Chemistry 1035 Test 4 November 20, 2003
THIS IS FORM C[FORM D #s in ( )]

Name ____________________________

Grades will be posted Friday Morning.
Have a wonderful break!

Choose the single best answer to each question.

1.(4) Which of the following has the same electron configuration as P^{3-} ions?

1. Ti^{4+} 2. Mn^{5+} 3. As^{3-} 4. Ne

Answer: 1. Ti^{4+}

For anions, electrons are added to the electron configuration. For cations, electrons are removed from the highest n orbital first:
P: $[Ne]3s^23p^3$; P^{3-}: **[Ar]** 3 electons are added to the 3p orbitals
Ti: $[Ar]4s^23d^2$; Ti^{4+}: **[Ar]** 2 electrons are removed from the 4s orbital and 2 more from the 3d orbitals.
As: $[Ar]4s^23d^{10}4p^3$; As^{3-}: $[Ar]4s^23d^{10}4p^6 = [Kr]$ 3 electons are added to the 4p orbitals
Ne: $1s^22s^22p^6$

2.(3) Which of the following elements has the lowest ionization energy?

1. Br 2. Al 3. Mg

4. Cs 5. Ni

Answer: 4. Cs

Ionization energy trend:
Increase
→ → Ionization energy increases as you proceed up any group and to the right in
Inc ↑ the periodic table.
↑

Relative position of these elements in the periodic table:
Mg Al
Br
Zr
Cs

Since cesium is in the bottom left of the periodic table, it has the smallest ionization energy..

3.(6) The lattice energy for ionic compounds increases as the charge on the ions _________________ and the size of the ions _________________

1. increases, increases
2. increases, decreases
3. decreases, increases
4. decreases, decreases

Answer: 2. increases, decreases

Two factors contribute to the value of lattice energy for an ionic compound. One is the magnitude of the charge of the ions; the greater the charge of the ions, the greater the value of the lattice energy. Lattice energy increases as the charge on the ions increase.

The other contributing factor is the size (radius) of the ions. The smaller the radius, the closer the ions can approach each other and the higher the value of the lattice energy. Lattice energy increases as the size of the ions decreases.

4.(5) Elements with __________ first ionization energies and ______________ electron affinities generally form anions.

1. low, very negative
2. high, slightly negative
3. low, positive or slightly negative
4. high, very negative

Answer: 4. high, very negative

The elements that form anions (negatively charged ions) are the nonmetals located on the right in the periodic table. Metals form cations (positively charged ions). Nonmetals do not lose electrons easily because they have relatively high ionization energies (the energy required to remove an electron to form a cation). Nonmetals also have high or very negative electron afffinities (the energy involved in adding an electron to form an anion), indicating that nonmetals readily gain electrons to form anions.
The trends for ionization and electron affinity are both as follows:

increase
→ →
Increase ↑
↑

So the nonmetals on the right side of the periodic table have high ionization energy and very negative electron affinity.

5.(1) How many unpaired electrons are there in the species V^{3+}?

1. zero 2. one 3. two 4. Three

Answer: 3. two

Since V has a +3 charge, 3 electrons are lost to form this ion. The first two electrons come from the 4s orbital (this orbital has the highest "n" value) and the last electron come from a 3d orbital:
V: [Ar]4s23d^3;
V^{3+}: [Ar]3d^2 _↑_   __↑_   ___   ___   ___
3d

The two remaining 3d electrons left are each unpaired for a total of two unpaired electrons.

6.(8) Analysis of an unknown substance showed that is has a high boiling point and is brittle. It conducts electricity when melted, but not as a solid. Which of the following substances would have those characteristics?

1. HCl 2. Al 3. KBr 4. I_2

Answer: 3. KBr

Ionic compounds have high lattice energies while the bond energies in covalent compounds are not as high. Ionic bonds are stronger than covalent bonds. The strength of ionic bonds results in ionic compounds being high melting and high boiling and being physically hard and crystalline. Ionic compounds do not conduct in the solid phase as the ions are locked in place in the solid, but will conduct when melted due to the presence of ions to carry current. Covalent compounds have relatively low melting and boiling points and are not electrical conductors since no ions are present. KBr is the only ionic compound in this list.

7.(11) Which bond is most polar?

1. C–O 2. Cl–F 3. C–F 4. Si–F

Answer: Si-F

A bond between two atoms of different elements is polar which means that the more electronegative atom in the bond pulls the shared electron cloud towards it, away from the less electronegative atom. Thus the more electronegative atom has a partial negative charge and the less electronegative atom has a partial positive charge. The electron cloud is shared unequally. The greater the difference between the electronegativity values on the two atoms in the bond, the more unequally the electrons are shared and the more polar the bond. The electronegativity trend in the periodic table is:

increase
→ →
Increase ↑ C O F
↑ Si Cl

The electronegativity difference between Si and F is greater than that for the other bonds since Si and F are more widely separated in the periodic table.

8.(2) Bond order is __________ proportional to bond energy and ____________ proportional to bond length.

1. directly, directly
2. inversely, inversely
3. directly, inversely
4. inversely, directly

Answer: 3. directly, inversely

Bond order refers to the number of bonds between two atoms. A single bond (1 shared pair) has a bond order of 1, a double bond (2 shared pairs) has a bond order of 2 and a triple bond (3 shared pairs) has a bond order of 3. The higher the bond order, the more pairs of electrons being shared between the two atoms. As the number of electron pairs between two atoms increases, the bond becomes stronger and the atoms are pulled closer together. So, a higher bond order results in a shorter, stronger bond. Bond length decreases as bond order increases (inversely proportionality) and bond energy increases as bond order increases (direct proportionality).

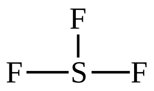
9.(14) In which of the following species does the central atom have three unshared (lone) pairs in its Lewis structure?

1. SF_3^+ 2. ClF_2^- 3. BrF_3 4. $GeCl_4$

Answer: 2. ClF_2^-

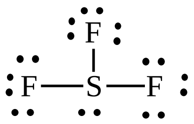
$6 + 3(7) - 1 = 26\ e^-$ $\longrightarrow$

SF_3^+



$$\begin{array}{rl} 26\ e^- & \\ -\ \ 6\ e^- & \text{in bonds} \\ \hline 20\ e^- & \text{left} \end{array}$$

$\xrightarrow[\text{octets}]{\text{complete}}$



$$\begin{array}{rl} 20\ e^- & \\ -\ \ 20\ e^- & \text{used} \\ \hline 0 & \end{array}$$

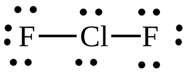
1 unshared pair on S

$7 + 2(7) + 1 = 22\ e^-$ $\longrightarrow$ F—Cl—F

ClF_2^-

$$\begin{array}{rl} 22\ e^- & \\ -\ \ 4\ e^- & \text{in bonds} \\ \hline 18\ e^- & \text{left} \end{array}$$

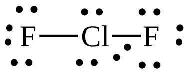
$\xrightarrow[\text{octet}]{\text{complete}}$



$$\begin{array}{rl} 18\ e^- & \\ -\ \ 16\ e^- & \text{used} \\ \hline 2\ e^- & \text{left} \end{array}$$

put on central
atom

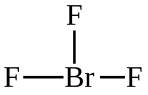
$\longrightarrow$



3 unshared pairs on Cl

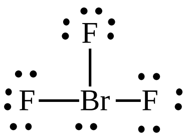
$7 + 3(7) = 28\ e^-$ $\longrightarrow$

BrF_3



$$\begin{array}{rl} 28\ e^- & \\ -\ \ 6\ e^- & \text{in bonds} \\ \hline 22\ e^- & \text{left} \end{array}$$

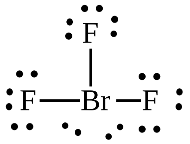
$\xrightarrow[\text{octets}]{\text{complete}}$



$$\begin{array}{rl} 22\ e^- & \\ -\ \ 20\ e^- & \text{used} \\ \hline 2\ e^- & \text{left} \end{array}$$

put on central
atom

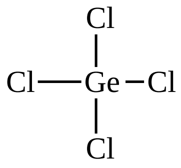
$\longrightarrow$



2 unshared pairs on Br

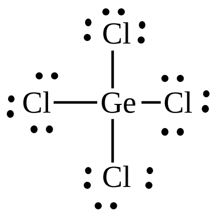
$4 + 4(7) = 32\ e^-$ $\longrightarrow$

$GeCl_4$



$32\ e^-$
$-\ 8\ e^-$ in bonds
$24\ e^-$ left

$\xrightarrow[\text{octets}]{\text{complete}}$



$24\ e^-$
$-\ 24\ e^-$ used
$0\ e^-$

no unshared pairs on Ge

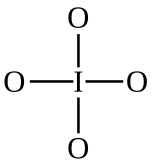
10.(16)The central atom in IO_4^- is surrounded by

1. 4 single bonds, no double bonds and no lone pairs of electrons.
2. 3 single bonds, 1 double bond and no lone pairs of electrons.
3. 3 single bonds, 1 double bond and 2 lone pairs of electrons.
4. 4 single bonds, no double bonds and 2 lone pairs of electrons.

Answer: 1. 4 single bonds, no double bonds and no lone pairs of electrons

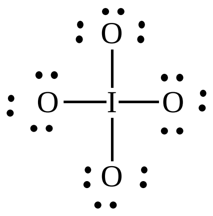
$7 + 4(6) + 1 = 32\ e^-$ $\longrightarrow$

IO_4^-



$32\ e^-$
$-\ 8\ e^-$ in bonds
$24\ e^-$ left

$\xrightarrow[\text{octets}]{\text{complete}}$



$24\ e^-$
$-\ 24\ e^-$ used
0

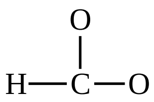
11.(12)Which of the following species has resonance structures?

1. $GeCl_4$
2. HCO_2^- (C is central atom)
3. BrF_3
4. SF_3^+

Answer: 2. HCO_2^-

$GeCl_4$: See Problem 9 for Lewis structure; only single bonds present: no resonance

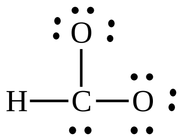
$1 + 4 + 2(6) + 1 = 18\ e^-$ $\longrightarrow$



$$\begin{array}{rl} 18\ e^- & \\ -\ \ 6\ e^- & \text{in bonds} \\ \hline 12\ e^- & \text{left} \end{array}$$

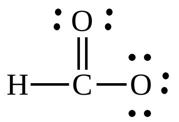
HCO_2^-

$\xrightarrow[\text{octets}]{\text{complete}}$

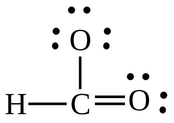


$$\begin{array}{rl} 12\ e^- & \\ -\ 14\ e^- & \text{used} \\ \hline -\ \ 2\ e^- & \text{too many - use another bond} \end{array}$$

$\longrightarrow$



or



Two possible Lewis structures: **resonance**

BrF_3: See Problem 9 for Lewis structure; only single bonds present: no resonance

SF_3^+: See Problem 9 for Lewis structure; only single bonds present: no resonance

12.(15)What is the formal charge on C and N in the molecule HCN?

1. C = +1, N = -1
2. C = 0, N = -1
3. C = +1, N = 0
4. C = 0, N = 0

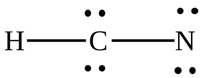
Answer: 4. C = 0; N = 0

$1 + 4 + 5 = 10\ e^-$ $H—C—N$

$$\begin{array}{rl} 10\ e^- & \\ -\ \ 4\ e^- & \text{in bonds} \\ \hline 6\ e^- & \text{left} \end{array}$$

HCN

$\xrightarrow[\text{octets}]{\text{complete}}$



$$\begin{array}{rl} 6\ e^- & \\ -\ 10\ e^- & \text{used} \\ \hline -\ \ 4\ e^- & \text{to many - use 2 additional bonds} \end{array}$$

$\longrightarrow$ $H—C\equiv N:$

Formal charge = # of valence electrons - (# of unshared electrons + 1/2 shared electrons)

C: 4 - (0 + 4) = 0
(C has no unshared electrons and 4 bonds)

N: 5 - (2 + 3) = 0
(N has 2 unshared electrons and 3 bonds)

13.(7) Pure covalent bonds in which the bonding electrons are shared equally between two atoms only exist for

1. period 2 nonmetals.
2. group 14 nonmetals.
3. metals.
4. bonds between atoms of the same element.
5. ionic compounds.

Answer: 4. Bonds between atoms of the same element

Electrons are shared equally between atoms with identical electronegativity values. A bond between atoms of the same element would be nonpolar as both atoms in the bond would have identical attraction for the shared electrons in the bond.

14.(9) Given the following bond energies, calculate ΔH° for the reactions

$$N_2(g) + 3\ F_2(g) \rightarrow 2\ NF_3(g)$$

Bond energies of N_2 = 945 kJ/mol, F_2 = 159 kJ/mol, N–F = 272 kJ/mol

1. -210 kJ 2. 606 kJ 3. -560 kJ 4. 878 kJ

ΔH° = ΔH°(reactant bonds broken) + ΔH°(product bonds formed)
Bond breaking is a positive quantity: energy is required to break bonds
Bond formation is a negative quantity: energy is released when bonds are formed.

Bonds broken in reactants: 1 N_2 bond: 1(945 kJ/mol) = 945 kJ
3 F-F bonds: 3(159 kJ/mol) = $\underline{477\text{ kJ}}$
total 1422 kJ

Bonds formed in products: 6 N-F bonds: 6(272 kJ/mol) = -1632 kJ
(Each of the 2 NF_3 molecules has 3 N-F bonds; the total energy is negative as energy is released in bond formation)

ΔH° = ΔH°(reactant bonds broken) + ΔH°(product bonds formed)
ΔH° = 1422 kJ + (-1632 kJ)
ΔH° = -210 kJ

15.(13)Which of the following species has an expanded octet?

1. HCO_2^- (C is central atom)
2. SF_3^+
3. $GeCl_4$
4. ClF_2^-

Answer: 4. ClF_2^-

Expanded octet: an atom with more than 8 electrons:

HCO_2^-: See Problem 11 for the Lewis structure. H has two electrons and C and O each have 8 electrons.

SF_3^+: See Problem 9 for the Lewis structure. All atoms have only 8 electrons.

$GeCl_4$: See Problem 9 for the Lewis structure. All atoms have only 8 electrons

ClF_2^-: See Problem 9 for the Lewis structure. **The Cl has 10 electrons in this structure.**

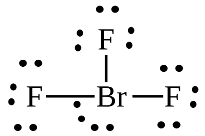
16.(18)What is the molecular geometry of BrF_3?

1. T-shaped
2. trigonal planar
3. trigonal pyramid
4. square planar

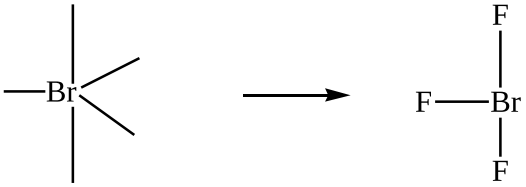
Answer: 1. T-shaped

See Problem 9 for the Lewis structure:

BrF_3



5 electron groups on Br



T-shaped

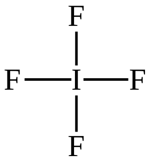
17.(22)Which of the following species has tetrahedral molecular geometry?

1. IF_4^-
2. XeF_4
3. SCl_4
4. $GeCl_4$

Answer: 4. $GeCl_4$

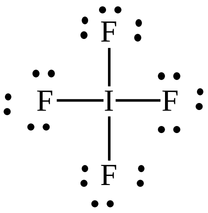
$7 + 4(7) + 1 = 36\ e^-$ →

IF_4^-



$$\begin{array}{rl} 36\ e^- & \\ -\ \ 8\ e^- & \text{in bonds} \\ \hline 28\ e^- & \text{left} \end{array}$$

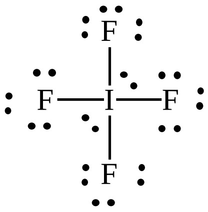
$\xrightarrow[\text{octets}]{\text{complete}}$



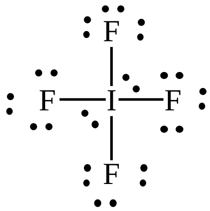
$$\begin{array}{rl} 28\ e^- & \\ -\ 24\ e^- & \text{used} \\ \hline 4\ e^- & \text{left} \end{array}$$

put on central atom

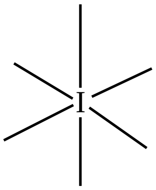
→



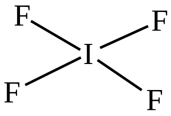
IF_4^-



6 electron groups on I →



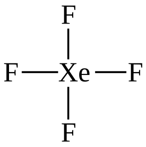
→



square planar

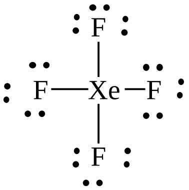
$8 + 4(7) = 36\ e^-$ →

XeF_4



$$\begin{array}{rl} 36\ e^- & \\ -\ \ 8\ e^- & \text{in bonds} \\ \hline 28\ e^- & \text{left} \end{array}$$

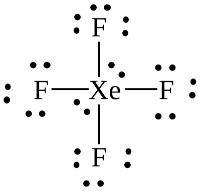
$\xrightarrow[\text{octets}]{\text{complete}}$



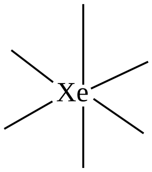
$$\begin{array}{rl} 28\ e^- & \\ -\ 24\ e^- & \text{left} \\ \hline 4\ e^- & \text{left} \end{array}$$

put on central atom

→



6 electron groups on Xe



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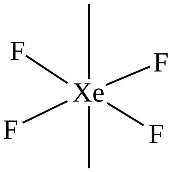
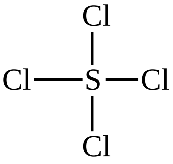


square planar

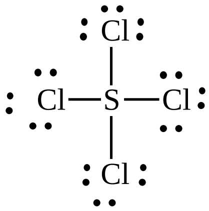
$6 + 4(7) = 34\ e^-$ $\longrightarrow$

SCl_4



$$\begin{array}{rl} 34\ e^- & \\ -\ \ 8\ e^- & \text{in bonds} \\ \hline 26\ e^- & \text{left} \end{array}$$

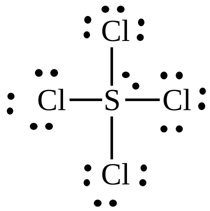
$\xrightarrow[\text{octets}]{\text{complete}}$



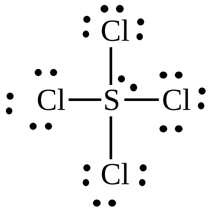
$$\begin{array}{rl} 26\ e^- & \\ -\ 24\ e^- & \text{used} \\ \hline 2\ e^- & \text{left} \end{array}$$

put on central atom

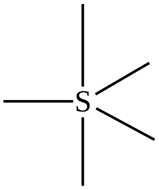
$\longrightarrow$



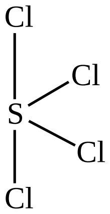
SCl_4



5 electron groups on S $\longrightarrow$



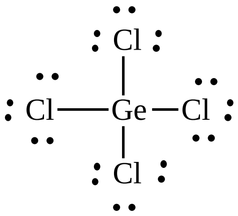
$\longrightarrow$



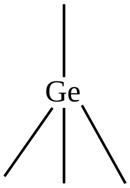
see-saw

$GeCl_4$: See Problem 9 for the Lewis structure:

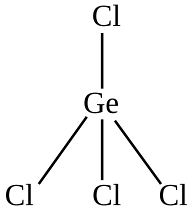
$GeCl_4$



4 electron groups on Ge



$\longrightarrow$



tetrahedral

18.(19)Which of the following species is planar (flat?)

1. $GeCl_4$ 2. H_2S 3. SF_3^+ 4. $TeCl_4$

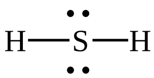
Answer: 2. H_2S

$GeCl_4$: See Problem 17 for the molecular geometry: tetrahedral, which is not a planar shape.

H_2S

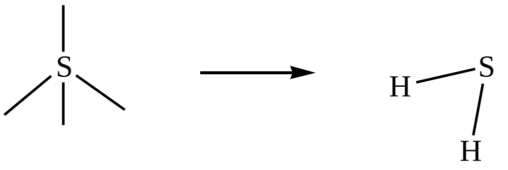
$$2(1) + 6 = 8\,e^- \longrightarrow \text{H—S—H} \qquad \begin{array}{rl} 8\ e^- & \\ -\ 4\ e^- & \text{in bonds} \\ \hline 4\ e^- & \text{left} \end{array}$$

$\xrightarrow[\text{octets}]{\text{complete}}$



$$\begin{array}{rl} 4\ e^- & \\ -\ 4\ e^- & \text{used} \\ \hline 0 & \end{array}$$

4 electron groups in S

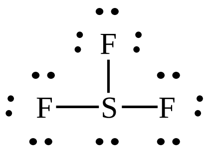


Bent

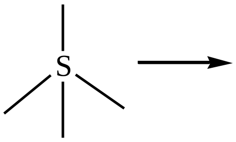
This is a planar molecule

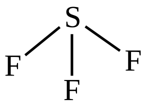
SF_3^+: See Problem 9 for the Lewis structure:

SF_3^+



4 electron groups on S



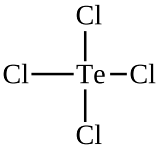


trigonal pyramid

This is not a planar structure.

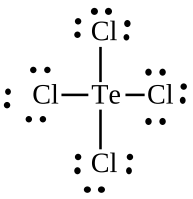
$6 + 4(7) = 34\ e^-$ →

$TeCl_4$



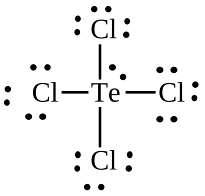
$34\ e^-$
$-\ 8\ e^-$ in bonds
$26\ e^-$ left

complete → octets

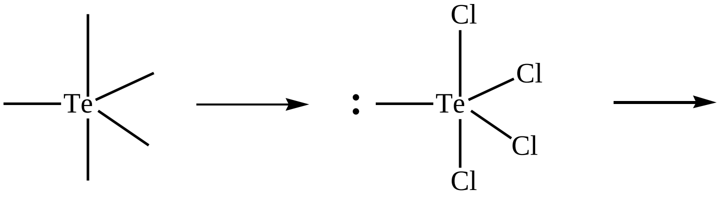


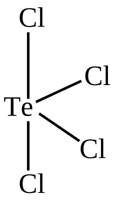
$26\ e^-$
$-\ 24\ e^-$ left
$2\ e^-$ left
put on central atom

→



5 electron groups around Te:





see-saw

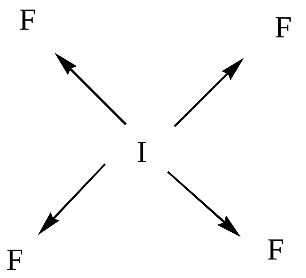
This is not a planar structure.

19.(20)Which of the following species is polar?

1. IF_4^- 2. SO_3 3. $TeCl_4$ 4. $GeCl_4$

Answer: 3. $TeCl_4$

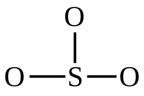
IF_4^-: See Problem 17 for the geometry:



Polar bonds cancel nonpolar

$6 + 3(6) = 24\ e^-$

SO_3

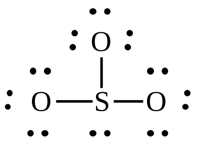


$24\ e^-$

$-\ 6\ e^-$ in bonds

$18\ e^-$ left

$\xrightarrow[\text{octets}]{\text{complete}}$

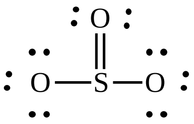


$18\ e^-$

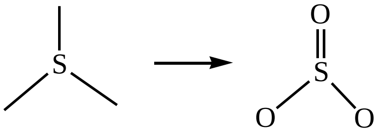
$-\ 20\ e^-$

$-\ 2\ e^-$ to many - use another bond

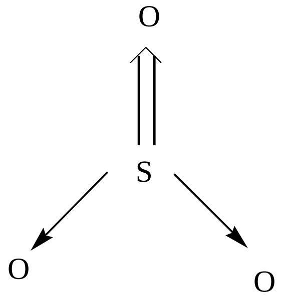
→



3 electron groups in S



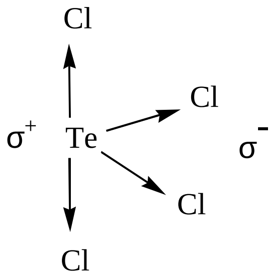
trigonal planar



Polar bonds cancel - nonpolar

$TeCl_4$: See Problem 18 for the geometry:

$TeCl_4$

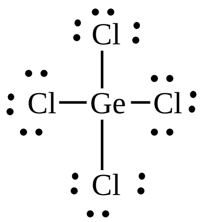


polar molecule

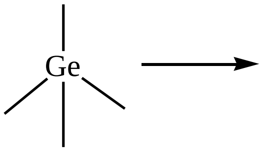
The polar bonds do not cancel and this is a polar molecule.

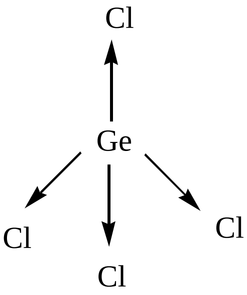
$GeCl_4$: See Problem 18 for the geometry:

GeCl$_4$



4 electron groups on Ge





All polar bonds cancel: nonpolar molecule

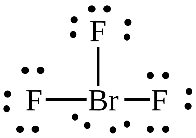
20.(21)Which statement below concerning BrF$_3$ is true?

1. The bonds are polar and the molecule is polar.
2. The bonds are nonpolar and the molecule is nonpolar.
3. The bonds are polar and the molecule is nonpolar.
4. The bonds are nonpolar and the molecule is polar.

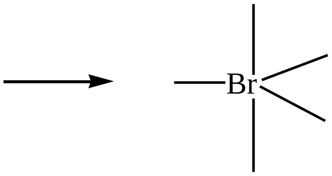
Answer: 1. The bond are polar and the molecule is polar.

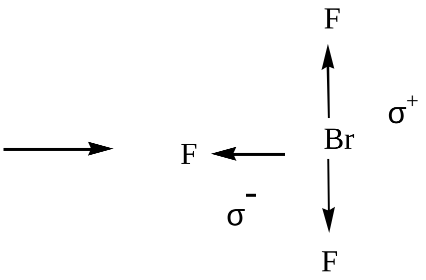
See Problem 9 for the Lewis structure:

BrF$_3$



5 electron groups on Br





polar molecule

Each of the 3 Br-F bonds are polar since F is more electronegative than Br and pulls the electron shared pair towards it. Since the structure is not symmetrical, there is a net movement of the electron cloud away from the Br towards the F atoms in the triangular plane. The molecule is polar since there is a "positive side" vs a "negative side" in the molecule.

21.(24)In the species SF_3^+ the central atom uses hybrid orbitals of this type.

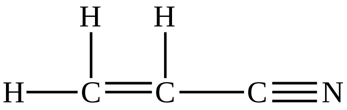
1. sp 2. sp^2 3. sp^3

4. sp^3d 5. sp^3d^2

Answer: 3. sp^3

See Problem 9 for the Lewis structure of SF_3^+.
Since the central atom S has 4 electron groups (3 shared pairs and one lone pair), it will require 4 hybrid orbitals: sp^3 hybrid orbitals occur in a set of 4.

22.(23)How many sigma (σ) bonds and pi (π) bond are in the following molecule?



$\mathrm{H{-}\overset{H}{\overset{|}{C}}{=}\overset{H}{\overset{|}{C}}{-}C{\equiv}N}$

1. 4 σ and 5 π 2. 4 σ and 2 π
3. 6 σ and 5 π 4. 6 σ and 3 π

Answer: 4. 6 σ and 3 π

Each single bond is a σ bond. There are 4 single bonds. The double bond is 1 σ bond and 1 π bond. The triple bond is 1 σ bond and 2 π bonds. That is a total of 6 σ and 3 π bonds.

23.(10)What is the total number of valence electrons in the species IF_4^-?

1. 36 2. 16 3. 30 4. 24

Answer: 1. 36

I: Group 17: 7 valence electrons
F: Group 17: 4(7) = 28 valence electrons (4 F atoms in the ion)
-: add one more electron due to the negative change: 1 valence electron
Total valence electrons: 7 + 28 + 1 = 36

24.(25)Which of the following geometries is associated with sp^3d hybrid orbitals?

1. octahedral
2. trigonal pyramid
3. square planar
4. see-saw

Answer: 4. see-saw

sp^3d hybrid orbitals are used by a central atom that needs to accommodate 5 pairs of electrons. The see-saw geometry arises when a central atom has 4 bonds and 1 lone pair for a total of 5 electron groups. Octahedral has 6 electron groups, square planar has 6 electron groups and trigonal pyramid has 4 electron groups.

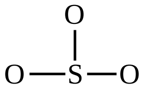
25.(17)The fact that SO_3 is planar while SF_3^+ is trigonal pyramidal is best explained by which of the following statements?

1. Fluorine is more electronegative than oxygen.
2. SO_3 has sp hybrid orbitals while SF_3^+ uses sp^2 hybrid orbitals.
3. S in SF_3^+ has a lone pair of electrons while S in SO_3 does not.
4. Fluorine is smaller than oxygen.

Answer: 3. S in SF_3^+ has a lone pair of electrons while S in SO_3 does not.

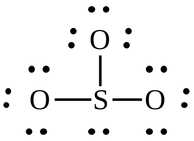
$6 + 3(6) = 24\ e^-$

SO_3



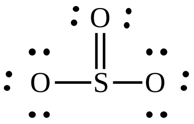
$$\begin{array}{rl} & 24\ e^- \\ - & 6\ e^- \ \text{in bonds} \\ \hline & 18\ e^- \ \text{left} \end{array}$$

$\xrightarrow[\text{octets}]{\text{complete}}$

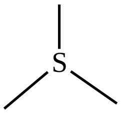


$$\begin{array}{rl} & 18\ e^- \\ - & 20\ e^- \\ \hline - & 2\ e^- \ \text{to many - use another bond} \end{array}$$

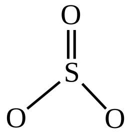
$\longrightarrow$



3 electron groups in S

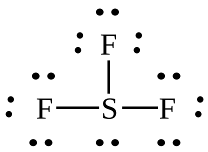


$\longrightarrow$

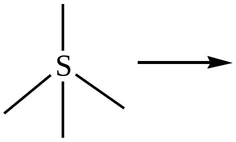


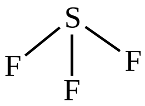
trigonal planar

SF_3^+



4 electron groups on S





trigonal pyramid

SF_3^+ has 3 bonds like SO_3 but a total of 4 electron groups since S has a lone pair. The presence of the lone pair changes the way in which the 3 bonds are arranged.