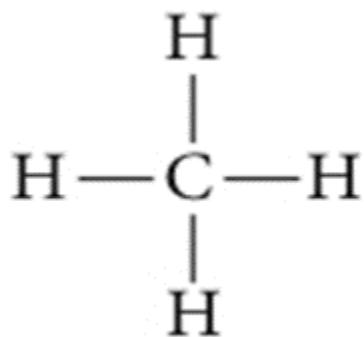
Based on the valence of carbon (as shown), one would expect the following Lewis Structure:



In reality, methane is tetrahedral. Therefore, the Lewis Structure of carbon actually looks like:

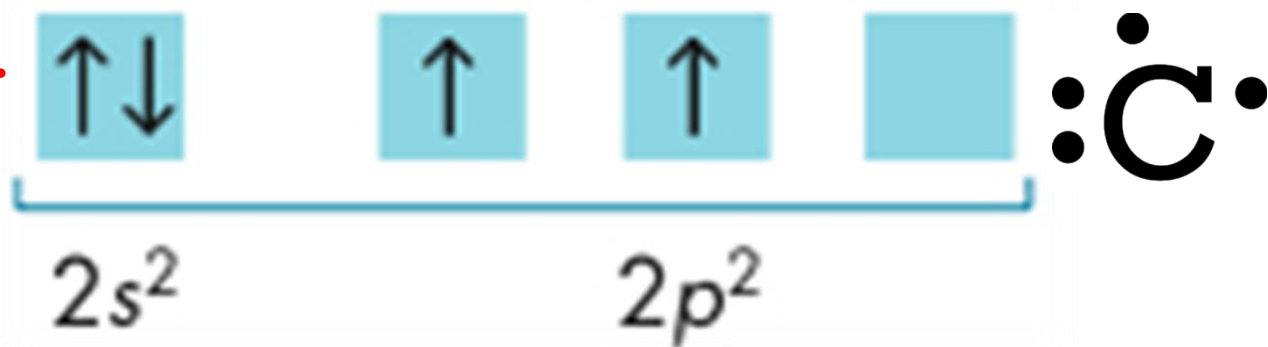


To adequately explain methane's actual shape, scientists created hybrid orbitals:



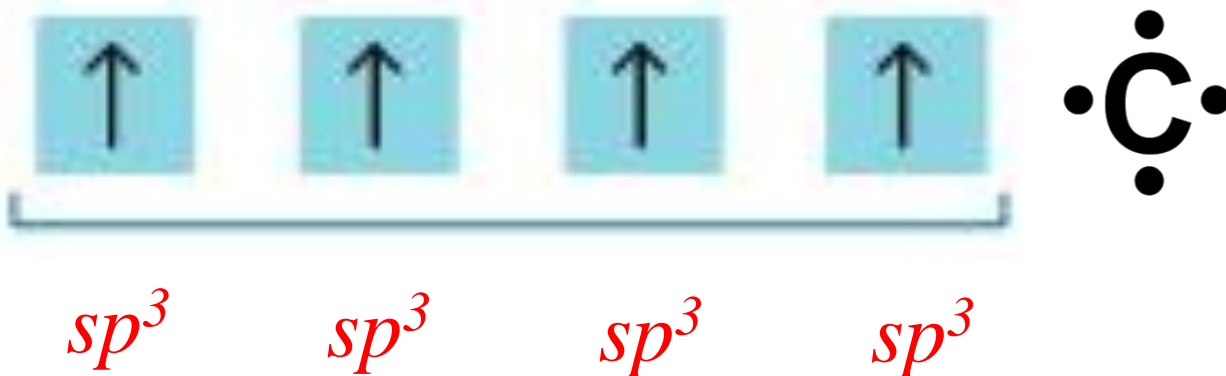
By manipulating carbon's valence electron configuration, scientists created "sp³" orbitals in place of 2s and 2p orbitals.

Original valence:



Hybrid valence:

*Composed of 1 s orbital
and 3 p orbitals*



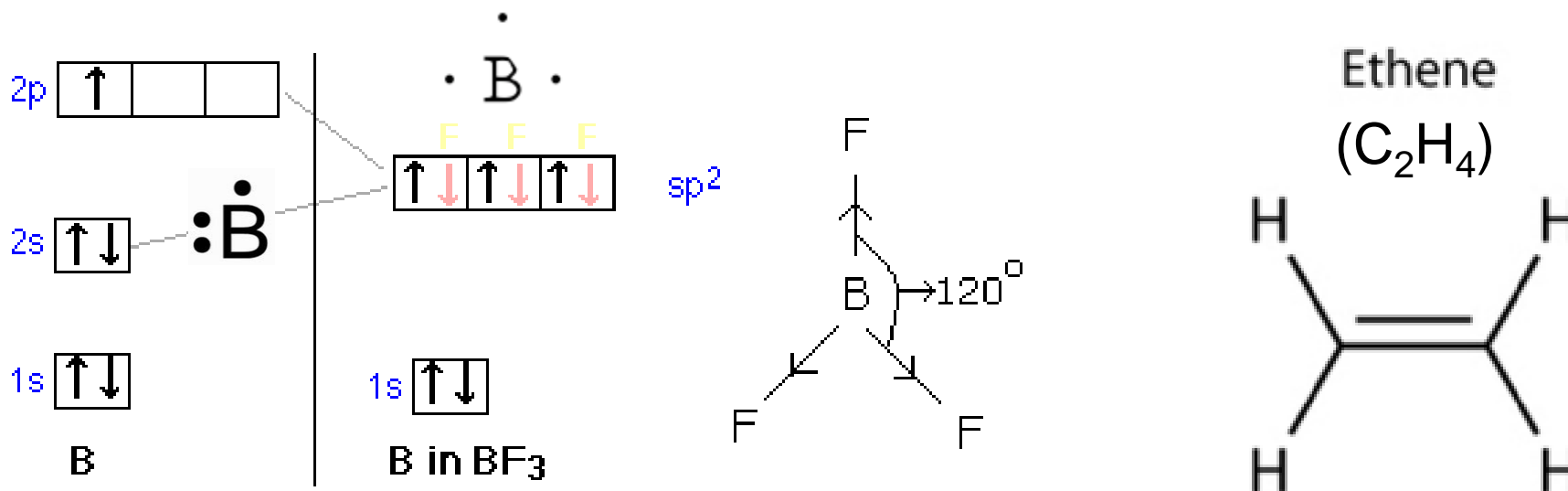
The hybrid orbitals allow for equidistant (tetrahedral) bonds

sp^2 Hybrid Orbitals

Explain the trigonal planar shape (*equilateral triangle*) for single bonded atoms like Boron & for double bonded molecules like Ethene.

sp^2 hybrid orbitals form from the combination of one 2s and two 2p atomic orbitals.

Each hybrid orbital is separated from the other two by 120° .

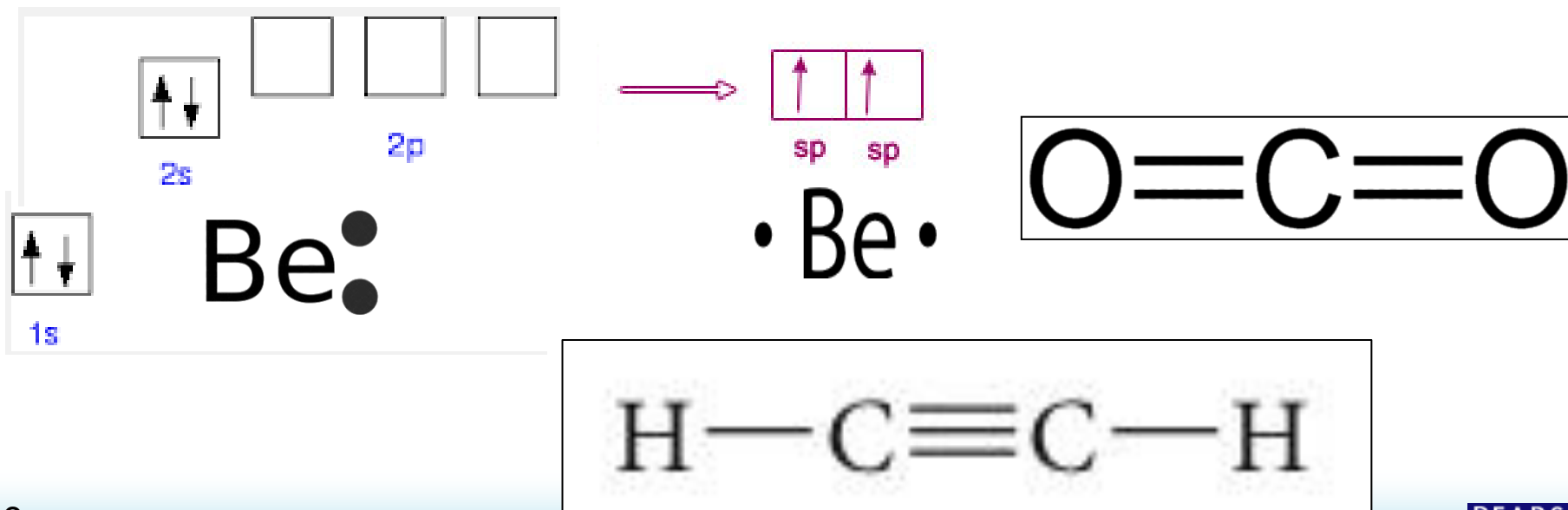


sp Hybrid Orbitals

Explain the **linear** shape for single bonded atoms like Beryllium (e.g. F-Be-F) & for double bonded molecules like carbon dioxide, and for triple bonded molecules like Ethyne (C_2H_2).

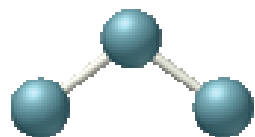
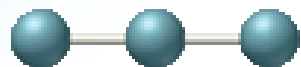
sp hybrid orbitals form from the combination of one 2s and one 2p atomic orbitals.

Each hybrid orbital is separated from the other two by 180° .





Match the molecule with the correct geometric symmetry. Arrange the examples. Give the hybrid orbitals that may be involved.



Bent



Linear



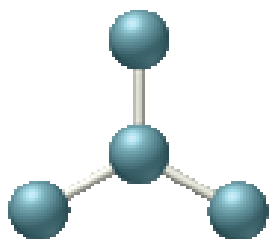
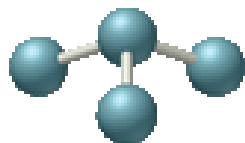
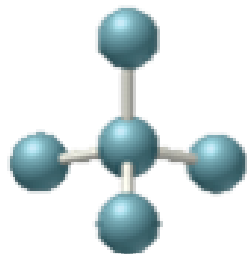
Pyramidal



Tetrahedral

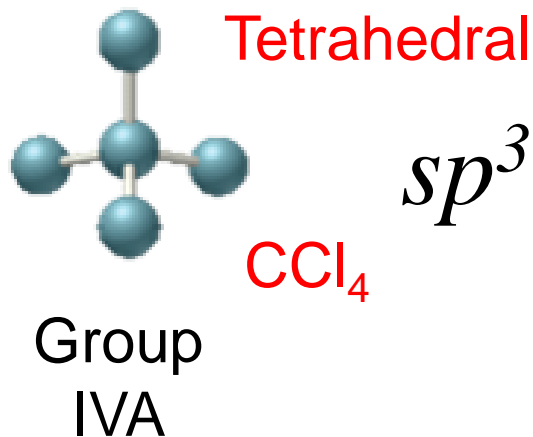
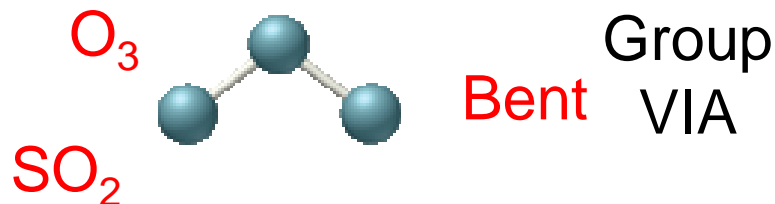
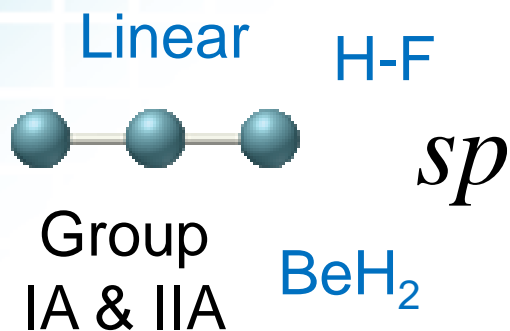


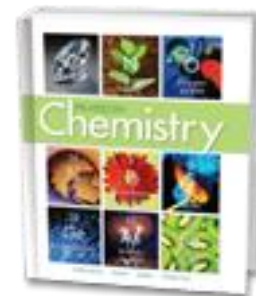
Trigonal planar





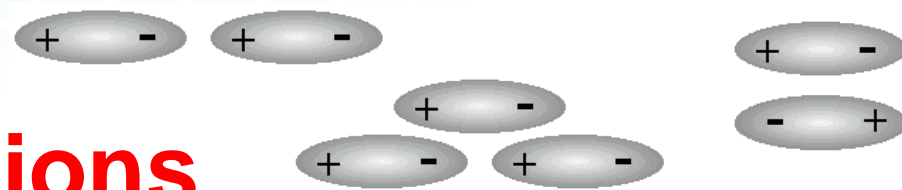
Match the molecule with the correct geometric symmetry. Arrange the examples. Give the hybrid orbitals that may be involved.





Hybrid Orbitals sp (linear), sp^2 (trigonal planar), and sp^3 (tetrahedral) Notes (4:43)

<https://screencast-o-matic.com/watch/cF6j2eYFFy>



Attractive Dipole-Dipole Interactions

Intermolecular attractions

MOLECULES can be attracted to each other by a variety of different forces called intermolecular attractions.

These are weaker than either ionic or covalent bonds that hold the individual ATOMS together.

Dispersion Forces are present between **ALL** molecules whether they are polar or nonpolar. They represent **TEMPORARY** “dipole” to “dipole” attractions).

Think of a SPECTATOR watching “TENNIS”. At any given instance, the ball is to the left or to the right (“Dipole”). On the average, however, the ball looks as if it is at the center (no overall charge).



Van der Waals Forces

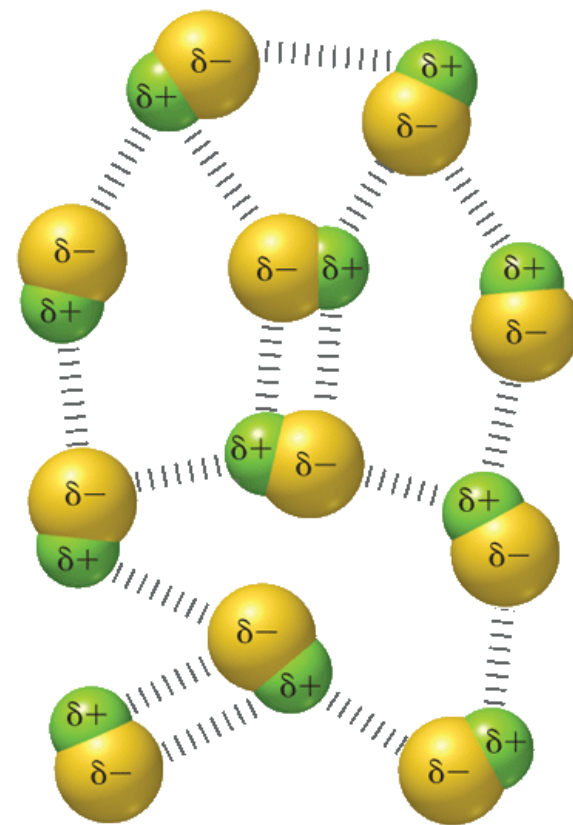


VDW forces consist of **dipole interactions** and **dispersion forces**. The slightly negative region of a polar molecule is weakly attracted to the slightly positive region of another polar molecule.

Dispersion forces, the weakest of all molecular interactions, are caused by the motion of electrons.

As electrons move, the charge within an atom shifts momentarily, producing a “dipole”.

The strength of dispersion forces generally increases as the number of electrons in a molecule increases (*e.g. down a group*).



Cl_2 , Br_2 , or I_2 molecules are diatomic halogens (same family/group), and have similar chemical properties and behaviors.

Each sample of molecules **is held WEAKLY but significantly together by dispersion forces.** These attractions are responsible for determining whether a molecular compound is a gas, a liquid, or a solid at a given temperature.

Boiling Temperature Noble Gases

Halogen	R.M.M	
F_2	38	g
Cl_2	71	g
Br_2	160	l
I_2	254	s

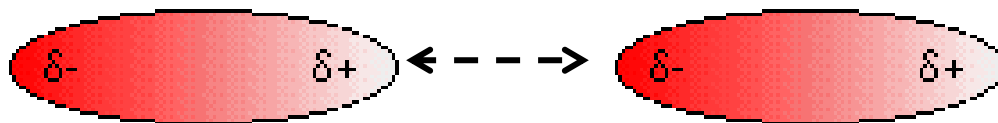
Iodine-53 is larger than bromine-35 which is larger than chlorine-17. So the dispersion forces are largest in iodine-53 and smallest in chlorine-17.

Therefore, iodine-53 has the highest melting & boiling points and is a solid at room temperature, bromine is a liquid, and chlorine is a gas (weakest VDW forces).

Dipole interactions are similar to, but much weaker than, ionic bonds.

Boiling Temperature Noble Gases

Halogen	R.M.M	
F ₂	38	g
Cl ₂	71	g
Br ₂	160	l
I ₂	254	s



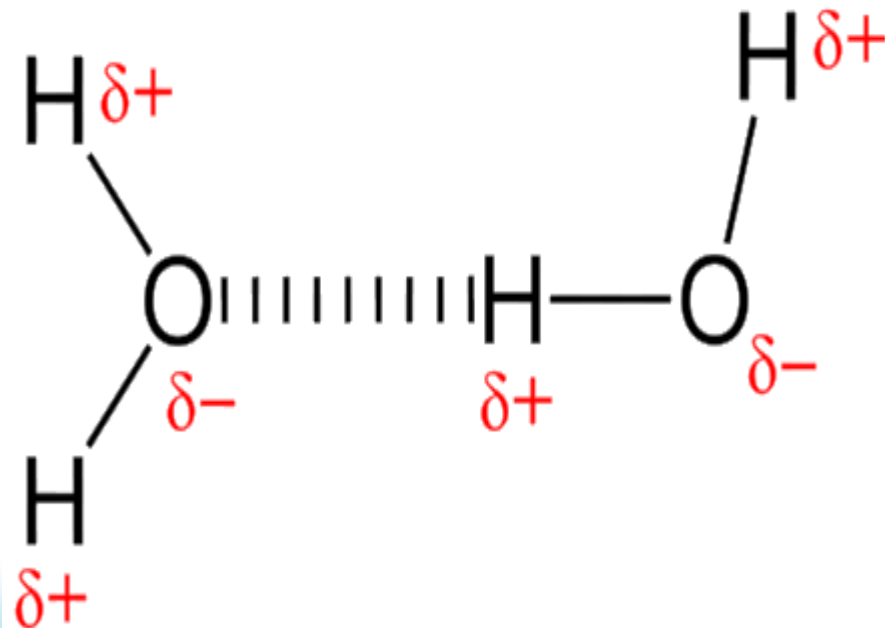
Hydrogen Bonds

A dipole – dipole attraction when hydrogen (*a very small atom*) bonds with a highly electronegative atom (e.g. O, F).

For water each O—H bond in the water molecule is highly polar (END $3.5 - 2.2 = 1.3$), and the oxygen “end” acquires a slightly negative charge because of its greater electronegativity while the hydrogen “ends” acquire a slightly positive charge.

The positive region of one water molecule attracts the negative region of another water molecule.

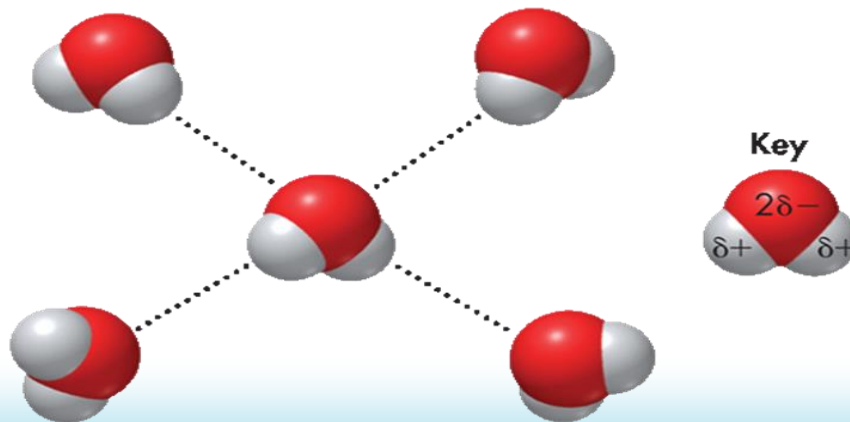
Hydrogen bonding in molecules like HF, NH₃, and H₂O causes unique and incredible properties.



When water freezes to ice a large proportion of SPACE is created and therefore, the DENSITY of ice is less than the density of liquid water (ICE FLOATS). All other solids sink in their own liquid.

Water has the second highest specific heat for liquids, meaning that water resists changes in temperature. (e.g. lakes, ponds, ocean take a long time to freeze and they do not freeze solid.)

Water also has the second highest Surface Tension, causing cohesion (“sticking to itself”), adhesion (“sticking” to a different substances), and capillary action (rising up a narrow tube).



Network Solids

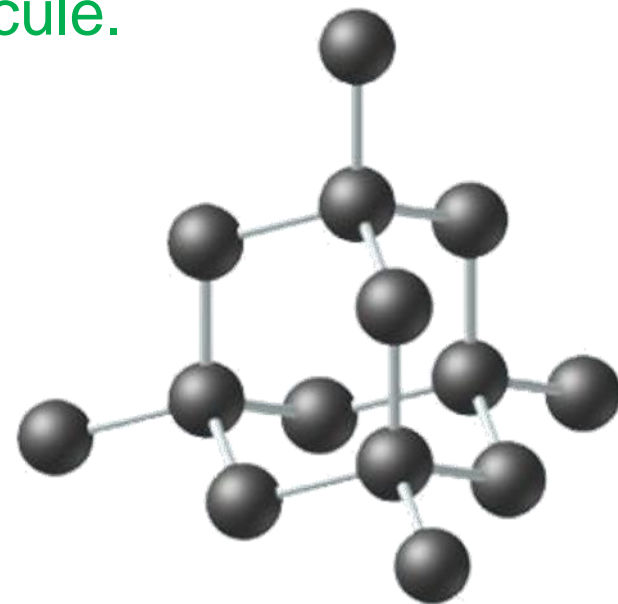
Large, stable molecules forming a “**network**” of **covalent bonds** that extends throughout the solid. Also called network crystals. *Covalent molecules are usually liquid or gas.*

high melting points ($> 1000\text{ C}$) due to a high strength of covalent bonds throughout the huge molecule.

Insoluble in all common solvents \rightarrow to form a liquid, most of the covalent bonds in the large crystal would need to be broken.

Poor conductors of heat and electricity

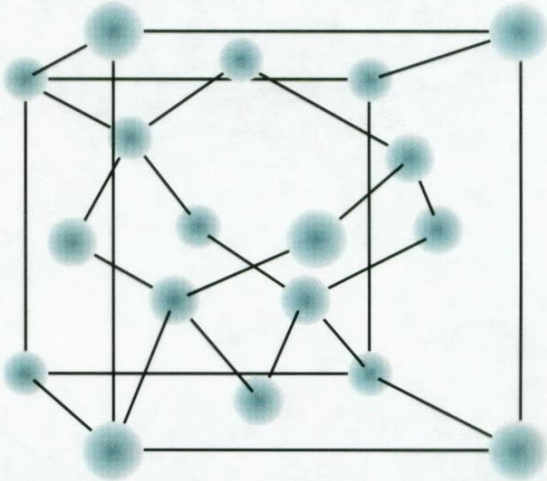
Hard crystals



8. 4 Polar Bonds & Molecules

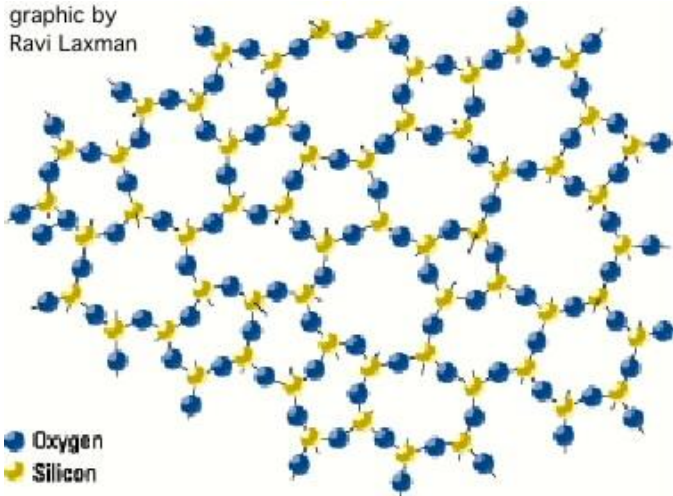
> Network Solids

Diamond → cutting tools, grindstones, jewelry



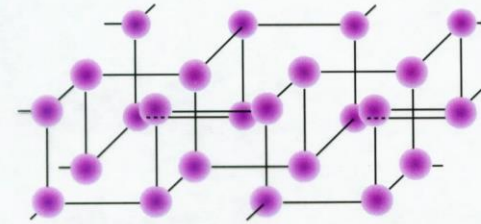
Diamond

graphic by Ravi Laxman

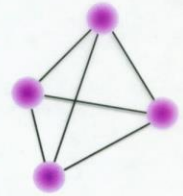


● Oxygen
● Silicon

Silicon Dioxide (SO₂) [silica] → Beach sand, glass, polished glass lenses in eyeglasses and optical instruments

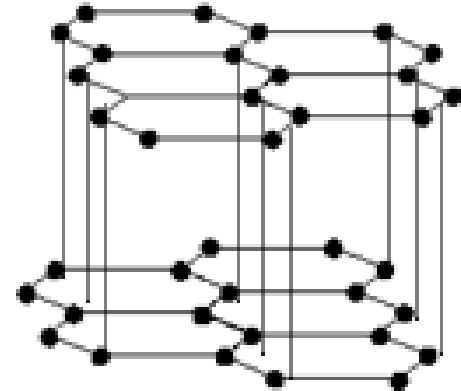


Black phosphorus



White phosphorus

Phosphorus → soft drinks, cleaning agents, water softening, fireworks, optical instruments



Graphite → paints (versus lead), lubricants & oils, pencils

Hydrogen bonds are strongest between the molecules of:

- a) HF (l) b) HCl (l) c) HBr (l) d) HI (l)

The attraction that causes iodine to be a solid when chlorine (another halogen) is a gas is due to:

- a) hydrogen bonds c) ionic bonds
b) covalent bonds d) Van der Waals forces

Which compound or molecule forms large crystals, is a poor conductor of heat & electricity, has a melting point over 1000 C, and is very hard?

- a) CO₂ b) diamond c) H₂O d) N₂O

Gives water special properties like LOWER density as a solid, surface tension, and high specific heat.

- a) hydrogen bonds c) ionic bonds
b) covalent bonds d) Van der Waals forces

Hydrogen bonds are strongest between the molecules of:

- a) HF (l) b) HCl (l) c) HBr (l) d) HI (l)

The attraction that causes iodine to be a solid when chlorine (another halogen) is a gas is due to:

- a) hydrogen bonds c) ionic bonds
b) covalent bonds d) Van der Waals forces

Which compound or molecule forms large crystals, is a poor conductor of heat & electricity, has a melting point over 1000 C, and is very hard?

- a) CO₂ b) diamond c) H₂O d) N₂O

Gives water special properties like LOWER density as a solid, surface tension, and high specific heat.

- a) hydrogen bonds c) ionic bonds
b) covalent bonds d) Van der Waals forces

Properties of Covalent Compounds

- Covalent bonds are strong and do not break easily
- Covalent compounds tend not to have strong intermolecular attractions
 - Intermolecular forces exist between molecules, not between atoms within a molecule
- As a result of these structural features, covalent compounds have these properties:
 - Do not ionize in solution
 - Poor conductors of electric charge
 - Poor conductors of heat
 - Low melting and boiling points



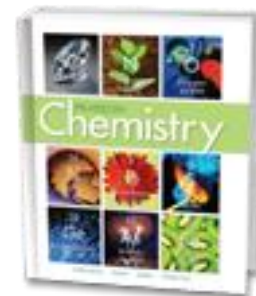
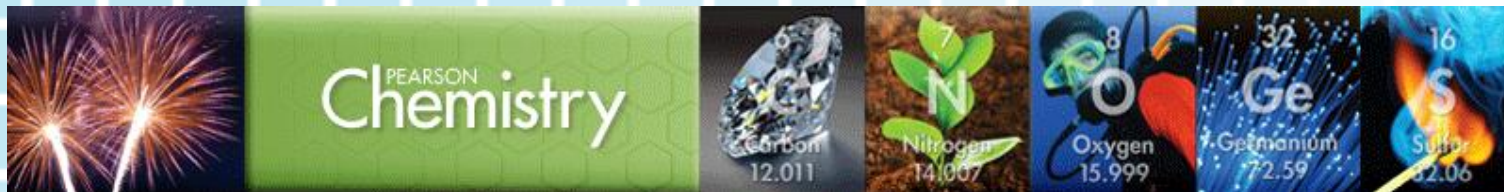
8.1 Molecular Compounds >

Intermolecular Attractions and Molecular Properties

This table summarizes some of the characteristic differences between ionic and covalent (molecular) substances.

Characteristics of Ionic and Molecular Compounds

Characteristic	Ionic Compound	Molecular Compound
Representative unit	Formula unit	Molecule
Bond formation	Transfer of one or more electrons between atoms	Sharing of electron parts between atoms
Type of elements	Metallic and nonmetallic	Nonmetallic
Physical state	Solid	Solid, liquid, or gas
Melting point	High (usually above 300° C)	High (usually below 300° C)
Solubility in water	Usually high	High to low
Electrical conductivity of aqueous solution	Good conductor	Poor to nonconducting



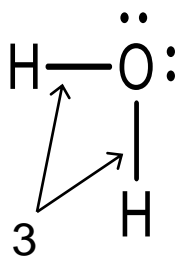
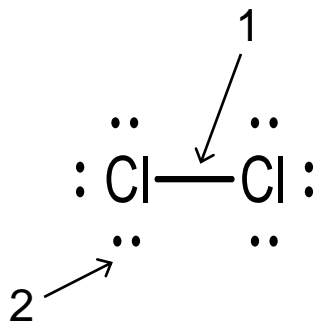
INTERmolecular Attractions: LDF, H-bonds, VDW,
Molecule-ion, Network Solids, Metallic Bonds (7:05)

<https://screencast-o-matic.com/watch/cF61rpYDm5>

Classify Electron Roles

TRY IT

Match each label below with the appropriate term or terms.



1:

- nonbonding electrons
- sigma bond
- represents two electrons

2:

- nonbonding electrons
- sigma bond
- represents core electrons

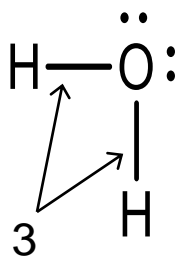
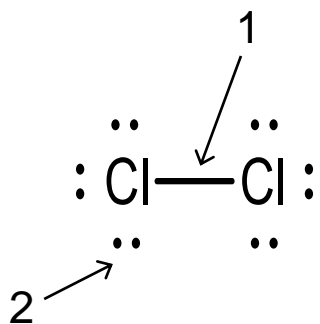
3:

- nonbonding electrons
- bonding electrons
- sigma bond

Classify Electron Roles

TRY IT

Match each label below with the appropriate term or terms.



1:

nonbonding electrons

sigma bond

represents two electrons

2:

nonbonding electrons

sigma bond

represents core electrons

3:

nonbonding electrons

bonding electrons

sigma bond

Drawing Lewis Structures of Molecules

The logo consists of a dark grey rounded rectangle with a white border. Inside, there are two horizontal orange bars, one above and one below the text "TRY IT" in a bold, white, sans-serif font.

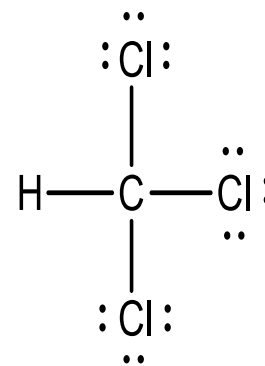
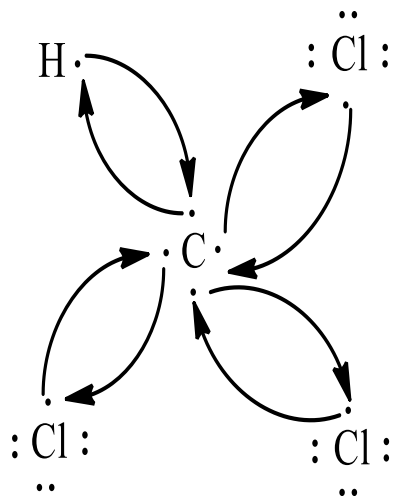
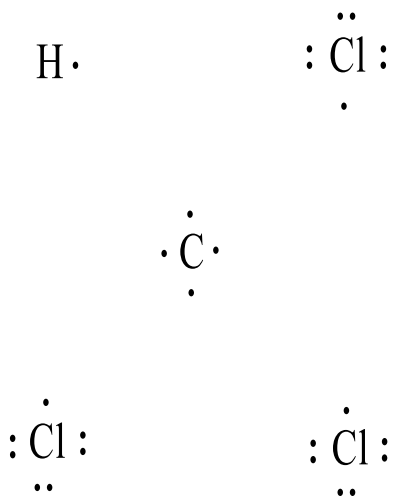
Draw the Lewis structure of CHCl_3 :

1. Draw the Lewis symbols of the atoms in the structure.
2. Form covalent bonds between atoms.
3. Draw the final structure with dashes to represent bonds.

Drawing Lewis Structures of Molecules

TRY IT

Draw the Lewis structure of CHCl_3 :



1. Draw the Lewis symbols of the atoms in the structure.

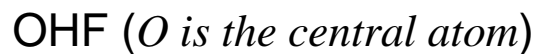
2. Form covalent bonds between atoms.

3. Draw the final structure with dashes to represent bonds.

Practice Drawing Lewis Structures of Simple Molecules



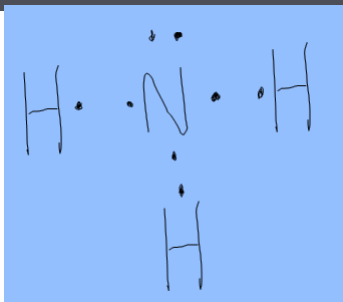
On a sheet of paper, draw the Lewis structures for the following compounds.



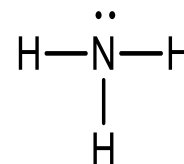
Practice Drawing Lewis Structures of Simple Molecules



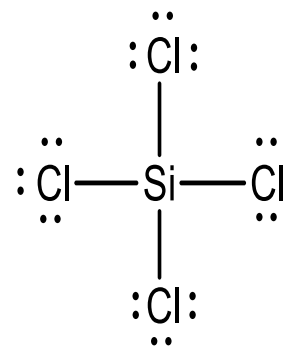
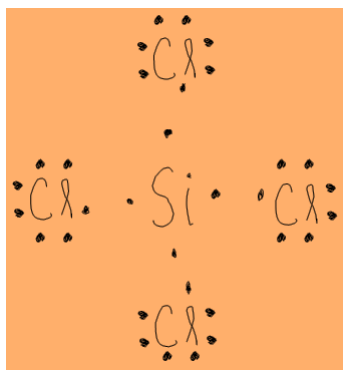
NH₃



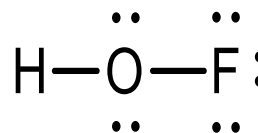
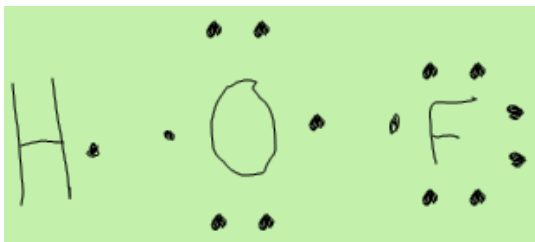
Do your structures resemble the following?



SiCl₄

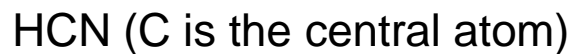
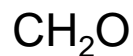


OHF

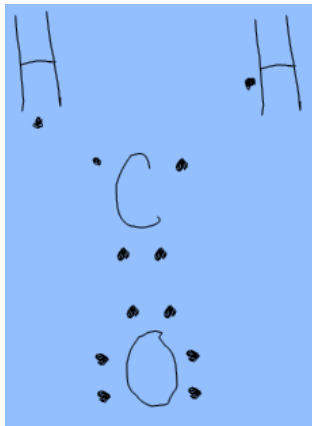
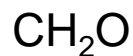


Practice Drawing Lewis Structures of Molecules Containing Multiple Bonds

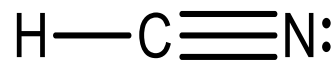
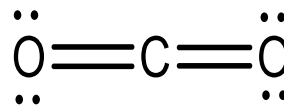
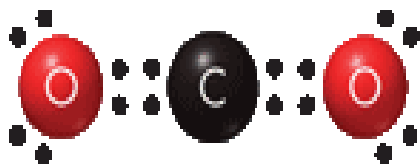
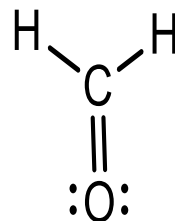
Draw the Lewis structures for the following compounds.



Practice Drawing Lewis Structures of Molecules Containing Multiple Bonds



Do your structures resemble the following?



Period	<i>s</i> -block	
	1 IA	
1	1.00794 1 H 1s ¹	+1 -1

KEY

Atomic Mass → 12.0111

Symbol → **C**

Atomic Number → 6

Electron Configuration → 1s²2s²2p²

Relative atomic masses are based on ¹²C = 12.00000

Selected Oxidation States → -4, +2, +4

s-block
GROUP

1 IA 2 IIA

New Designation

Former Designation
(prior to 1984 IUPAC decision)

	<i>s</i> -block		<i>d</i> -block									
	Transition Elements		GROUP									
	1 IA	2 IIA	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIII	9 VIII	10 VIII		
2	3 Li 1s ² 2s ¹	4 Be 1s ² 2s ²										
3	11 Na [Ne]3s ¹	12 Mg [Ne]3s ²										
4	19 K [Ar]4s ¹	20 Ca [Ar]4s ²	21 Sc [Ar]3d ¹ 4s ²	22 Ti [Ar]3d ² 4s ²	23 V [Ar]3d ³ 4s ²	24 Cr [Ar]3d ⁵ 4s ¹	25 Mn [Ar]3d ⁵ 4s ²	26 Fe [Ar]3d ⁶ 4s ²	27 Co [Ar]3d ⁷ 4s ²	28 Ni [Ar]3d ⁸ 4s ²	29 Cu [Ar]3d ¹⁰ 4s ¹	
5	37 Rb [Kr]5s ¹	38 Sr [Kr]5s ²	39 Y [Kr]4d ¹ 5s ²	40 Zr [Kr]4d ² 5s ²	41 Nb [Kr]4d ⁴ 5s ¹	42 Mo [Kr]4d ⁵ 5s ¹	43 Tc [Kr]4d ⁵ 5s ¹	44 Ru [Kr]4d ⁷ 5s ¹	45 Rh [Kr]4d ⁸ 5s ¹	46 Pd [Kr]4d ¹⁰ 5s ⁰	47 Ag [Kr]4d ¹⁰ 5s ¹	
6	55 Cs [Xe]6s ¹	56 Ba [Xe]6s ²	57 La-Lu	72 Hf [Xe]4f ¹⁴ 5d ² 6s ²	73 Ta [Xe]4f ¹⁴ 5d ³ 6s ²	74 W [Xe]4f ¹⁴ 5d ⁴ 6s ²	75 Re [Xe]4f ¹⁴ 5d ⁵ 6s ²	76 Os [Xe]4f ¹⁴ 5d ⁶ 6s ²	77 Ir [Xe]4f ¹⁴ 5d ⁷ 6s ²	78 Pt [Xe]4f ¹⁴ 5d ⁹ 6s ¹	79 Au [Xe]4f ¹⁴ 5d ¹⁰ 6s ¹	
7	(223) Fr [Rn]7s ¹	226.025 Ra [Rn]7s ²	89 Ac-Lr	(261) Unq*	(262) Unp	(263) Unh	(262) Uns	(262) Uno	(262) Une			

* The sys
103 wil

masses are
2.00000

s-block
18
0

ation States

4.00260	0
He	
2	
1s ²	

p-block
GROUP

			13 IIIA	14 IVA	15 VA	16 VIA	17 VIIA	18 0			
			10.81 B 5 1s ² 2s ² 2p ¹	12.0111 C 6 1s ² 2s ² 2p ²	14.0067 N 7 1s ² 2s ² 2p ³	15.9994 O 8 1s ² 2s ² 2p ⁴	18.998403 F 9 1s ² 2s ² 2p ⁵	20.179 Ne 10 1s ² 2s ² 2p ⁶			
			26.98154 Al 13 [Ne]3s ² 3p ¹	28.0855 Si 14 [Ne]3s ² 3p ²	30.97376 P 15 [Ne]3s ² 3p ³	32.06 S 16 [Ne]3s ² 3p ⁴	35.453 Cl 17 [Ne]3s ² 3p ⁵	39.948 Ar 18 [Ne]3s ² 3p ⁶			
10	11 IB	12 IIB	58.69 Ni 28 [Ar]3d ⁸ 4s ²	63.546 Cu 29 [Ar]3d ¹⁰ 4s ¹	65.39 Zn 30 [Ar]3d ¹⁰ 4s ²	69.72 Ga 31 [Ar]3d ¹⁰ 4s ² 4p ¹	72.59 Ge 32 [Ar]3d ¹⁰ 4s ² 4p ²	74.9216 As 33 [Ar]3d ¹⁰ 4s ² 4p ³	78.96 Se 34 [Ar]3d ¹⁰ 4s ² 4p ⁴	79.904 Br 35 [Ar]3d ¹⁰ 4s ² 4p ⁵	83.80 Kr 36 [Ar]3d ¹⁰ 4s ² 4p ⁶
106.42 Pd 46 [Kr]4d ¹⁰ 5s ⁰	107.868 Ag 47 [Kr]4d ¹⁰ 5s ¹	112.41 Cd 48 [Kr]4d ¹⁰ 5s ²	114.82 In 49 [Kr]4d ¹⁰ 5s ² 5p ¹	118.71 Sn 50 [Kr]4d ¹⁰ 5s ² 5p ²	121.75 Sb 51 [Kr]4d ¹⁰ 5s ² 5p ³	127.60 Te 52 [Kr]4d ¹⁰ 5s ² 5p ⁴	126.905 I 53 [Kr]4d ¹⁰ 5s ² 5p ⁵	131.29 Xe 54 [Kr]4d ¹⁰ 5s ² 5p ⁶			
195.08 Pt 78 [Xe]4f ¹⁴ 5d ⁹ 6s ¹	196.967 Au 79 [Xe]4f ¹⁴ 5d ¹⁰ 6s ¹	200.59 Hg 80 [Xe]4f ¹⁴ 5d ¹⁰ 6s ²	204.383 Tl 81 [Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ¹	207.2 Pb 82 [Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ²	208.980 Bi 83 [Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ³	(209) Po 84 [Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ⁴	(210) At 85 [Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ⁵	(222) Rn 86 [Xe]4f ¹⁴ 5d ¹⁰ 6s ² 6p ⁶			

Polyatomic Ions

Name	Formula	Name	Formula
perPhosphate	$(\text{PO}_5)^{-3}$	perCarbonate	$(\text{CO}_4)^{-2}$
Phosphate	$(\text{PO}_4)^{-3}$	Carbonate	$(\text{CO}_3)^{-2}$
Phosphite	$(\text{PO}_3)^{-3}$	Carbonite	$(\text{CO}_2)^{-2}$
hypoPhosphite	$(\text{PO}_2)^{-3}$	hypocarbonite	$(\text{CO})^{-2}$
perChlorate	$(\text{ClO}_4)^{-1}$	perNitrate	$(\text{NO}_4)^{-}$
Chlorate	$(\text{ClO}_3)^{-1}$	Nitrate	$(\text{NO}_3)^{-}$
Chlorite	$(\text{ClO}_2)^{-1}$	Nitrite	$(\text{NO}_2)^{-}$
hypoChlorite	$(\text{ClO})^{-1}$	Hyponitrite	$(\text{NO})^{-}$
perSulfate	$(\text{SO}_5)^{-2}$	perChromate	$(\text{CrO}_5)^{-2}$
Sulfate	$(\text{SO}_4)^{-2}$	Chromate	$(\text{CrO}_4)^{-2}$
Sulfite	$(\text{SO}_3)^{-2}$	Chromite	$(\text{CrO}_3)^{-2}$
hyposulfite	$(\text{SO}_2)^{-2}$	Hypochromite	$(\text{CrO}_2)^{-2}$
Acetate	$(\text{C}_2\text{H}_3\text{O}_2)^{-1}$	Cyanide	$(\text{CN})^{-1}$
Hydroxide	$(\text{OH})^{-1}$	Manganate	$(\text{MnO}_4)^{-2}$

Ammonium $(\text{NH}_4)^{+1}$
