



CHEM 110 Exam 2 - Practice Test 1 - Solutions

1D

1 has a triple bond. 2 has a double bond. 3 and 4 have single bonds. The stronger the bond, the shorter the length.

2A

A 1:1 ratio means there must be the same amount of Carbons as there are Oxygens.
A 2:1 ratio means there must be twice as many hydrogen's as there are carbons.
The only choice that satisfies these requirements is choice 1, since in choice 2 there are too many oxygens, and in choice 3 the formal charges on oxygen are -1.

3A

The correct Lewis structures are imperative to answer this question with confidence.

IF₃ – the central iodine atom will have 3 single bonds and two lone pairs of electrons. The correct VSEPR is AX₃E₂, which corresponds to T-shaped.

PCl₅ - the central phosphorus atom will have 5 single bonds and no lone pairs of electrons. The correct VSEPR is AX₅, which corresponds to trigonal bipyramidal.

NH₃ - the central nitrogen atom will have 3 single bonds and one lone pair of electrons. The correct VSEPR is AX₃E₁, which corresponds to trigonal pyramidal.

4E

Even though the overall charge of the phosphate ion is -3, the central atom, phosphorus, has a formal charge of +1 in this Lewis structure.

$$\text{Formal charge} = \# \text{ valence } e\text{'s} - (\# \text{ unshared valence } e\text{'s} + 1/2 \# \text{ shared } e\text{'s})$$

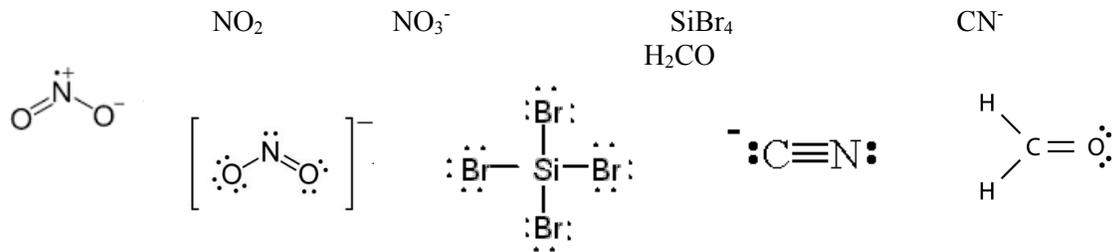
Phosphorous has 5 valence electrons (Group 5A) and is bonded to four oxygens with single bonds, and has no lone pairs. So,

$$\text{Formal charge} = 5 - (0 + 1/2 * 4) = +1$$

The formal charges of each individual atom add up to the total charge of the ion. This is a good way to check your answer.

5A

There are three cases of violations to the octet rule (8 e's surrounding atom): expanded valence shells (>8 e's), electron deficient molecules (<8 e's), and odd-electron molecules.



All central atoms have 8 valence e's, except for NO₂. This is the case in which the molecule has an odd number of electrons

6D

Again, the correct Lewis structures are essential to determining polarity.

SF₄ – This molecule has one lone electron pair on the central atom, which always means it's polar.

CH₃Br – Here Br acts as the odd-man-out creating a polar bond that isn't countered anywhere else in the molecule, thus it is polar.

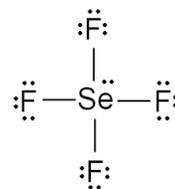
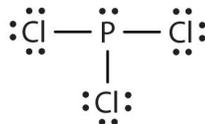
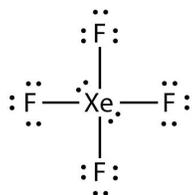
H₂S – Very similar to water, the sulfur has two lone pairs of electrons, but since the molecule's geometry is bent shaped, the lone pairs cannot cancel each other out, and thus it's polar.

XeF₂ – Here the central atom has two single bonds and three lone pairs of electrons. The corresponding shape is linear, which is always nonpolar if the surrounding atoms (in this case the fluorines) are the same. Thus we have our correct answer.

C₄H₉OH – As a rule, any molecule with an OH attached to it, is almost certainly going to be polar (and will also undergo hydrogen bonding).

7C

Draw the lewis structures for each and write the AX_mE_n formula and assign the associated molecular geometry.



AX_4E_2 : Square planar

AX_3E_1 : Trigonal pyramidal

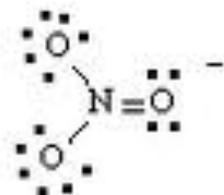
AX_4E_1 : Seesaw

So, i and ii are correctly assigned. Always remember exceptions to the octet rule!

8D

This is one of the octet rule exceptions dealing with odd-numbered electrons. Whenever you have an odd number of electrons the central atom will usually be forced to have 7 electrons surrounding it rather than the usual 8. In the case of Nitrogen Dioxide, nitrogen has a single bond to one of the oxygens, a double bond to the other oxygen, and ONLY ONE lone electron that it keeps for itself. Thus the formal charge is: $5 - (1 + 3) = +1$.

9B



The least amount of formal charge around each individual atom is accomplished through this structure best.

10(i)C

The sigma bonds are always formed using the hybrid orbitals (except for H atoms which only use s-orbitals) and so we must find the hybridization of C, which has two electron groups, and therefore must be sp .

10(ii)B

Pi bonds will always be formed using overlapping p-orbitals (not hybrid orbitals).

11D

We first want to look at the Lewis structure (which shows three single bonds and a lone electron pair on the central atom), thus the shape is trigonal pyramidal and the ideal angle is 109.5° . However, in almost every case, the angle is less than that since the lone electron pair pushes together the other atoms, and so we should look for a value a little lower than 109.5 , like 108° in this case.

12E

Electronegativity increases going up and to the right of the periodic table; fluorine is the most electronegative element. The bond with the greatest difference in electronegativity between both atoms will be the most polar. B and F are separated by the most space compared to the other bonds, so that will be the most polar bond.

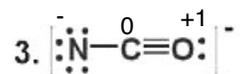
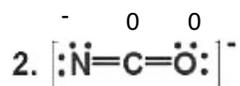
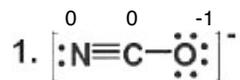
13C

The central C is double bonded to O and single bonded to each Cl. As usual, there are no lone electrons on the C. Thus, C has three electron groups, and is sp^2 .

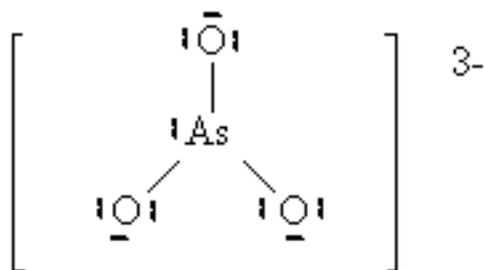
14B

Draw out possible Lewis structures for NCO^- and assign formal charges to each of the atoms.

The Lewis structures below obey the octet rule; there are other Lewis structures that do not. The most stable Lewis structure will only involve 0, -1, and +1 charges. This eliminates Lewis structure 3. So, the question is if the -1 charge will go on the oxygen or nitrogen. If there are two Lewis structures that have the same charges but on different atoms, put the charge on the more electronegative atom. Oxygen is more electronegative than nitrogen, so Lewis structure 1 is the best.



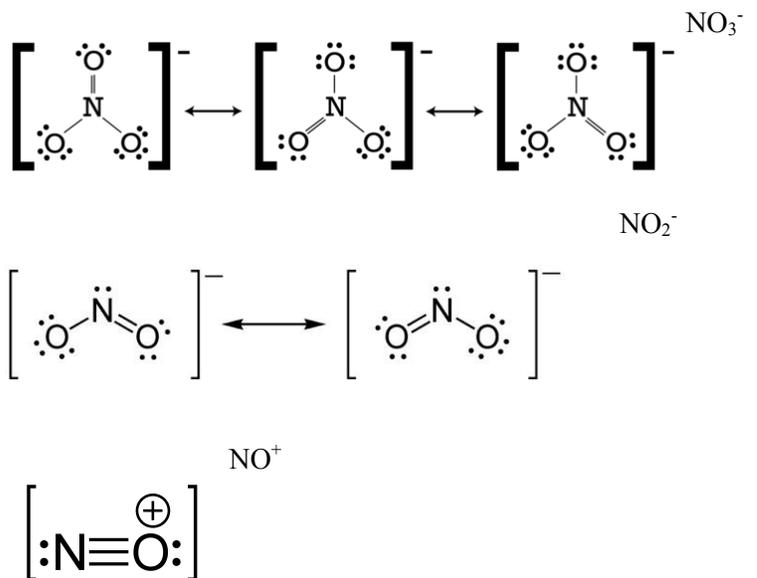
15B



Due to the -3 charge, this is the best way for the formal charges to be spread out among the atoms. So, we only have 1 electron pair on As.

16B

It is necessary to draw out possible resonance structures for each ion.



So, in NO_3^- and NO_2^- the π electrons are delocalized (spread throughout the ion). The average bond order between nitrogen and oxygen can be calculated by adding the total number of bonds (e.g. double bond=2 bonds) and dividing by the number of oxygen atoms nitrogen is bonded to.

NO_3^- : 4 bonds/3 oxygen atoms = 4/3 bond order

NO_2^- : 3 bonds/2 oxygen atoms = 3/2 bond order

NO^+ only has one structure that is all-octet, so its bond order is 3.

With increasing bond order, the bonds get shorter because the atoms are bonded closer together. So, B is the correct answer.

17D

sp^3d^2 corresponds to a central atom with 6 electron groups, and that can only mean octahedral.

18E

A single bond contains one sigma bond.

A double bond contains one pi bond and one sigma bond.

A triple bond contains two pi bonds and one sigma bond.

The molecule contains 21 single bonds, and 4 double bonds, which equates to **25 sigma bonds, and 4 pi bonds**.

As for the lone pairs:

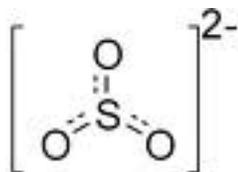
When oxygen has 4 shared electrons it will have 2 lone pairs.

When nitrogen has 6 shared electrons it will have 1 lone pair.

Thus the two oxygens equate to 4 lone pairs, and the four nitrogens equate to four lone pairs, making **8 lone pairs in total**.

19D

A structure with delocalized bonding is one that has resonance. This is a common occurrence in compounds with double bonds. The one double bond in SO_3 can be between sulfur and any three of the oxygens, so the hybrid structure shown below is the true representation of the structure.



20E

Sigma bonds result from the head-on overlap of orbitals along the internuclear axis.

Sigma bonds have more overlap of the orbitals than pi bonds, so they are stronger interactions than pi bonds. A triple bond contains one sigma and two pi bonds. Atoms cannot rotate around pi bonds because it would decrease the overlap between the parallel orbitals. The sideways overlap of two parallel, unhybridized p orbitals is an example of a pi bond, so E is true.

21A

- I. For NO_2^- the Lewis structure shows N with a single bond to O, a double bond to another O, and a lone pair. This is 3 electron domains surrounding N, which suggests sp^2 .
- II. For CO_2 the Lewis structure shows C with a double bond to each O, and no lone pairs. This is 2 electron domains surrounding C, which suggests sp .
- III. For O_3 , the Lewis structure shows the central O with a single bond to O, and double bond to O, and a lone pair. This is 3 electron domains, which suggests sp^2 .

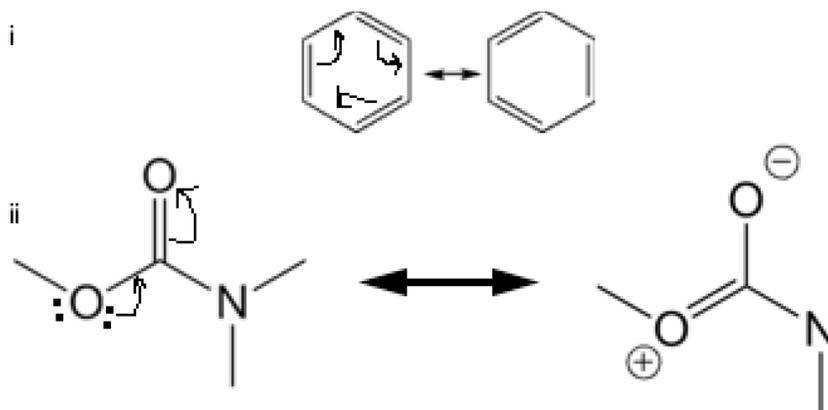
22B

The best Lewis structure will give each atom an octet and have the fewest charges overall. Since it's an uncharged molecule it is likely that the formal charge of each atom will be 0. For these criteria, nitrogens will have 3 bonds (e.g. double bond=2 bonds) and 1 lone pair, carbons will have 4 bonds, and oxygens will have 2 bonds and 2 lone pairs. The only structure that does this is B.

23D

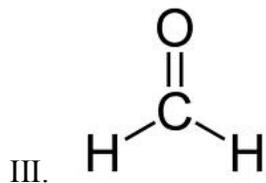
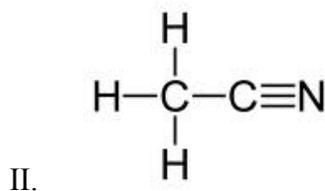
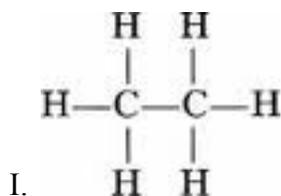
Resonance structures arise when there are multiple Lewis structures that can be drawn. The connectivity of the atoms needs to be the same. So, iii does not show resonance structures because the connectivity isn't the same. These molecules are isomers of each other.

The figure to the right shows the electron-pushing mechanism.



24D

A structure with free rotation around all its bonds would have all sigma bonds (single bonds) and zero pi bonds (double and triple bonds).



25E

The overall dipole moment of a molecule is the vector sum of the individual dipole moments of each bond, taking into account the molecule's geometry. HF is a linear molecule that has two atoms with the largest difference in electronegativities, so it is the most polar molecule and therefore the molecule with the greatest dipole moment.

LionTutors

CHEM 110 Exam 2 - Practice Test 2 – Solutions

1A

The total number of electron domains attached to an atom determines its hybridization. Electron domains can be either a bond or a pair of nonbonding electrons. NO_2^- has three electron domains, which are two bonds between N and an O and one nonbonding electron pair, so it has three sp^2 hybrid orbitals. CO_2 and O_3 have two and three electron domains, respectively, so CO_2 has two sp hybrid orbitals and O_3 has three sp^2 hybrid orbitals.

2E

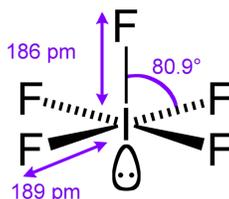
Octahedral is the electron domain geometry of sp^3d^2 hybridization.

3A

The “formula” for seesaw molecular geometry is AX_4E . So, there is one lone pair. Assuming the molecule is not charged, an atom that has 4 single bonds and one lone pair would need to have 6 valence electrons to achieve a formal charge of 0. Therefore, sulfur could be the central atom because it is in group 6.

4C

Remember “penta-“ means 5, “tetra-“ means 4, “tri-“ means 3, “di-“ means 2, and “mono-“ means 1. Of the choices only IF_5 has six electron domains, so it has an octahedral electron domain geometry.



5C

Polarity is a measure of differing electronegativity values for the two elements that form the bond. Another way to look at this is just to measure how far apart the elements are from each other on the periodic table.

Using this method, we get $\text{Mg} - \text{F} > \text{N} - \text{F} > \text{Br} - \text{Br}$.

Note: Bonds involving the same element (like $\text{Br} - \text{Br}$) have zero polarity and are called non-polar.

6E

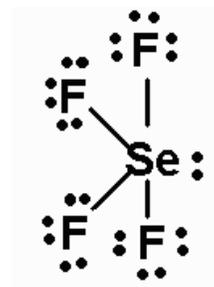
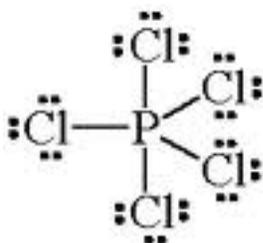
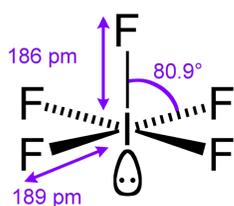
Remember that valence electrons correspond directly to the group number of the element. Since Br and F are both in group 7, we add 7 electrons for Br and 28 electrons for the four Fs. We also add one more electron for the negative charge. The total is 36 electrons.

7D

The only elements that can violate the octet rule (meaning that the central atom does not have exactly eight electrons surrounding it) are those in periods 3 and higher as well as B and Be. With that in mind you don't even need to draw Lewis structures to know the answer must be BH_3 .

8C

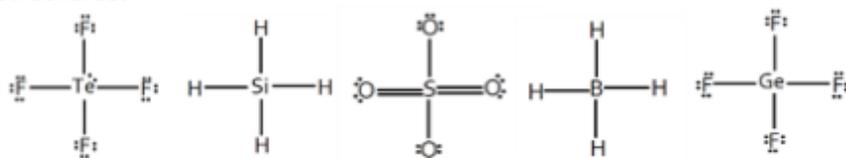
Trigonal bipyramidal molecular geometry corresponds to a central atom with five electron domains attached to it, none of which are lone electron pairs. PCl_5 is the only compound that fits this structural setup.



9

- A. TeF_4
- C. SO_4^{2-}

To solve this problem you must draw the Lewis Structures. Below are each of the Lewis structures.

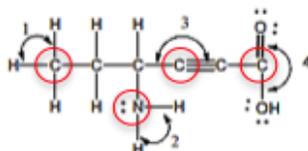


SO_4^{2-} could be drawn without breaking the octet rule, but this structure has more acceptable formal charges.

10

(e) $3 > 4 > 1 > 2$

In order to determine the bond angles, start by identifying the central atom associated with each bond angle. The central atom is the vertex of the angle. For Bond Angles 1, 3, and 4, the central atom is a Carbon. The central atom in Bond Angle 2 is Nitrogen. These central atoms are marked below.



Now to determine the bond angles, you must determine the geometry of each separate bond.

Bond Angle 1:

The central atom C has 4 different atoms bonded to it (AX_4E_0) tetrahedral = 109.5°

Bond Angle 2:

The central atom N has 3 different atoms bonded to it (AX_3E_1) tetrahedral = $<109.5^\circ$
Remember that lone electron pairs cause the bond angles to slightly decrease.

Bond Angle 3:

The central atom C has 2 different atoms bonded to it (AX_2E_0) linear = 180°

Bond Angle 4:

The central atom C has 3 different atoms bonded to it (AX_3E_0) tetrahedral = 120°

11

(e) E = I, T-Shaped

First determine the shape, and then decide what the center molecule is.

The shape is T-Shaped. You can see the center molecule has 5 electron domains (Trigonal Bi-pyramidal) and 2 lone electron pairs meaning T- Shaped. To count electron domains, just count how many bonds the central atom is forming and count the number of lone electron pairs.

Or you can use the AXE formula to see that it is AX_3E_2 which corresponds to T shaped. Remember A stands for the center atom, X is the number of atoms bonded to the central atom, and E is the number of lone electron pairs on the central atom. (As an aside, Adding X and E will also tell you the number of electron domains.)

Now that we know the molecule is T-shaped, we can eliminate each answer choice except B and E leaving Sulfur and Iodine as the only possibilities for E. Again there are different methods to accomplish this, but the best way is to check the formal charge on each element if they were the central molecule.

If sulfur were the central atom, the formal charge would be:

$$(6 \text{ valence}) - (3 \text{ bonds}) - (4 \text{ lone electrons}) = -1$$

If I were the central atom, the formal charge would be:

$$(7 \text{ valence}) - (3 \text{ bonds}) - (4 \text{ lone electrons}) = 0$$

Iodine would be the best fit for the central atom.

12

(a) C_3H_4O

Find the percent composition of each answer choice. This is the most assured way to complete this problem. You could also draw them to check which do not have a double bond.

$$a) \text{ mass \% } C = \frac{3(12g)}{3(12g) + 4(1g) + 1(16g)} = 64.3\%$$

$$b) \text{ mass \% } C = \frac{3(12g)}{3(12g) + 7(1g) + 1(16g)} = 61.0\%$$

$$c) \text{ mass \% } C = \frac{3(12g)}{3(12g) + 2(1g) + 1(16g)} = 66.6\%$$

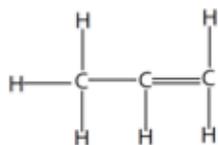
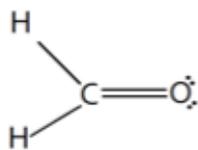
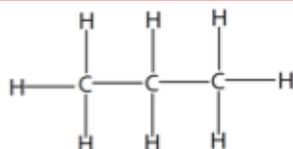
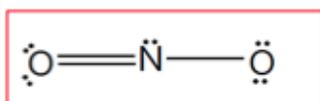
$$d) \text{ mass \% } C = \frac{4(12g)}{4(12g) + 7(1g) + 1(16g)} = 67.6\%$$

13

(b) NO_2

We only have delocalized π bonding when the structure has resonance. Recall that resonance occurs when electrons have the ability to move without changing anything else in the molecule.

Here are the Lewis structures:



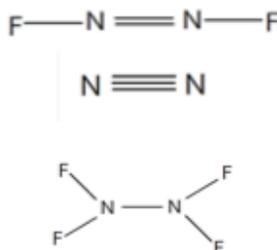
Only NO_2 can have delocalized π bonding. No other molecule can move the double bond without also moving a hydrogen, or breaking the octet.

14

(d) N_2 N_2F_2 N_2F_4

The magnitudes of the lengths given do not physically mean anything to us in this problem. If you look at the answer choices, you see that the lengths are arranged in increasing order, so the purpose of this problem is to arrange the molecules by increasing bond length. In order to accomplish this, you have to draw each molecule.

Here are the molecules (lone electrons not shown):



Since N_2 has a triple bond, the N-N bond length is the smallest. The double bond in N_2F_2 is shorter than the single bond in N_2F_4 . So $N_2 < N_2F_2 < N_2F_4$

15

(d) DA

Recall that polarity is defined as the difference of electronegativities. To solve this problem, find the difference between the electronegativities; the highest magnitude is the most polar.

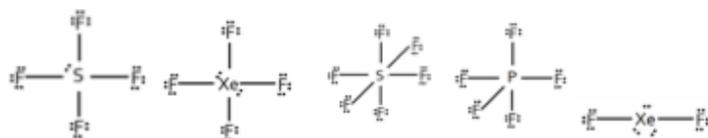
- a) $B - C = (2.01 - 2.04) = 0.03$
- b) $A - B = (1.01 - 2.01) = -1$
- c) $C - D = (2.04 - 3.05) = -1.01$
- d) $D - A = (3.05 - 1.01) = 2.04$
- e) $B - D = (2.01 - 3.05) = -1.04$

16

(a) SF₄

Draw each molecule. Usually with this type of problem, only one of the answer choices will actually be polar.

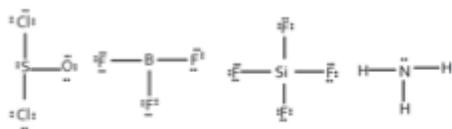
- (a) SF₄ – See- Saw, Always Polar
- (b) XeF₄ - Square Planer, non- Polar
- (c) SF₆ – Octahedral, non-Polar
- (d) PF₅ – Trigonal Pyramidal, non-Polar
- (e) XeF₂- Linear, non-Polar.



17

(b) 1 and 4 only

Draw them. Determine which is polar.



- 1. SOCl₂ Pyramidal- Always Polar
- 2. BF₃ Trigonal Planar, non-polar
- 3. SiF₄ tetrahedral, non-polar
- 4. NH₃ pyramidal, Always polar

18E

(a) The π electrons in SO_3^{-2} are delocalized. - TRUE

This is true. Structures with resonance have delocalized bonding.

(b) All C–C bond lengths in benzene are the same. - TRUE

For structures with resonance, we look at the bond order which is similar to an average bond length.

(c) More energy is required to break a carbon–nitrogen triple bond than energy needed to break a carbon–nitrogen double bond. – TRUE

Triple bonds require more energy to be broken.

(d) Hybridization allows for all F–C–F bond angles in CF_4 to be the same. – TRUE

It's the reason for hybridization.

(e) The bonding in benzene, C_6H_6 , includes 12 sigma (σ) bonds and 3 localized pi (π) bonds. – FALSE

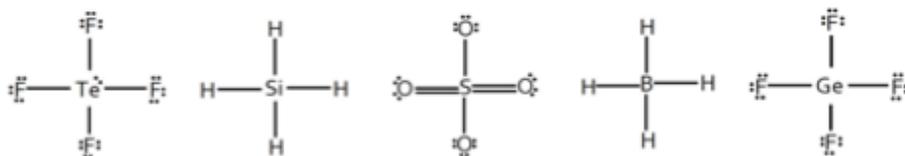
The pi bonds are delocalized.

19

(a) TeF_4

Believe it or not, you have to draw the Lewis structures for each of the answer choices!
Whichever does not have 4 electron domains is the answer.

Here are the Lewis structures:



20

(e) None of the above.

Count the number of electron domains on each carbon, then assign the correct hybridization. You can use AXE to help you. (A= central atom, X= number of elements bonded to central atom, E= number of lone electron pairs on central atom; Number of electron domains is X + E) Also, recall the sum of the hybridization (s + p + d) must be equal to the number of electron domains.

C1 \rightarrow AX₃E₀ \rightarrow 3 electron domains \rightarrow sp²

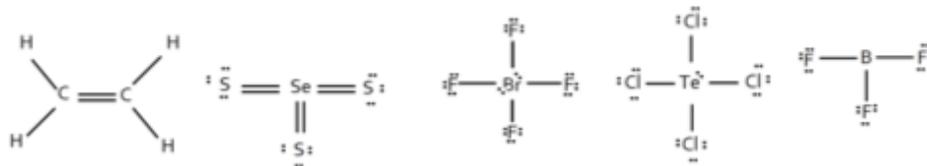
C2 \rightarrow AX₃E₀ \rightarrow 3 electron domains \rightarrow sp²

C3 \rightarrow AX₂E₀ \rightarrow 2 electron domains \rightarrow sp

21

(d) TeCl₄

There is a really quick and easy way of answering this problem. Just kidding. Draw each Lewis structure and determine which is not planar. Any electron domains over 3 can be 3D.



(a) C₂H₄ – Organic and planar. (no electron domains over 3)

(b) SeS₃ – Trigonal Planar

(c) BrF₄⁻ - Square Planar

(d) TeCl₄ – See Saw – 3D

(e) BF₃ – Trigonal Planar

22

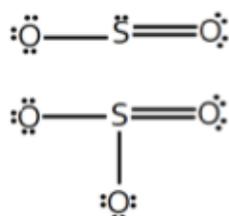
(e)

Each carbon is sp² as each one has three electron domains surrounding it, except for the two carbons that are triple bonded, which have only two electron domains and are thus sp hybridized.

23

- (a) All S–O bonds in SO_3^{2-} are the same because of resonance.
 (b) Each of the S–O bonds in SO_3^{2-} are weaker than S–O bonds in SO_2 .

Below are the Lewis structures for these molecules, the bonds lengths in each molecule will be the same because of resonance. The bond lengths in SO_2 are essentially $3/2 = 1.5$ and the bond lengths in SO_3^{2-} are $4/3 = 1.33$. These values are called the bond order and are a result of resonance and delocalized bonding. Answers A and B are both correct. With answer choice A, make sure they mention the correct number of resonance forms. (There would be three resonance forms for SO_3^{2-} and 2 for SO_2). For B) each of the SO_3^{2-} bonds are 1.33 which is longer and weaker than the SO_2 (1.5) bonds.



24

- (b) Trigonal bipyramidal

The molecule only contains Cl and F, so the $\text{mass \% Cl} + \text{mass \% F} = 100$. since we know mass percent Cl, we can solve for mass % F. (61.65 %)

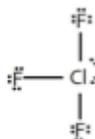
Now we can assume we have 100 grams of the substance meaning we have 38.35 grams of Cl and 61.65 grams of F. With these masses, we can determine the number of moles, and then find the mole to mole ration known as the empirical formula.

$$38.35 \text{ g Cl} * \frac{1 \text{ mole Cl}}{35.45 \text{ g Cl}} = 1.08 \text{ mole Cl}$$

$$61.65 \text{ g F} * \frac{1 \text{ mole F}}{18.99 \text{ g F}} = 3.25 \text{ mole F}$$

Divide through by the smallest number of moles to get the ratio $\rightarrow \text{ClF}_3$

Now draw this molecule:



The central atom Cl has 5 electron domains so this molecule is Trigonal bi-pyramidal. ($\text{AX}_3\text{E}_2 \rightarrow \text{X} + \text{E} = 5$ electron domains.)