Dr. Zellmer | Chemistry 1210 | All sections
Time: 7 PM Sun. | Autumn Semester 2022 | November 6, 2022
40 min | Quiz IX

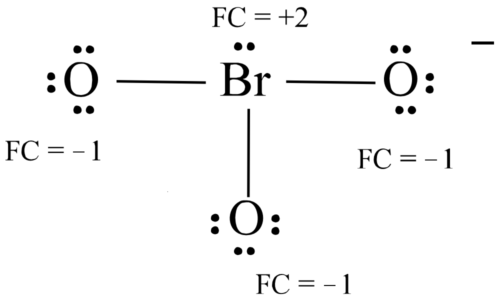
Name ______________KEY__________________ Rec. TA/time ___________________________________

1. (5 pts) Draw the Lewis structure of the bromate ion, BrO_3^- conforming to the Lewis octet (noble gas) rule, **and** put the formal charges on each atom. **(Must show work or explain what you are doing and show and account for all valence electrons and formal charges.)**

a) Lewis structure of BrO_3^- ; Br: Grp 7A (7 val e^-) O: Grp 6A (6 val e^-)

Br O -1 chg
1) A = $1(7e^-) + 3(6e^-) + 1\ e^- = 26\ e^-$ available

2) Draw skeleton structure. The more EN O atoms attached to Br This accounts for 6 e^-.



3) Put 6 e^- on each O to fulfill octet. This uses 18 e^-.

4) # e^- left = $26\ e^- - (6\ e^- + 18\ e^-) = 2\ e^-$
These go on the Br atom to fulfill octet.

NOT trigonal planar, tetrahedral or flat T- shaped

5) Formal Charge:

Divide e^- in bonds equally between atoms (each atom gets ½ the e^- involved in the bond).

lpe^- assigned to the atom they're on.

Subtract this total from the valence e^- on the atom.

Ex: For Br: 7 - (3 + 2) = +2
For O: 6 - (1 + 6) = −1

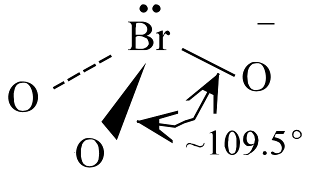
FC should add up to give overall chg., in this case −1 (as they do)

***** The following was not asked for *****

b) What is the shape of this species (polyatomic ion)?

4 e^- pairs (3 atoms & 1 lpe^-) on Br
All 4 e^- pairs are arranged in a tetrahedron
so angles start out close to 109.5°

- describe molecular shape by considering **only atoms** an **NOT lone pair** e^-



trigonal pyramidal

c) What is the O-Br-O bond angle?

< 109.5° or ~ 109.5° (but **NOT exactly** 109.5°)

The e^- pairs (lpe^-) on Br causes the angles to be slightly diff. than the 109.5° found in a perfect tetrahedral molecule (4 identical surrounding atoms).
For 4 "things" (atoms & lpe^-) around a central atom, if all 4 things are **NOT identical** then angles are < 109.5° or ~ 109.5° (but NOT exactly 109.5°).

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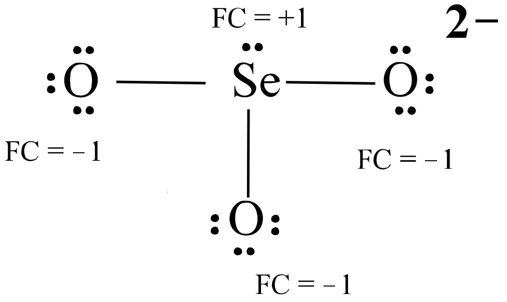
2. (5 pts) Which of the following is the best Lewis structure, conforming to Octet rule, for SeO_3^{2-}?

a) Lewis structure of SeO_3^{2-} ; Se: Grp 6A (6 val e^-) O: Grp 6A (6 val e^-)

1) $A = \overset{Se}{1(6e^-)} + \overset{O}{3(6e^-)} + \overset{-2\ chg}{2\ e^-} = 26\ e^-$ available

2) Draw skeleton structure. The more EN O atoms attached to Se This accounts for 6 e^-.



NOT trigonal planar, tetrahedral or flat T- shaped

3) Put 6 e^- on each O to fulfill octet. This uses 18 e^-.

4) # e^- left = 26 e^- - (6 e^- + 18 e^-) = 2 e^-
These go on the Se atom to fulfill octet.

5) Formal Charge:

Divide e^- in bonds equally between atoms (each atom gets ½ the e^- involved in the bond).

lpe^- assigned to the atom they're on.

Subtract this total from the valence e^- on the atom.

Ex: For Se: 6 - (3 + 2) = +1
For O: 6 - (1 + 6) = −1

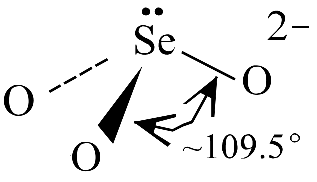
FC should add up to give overall chg., in this case −2 (as they do)

***** The following was not asked for *****

b) What is the shape of this species (polyatomic ion)?

4 e^- pairs (3 atoms & 1 lpe^-) on Se
All 4 e^- pairs are arranged in a tetrahedron
so angles start out close to 109.5°

- describe molecular shape by considering **<u>only</u> atoms** an **<u>NOT</u> lone pair e^-**



trigonal pyramidal

c) What is the O-Se-O bond angle?

< 109.5° or ~ 109.5° (but **<u>NOT</u> exactly** 109.5°)

The e^- pair (lpe^-) on Se causes the angles to be slightly diff. than the 109.5° found in a perfect tetrahedral molecule (4 identical surrounding atoms).
For 4 "things" (atoms & lpe^-) around a central atom, if all 4 things are **<u>NOT</u> identical** then angles are < 109.5° or ~ 109.5° (but NOT exactly 109.5°).

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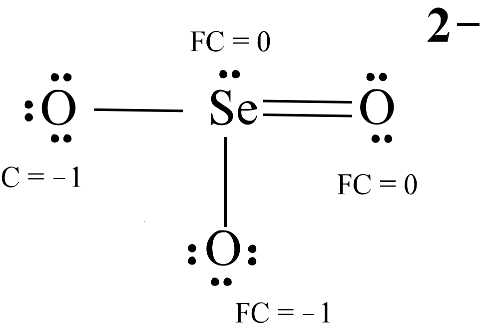
3. (5 pts) Draw the Lewis structure of the selenite ion, SeO_3^{2-}, conforming to the Formal Charge rules, **and** put the formal charges on each atom. **(Must show work or explain what you are doing and show and account for all valence electrons and formal charges.)**

Use the work done above in #6 to come up with this structure by remembering that from row 3 down atoms can get more than an octet around them, like I showed for SO_4^{2-} in lecture.



Still trigonal pyramidal

Atoms in row 3 and below can have more than an octet. Thus, the Se atom can form another bond to lower the formal charges. The above structure is one of three resonance structures with a double bond between the Se and one of the O atoms. The double bond appears to "move" between O atoms and the Se atom. However, remember for resonance structures none of them actually exists. Instead the actual structure is an average of all the resonance structures.

Based on formal charge rules, this structure would be better than the one in question 6. In the structure in this question, the formal charges are lower in all than those in question 6. In this structure only two atoms have formal charges that are nonzero, whereas the one in question 3 all atoms have nonzero formal charges. Also, the sum of the absolute values of the formal charges in the above structure is 2, which is lower than that in question 6, which is 4.

Actually, there are 4 resonance structures altogether, the 3 which look like the one above with the double bond in the different Se-O positions and the one above in question 6 which conforms to the octet rule.

Remember, there is disagreement on whether the octet rule or formal charge rules should take precedence. We would ask you to draw the structure conforming to one or the other rule, like I did in question 6 and this question.

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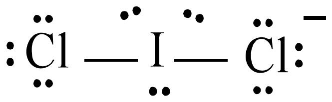
4. (5 pts) Draw the Lewis structure for ICl_2^-. How many lone pair(s) of electrons are there in the valence shell of the central atom?

I: 7A Cl: 7A −1 chg
I Cl 1 e^-
1) A = 1(7e-) + 2(7e-) + (1e-) = 22 val e^-

2) Draw skeleton structure. The more EN Cl atoms attached to I This accounts for 4 e^-.



Iodine is in the middle

3) Put 6 e^- on each Cl to fulfill octet. This uses 12 e^-.

4) # e^- left = 22 e^- - (4 e^- + 12 e^-) = 6 e^-
These go on the I atom. This gives 10 e- on the I atom, more than an octet.

I has 2 bonds & 3 lpe$^-$ in this case since it is in the middle (when halogens are on the outside they have only 1 bond)

The ED geometry around the I is trigonal bipyramidal and the molecular geometry is linear with a bond angle of 180°

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5. (5 pts) Draw the resonance structures for the nitrite ion, NO_2^-. These indicate,

A. An electron pair in the molecule alternates back and forth between the two nitrogen-oxygen bonds so that the two different bonds seem to exchange positions.

B. The two bonds in the nitrite ion are of equal length and the electron distribution in the two nitrogen-oxygen bonds is identical with an electron pair being shared by all three atoms to give a bond order of 1.5.

C. The electron distribution in the two nitrogen-oxygen bonds differs, as do their bond lengths.

D. The molecule revolves around an axis through the nitrogen atom between the two oxygen atoms so the two different nitrogen-oxygen bonds seem to exchange positions.

E. The nitrite ion exists as two different molecules which simultaneously exist.

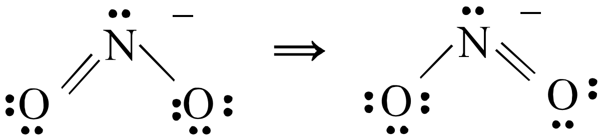
N (5A) O (6A) −1 chg
1) A = 1(5e−) + 2(6e−) + (1e−) = 18 val e−

2) Draw skeleton structure. The more EN O atoms attached to N. This accounts for 4 e^-.

3) Put 6 e^- on each O to fulfill octet. This uses 12 e^-.

4) # e^- left = 18 e^- - (4 e^- + 12 e^-) = 2 e^-

2 more e- to put on N. However, this still would not give an octet on N. Must move a pair from an O atom to form double bond between N and O. There are 2 oxygen atoms for which you could do this. Thus the double bond could be drawn in 2 different positions leading to 2 resonance structures.



2 resonance structures (double bond can be in 1 of the 2 different positions)
Bent, angles of **approx.** 120°

To fulfill the octet rule on nitrogen there has to be a double bond between the N and an O atom. The double bond could be drawn in 2 different positions leading to 2 resonance structures. The double bond isn't moving or flipping back and forth between the two regions. There aren't two separate molecules. The two bonds are identical (have the same bond length and electron distribution). There's simply no easy way to draw one single Lewis Structure for this molecule. What it represents is that the "extra" pair of electrons represented in the double bond is actually being shared equally in both regions (actually shared by all three atoms in a molecular orbital which spans all three atoms). It's like the "extra" bond is shared equally to give a bond order of 1.5.

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6. (8 pts) Draw the all the possible resonance structures of the cyanate ion, NCO^-, conforming to the Lewis octet rule **and** put the **formal charges** on each atom. **Also**, **indicate** which would likely be the **dominate structure**. **(Must show work or explain what you are doing and show and account for all valence electrons and formal charges. Also, explain your reasoning for your choice of the dominate structure.)**

a) Lewis structure of NCO^- ; N: Grp 5A (5 val e^-) C: Grp 4A (4 val e^-) O: Grp 6A (6 val e^-)

N C O -1 chg

1) $A = 1(5e^-) + 1(4e^-) + 1(6e^-) + 1\ e^- = 16\ e^-$ available

2) Draw skeleton structure. The more EN N & O atoms attached to C.
This accounts for 4 e^-. This will leave 12 e^-.

N - C - O

3) Put 6 e^- on N & O to fulfill octet. This uses 12 e^-.

$:\underset{\cdot\cdot}{\overset{\cdot\cdot}{N}} - C - \underset{\cdot\cdot}{\overset{\cdot\cdot}{O}}:$

4) $\#\ e^-$ left $= 16\ e^- - (4\ e^- + 12\ e^-) = 0\ e^-$

The C doesn't have an octet (C normally forms 4 bonds).
Need to "take" e^- from N or O to form multiple bonds.
Get resonance structures.

5) Formal Charge:

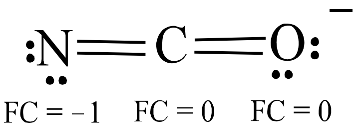
Divide e^- in bonds equally between atoms
(each atom gets ½ the e^- involved in the bond).

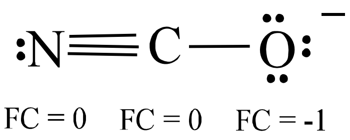
lpe$^-$ assigned to the atom they're on.

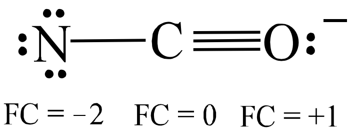
Subtract this total from the valence e^- on the atom.

Ex: For N: $5 - (2 + 4) = -1$
For C: $4 - (4 + 0) = 0$
For O: $6 - (2 + 4) = 0$

FC should add up to give overall chg., in this case −1 (as they do)







The **second structure** would be the **dominate one** in this case since the neg. FC is on the more EN O atom. The third structure would not contribute much since the sum of the absolute values of the FC is the largest and the N atom has a -2 and the O has a +1, which would not normally be the case since O is more EN.

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7. (3 pts) Which of the following does **<u>NOT</u>** conform to the Lewis "octet" (noble gas) rule?

1) NO_2^- 2) PF_5 3) HCN 4) BCl_3 5) CO

PF_5 and BCl_3

There are several instances when the octet rule is not obeyed.

There are cases when there is less than an octet and there's an even number of electrons. This occurs for **Be** (group 2A) and **B** and **Al** (group 3A).

Be gets only 4 e– around it and forms 2 bonds (group 2A)

B and Al get only 6 e– around them and form 3 bonds (group 3A)

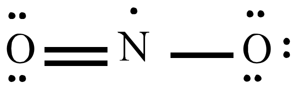
•Be • $\overset{\bullet}{\bullet\text{B}\bullet}$ (Al like B since they are in the same group, 3A)

2 bonds 3 bonds
4 e– 6 e–

There's also the cases when an atom can get more than an octet of electrons. This occurs for atoms from the 3rd row down. It does NOT occur for atoms in row 2. Examples are PF_5 (10 e–), SF_4 (10 e–), ClF_3 (10 e–), XeF_2 (10 e–), SF_6 (12 e–), BrF_5 (12 e–), XeF_4 (12 e–) and others.

There are some systems with an odd # of electrons. These will not get an octet on one of the atoms and there will be an unpaired electron. This occurs most often with Nitrogen containing compounds (group 5A).

> NO_2 has 17 valence electrons and it winds up with an unpaired e– in the molecule and it's usually shown on the N atom.



While H does not get an "octet" it does get 2 electrons around it and looks like He, a noble gas. That's why in lecture I called the "octet" rule the "noble-gas" rule. Hydrogen gets a complete outer shell like He when it forms bonds. That's what the "octet" rule is about, getting a complete outer shell like a noble gas. That is what occurs for most atoms but H can't get 8 electrons around it, just 2 electrons like the noble gas He, but it does get a complete outer shell.

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8. (3 pts) Of the possible bonds between nitrogen atoms (single, double, and triple), this of the following are ture? Multiple answers possible.

a) a triple bond is longer than a single bond
b) a double bond is stronger than a triple bond
c) a single bond is stronger than a triple bond
d) a double bond is longer than a triple bond
e) a single bond is stronger than a double bond
f) a triple bond is stronger than a double bond

Both D and F are true. When the same two atoms are involved the following is true for the bond lengths (in order of decreasing length) and strengths (in order of increasing strength) :

single < double < triple

double bonds are shorter than single bonds and triple bonds are shorter than double bonds

double bonds are stronger than single bonds and triple bonds are stronger than double bonds

Think of bonds like springs. A double bond is a shorter fatter spring than a single bond and thus it takes more energy to break a double bond. A triple bond is a shorter fatter spring than a double bond and thus it takes more energy to break a triple bond.

9. (3 pts) Which of the following bonds is most likely to be the **shortest**?

a) $C = C$ b) $C = N$ c) $O = O$ d) $C - C$ **e) $C \equiv C$**

While it can be difficult to compare bond lengths and strengths when different atoms are involved, generally speaking triple bonds are shorter and stronger than double bonds.

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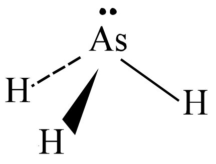
10. (6 pts) Consider the following molecules and list their **molecular shapes** (NOT the electron domain geometries), **bond angles** and whether they are **polar or nonpolar**. (**Provide the Lewis structure and a short explanation for your choices.**)

1) AsH_3

trigonal pyramidal (≈ 109.5° angles)
4 things around As and not all identical
(3 H and **1 lpe**$^-$)

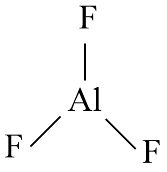
Polar (like **:**NH_3,**:**PH_3, **:**PCl_3, etc.)



2) AlF_3

trigonal planar (120° angles, exactly)
3 things around Al and **all identical**
(3 atoms and **NO lpe**$^-$)

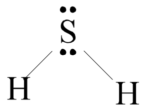
Nonpolar (like BCl_3, BF_3 $AlCl_3$ etc.)



3) H_2S

bent: 4 things around S and not all identical (2 H and **2 lpe**$^-$)
lpe$^-$ on O (≈ 109.5° angles)

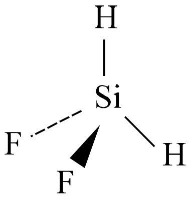
Polar (like $H_2\ddot{\text{O}}$**:**, $H_2\ddot{\text{S}}$**:**, etc.)



4) SiH_2F_2

tetrahedral (≈109.5° angles)
all 4 atoms on Si are **not** identical

Polar (like $CHCl_3$, CH_3Cl, CH_2Cl_2, ect.)



NOTE: The **5 basic symmetric** molecular **geometries** (linear, trigonal planar, tetrahedral, trigonal bipyramidal and octahedral) can give **non**polar if all *surrounding* **atoms** (or groups of atoms) are **identical** to each other (although they don't all have to be identical for the trigonal bipyramidal, octahedral or square planar molecular geometries.

See Tables 9.2 and 9.3 on pages 350 and 353 in the textbook.

***** continued on the next page *****

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10. (Cont.)

* **NOTE**: If there are 1 or more **lpe**$^-$ on the central atom the molecule will **generally** be **polar**.

There are two shapes that can be exceptions to this 'rule'. (See Table 9.3 on page 353 of the textbook.)

The linear shape resulting from the trigonal bipyramidal e$^-$-pair geometry will be nonpolar if the 2 atoms around the central atom are identical to each other (e.g. XeF_2 has 3 lpe$^-$ in a trigonal planar geometry around the Xe and the 2 F atoms linear to each other).

The square planar shape resulting from the octahedral e$^-$-pair geometry will be nonpolar if the 4 atoms around the central atom are identical to each other (e.g. XeF_4 has 4 F atoms in a square planar arrangement with 2 lpe$^-$ linear to each other, 1 above and 1 below the square planar arrangement).

For AlF_3 (#2) it's trigonal planar & **all atoms** *surrounding* Al are **identical** so the molecule is **non**polar.

H_2S and AsH_3 (3 & 1) neutral molecules which are bent or trigonal pyramidal with **lpe**$^-$ on the central atom are **polar**. In general, molecules that have **<u>1</u>** or **<u>more</u>** **<u>lpe</u>**$^-$ on the **<u>central</u>** **<u>atom</u>** are **<u>polar</u>**. (See discussion above about exceptions.)

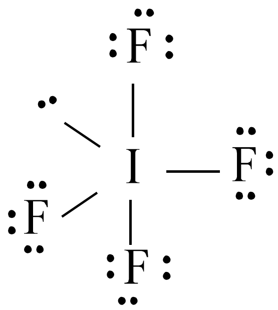
For SiH_2F_2 (#4) it is tetrahedral but all surrounding atoms are **<u>NOT</u>** **identical** so the molecule is **<u>polar</u>**.

11. (6 pts) Draw the Lewis structure of IF_4^+. What is its electron-domain geometry? What is its molecular geometry? What are the bond angles? (**Show work or explain.**)

I: 7A F: 7A
I F

1) A = 1(7e-) + 4(7e-) - (1e-) = 34 val e−

2) Draw skeleton structure. The more EN F atoms attached to I. This accounts for 10 e$^-$.



Iodine is in the middle

3) Put 6 e$^-$ on each F to fulfill octet. This uses 24 e$^-$.

4) # e$^-$ left = 34 e$^-$ - (8 e$^-$ + 24 e$^-$) = 2 e$^-$
These go on the I atom. This gives 10 e- on the I atom, more than an octet.

I has 4 bonds & 1 lpe$^-$ in this case since it is in the middle (when halogens are on the outside they have only 1 bond, F is always on the outside, never in the middle)

The **ED** geometry around the I is **trigonal bipyramidal**. There are **5 electron domains** around the Br with angles of 90° and 120°.

The **molecular geometry** is **seesaw** with **bond angles** of ~ **90° and ~ 120°**. The electron pairs push the F atoms closer together so bond angles are less than or approx. 90° and 120°.

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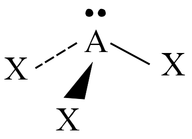
12. (4 pts) Which of the following species is **<u>trigonal</u> <u>pyramidal</u>**?

1) SO_3^{2-} 2) NO_2^- 3) CO_3^{2-} 4) BCl_3 5) NF_3

a) 3 b) 4, 5 c) 2, 5 d)* 1, 5 e) 1, 3, 5

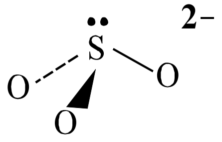
In order for a molecule to be **trigonal pyramidal** there has to be **3 atoms** and **1 lpe$^-$** around the central atom.



trigonal pyramidal: 3 atoms and 1 lpe$^-$ surrounding A
angle < 109.5° or ≈109.5°

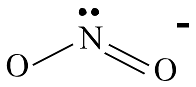
*1) SO_3^{2-}

3 atoms and 1 lpe$^-$ on S (4 things total)
trigonal pyramidal
angle ≈ 109.5°
(neither polar nor nonpolar since it's an ion)



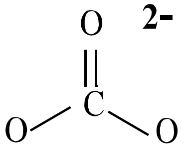
2) NO_2^-

2 atoms and 1 lpe$^-$ on S (3 things total)
Bent
angle ≈ 120 °
(neither polar nor nonpolar since it's an ion)



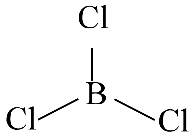
3) CO_3^{2-}

3 atoms and no lpe$^-$ on C (3 things total)
trigonal planar
angle = 120° (exactly)
(neither polar nor nonpolar since it's an ion)



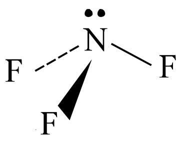
4) BCl_3

3 atoms and no lpe$^-$ on B (3 things total)
trigonal planar
angle = 120° (exactly)
nonpolar



*5) NF_3

3 atoms and 1 lpe$^-$ on N (3 things total)
trigonal pyramidal
angle≈ 109.5 °
polar



1 & 5, SO_3^{2-} and NF_3, are trigonal pyramidal

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13. (6 pts) Consider the following molecules and select those that are **<u>non</u>**polar.

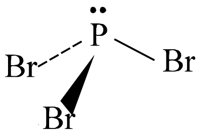
1) PBr_3 2) BH_3 3) H_2S 4) CH_2Cl_2 5) CS_2

a) 1, 2 b) 1, 3 c) 2, 4 d)* 2, 5 e) 2, 4, 5

1) PBr_3

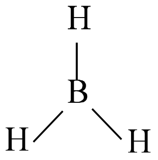
trigonal pyramidal: 4 things around P and they are not all identical (3 Br and **1 lpe**$^-$) lpe$^-$ on P (≈ 109.5° angles)

Polar (like **:**NH_3, **:**NF_3, **:**PH_3, **:**PF_3, etc.)



*2) BH_3

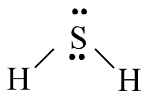
trigonal planar (120° angles, exactly) 3 things around B and **<u>all</u> identical** (and no lpe$^-$)
Nonpolar (like BH_3, BF_3, $AlCl_3$ etc.)



3) H_2S

bent: 4 things around S and not all identical (2 H and **2 lpe**$^-$) lpe$^-$ on O (≈ 109.5° angles)

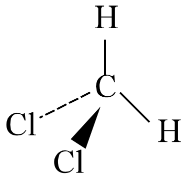
Polar (like $H_2\ddot{O}$**:**, $H_2\ddot{S}$**:**, etc.)



4) CH_2Cl_2

tetrahedral (≈109.5° angles) all 4 atoms on C are **<u>not</u>** identical

Polar (like $CHCl_3$, CH_3Cl, CH_2F_2, ect.)



*5) CS_2

linear (180° angles) the 2 atoms on C are **identical**

Nonpolar (like CO_2)

<u>Note</u>: The symmetric shapes (linear, trigonal planar, tetrahedral, trigonal bipyramidal, octahedral and square planar) are the only ones that can give nonpolar. For linear, trigonal planar and tetrahedral all surrounding atoms have to be identical to each other for the molecule to be nonpolar.

For CH_2Cl_2 (#4) all surrounding atoms are **<u>NOT</u> identical** so the molecule is **<u>polar</u>**.

Also, the molecules (1 & 3) with **lpe**$^-$ are **polar**. Generally, if there are 1 or more lpe$^-$ on the central atom, the molecule will be polar (at least for molecular shapes coming from ED geometries involving 2, 3 or 4 ED). The exceptions to this are for the linear molecular geometry arising from the trigonal bipyramidal ED geometry and the square planar molecular geometry arising from the octahedral ED geometry. These can be nonpolar if all the atoms surrounding the central atom are identical even though there are also **lpe**$^-$ on the central atom.

2 & 5 are Nonpolar

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