You can preview this quiz, but if this were a real attempt, you would be blocked because:

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Question **1**

Not yet answered

Marked out of 10.00

Interpreting Lewis Structures. A Lewis structure is a visual representation of how the atoms are connected and how the valence or outermost electrons of the atoms are distributed in a molecule or polyatomic ion. When atoms form covalent bonds, only the valence (or outermost) electrons are involved. In a Lewis structure,

- each atom is represented by its symbol.
- valence electrons are represented by dots or lines.
- each dot represents one valence electron.
- a dash or line is equivalent to dots.
- lines or dots drawn between symbols of two atoms represent valence electrons by the two atoms.

Answer the questions below. You may find it useful to watch the following video first.

Consider the Lewis structure for HCN as shown below:

How many valence electrons are shared between C and N?

How many valence electrons are shared between H and C?

How many valence electrons are unshared?

What is the total number of valence electrons shown?

Of all the electrons that this molecule has, how many are not shown in the Lewis structure?

A pair of electrons shared between two atoms is called a bonding pair while an unshared pair of electrons is called a lone pair. If two atoms share one bonding pair, they are said to have a single bond; if they are sharing two bonding pairs, they are said to have a double bond; if they are sharing three bonding pairs, they are said to have a bond. In the Lewis structure for HCN above,

Which two atoms are linked by a single bond?

Which two atoms are linked by a double bond?

Which two atoms are linked by a triple bond?

How many bonding pairs are there?

Which atom has a lone pair?

Question **2**

Not yet answered

Marked out of 19.00

Drawing Lewis Structures. A Lewis structure shows valence electrons only. So, the first thing we need to do is determine the number of valence electrons that should be shown. The group number in the traditional American periodic table corresponds to the number of valence electrons of an atom belonging to that group (as predicted by quantum theory). One exception is He, which is located in column VIIIA but has only two valence electrons. Calculate the number of valence electrons that should be shown in the Lewis structures for the molecules and polyatomic ions listed below. Keep in mind that positive and negative charges (for polyatomic ions) mean electrons lost or gained.

Answer the following questions. You may find it useful to first watch this video:

The H_2O molecule has

- [ ] H atoms, each one contributing [ ] valence electron and
- [ ] O atom, contributing [ ] valence electrons.
- a total of [ ] valence electrons.

The H_3O^+ ion has

- [ ] H atoms, each one contributing [ ] valence electron and
- [ ] O atom, contributing [ ] valence electrons.
- But a charge of +1 means the entire structure has [ ] [ ] electron. Therefore, the total number of valence electrons is [ ].

The SO_4^{2-} ion has

- [ ] O atoms, each one contributing [ ] valence electrons and
- [ ] S atom, contributing [ ] valence electrons.

- A charge of -2 means the entire structure has [ ] [ ] electrons. Therefore, the total number of valence electrons is [ ].

Question **3**

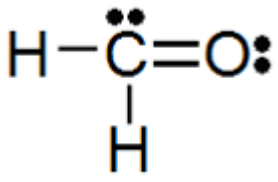
Not yet answered

Marked out of 13.00

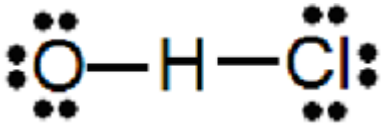
The octet/duet rule. When atoms covalently bond with other atoms, they *tend* to acquire a noble-gas-like configuration. H *always* forms just one single bond; when it does, it appears to be surrounded by a "duet" (two electrons) just like a He atom. Other atoms *tend* to share valence electrons such that they appear to be surrounded by an "octet" (8 valence electrons). Watch this video for an elaboration:

A structure is not necessarily incorrect just because an atom does not have an octet. However, atoms in period 2 can have less than an octet but can never have more than an octet. Atoms in period 3 and beyond can be surrounded by more than an octet. What is wrong, if any, with the structures shown below? You may find it useful to first watch this video:

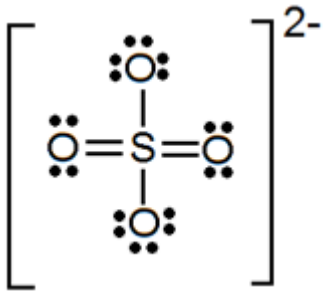
The structure below is incorrect because _____. Atoms in the _____ row of the periodic table can be surrounded by a maximum of _____ electrons.



The structure below is incorrect because it is showing [ ] valence electrons when it should only be showing [ ]. Furthermore it is showing an H atom with [ ] single bonds; a hydrogen atom cannot be surrounded by more than [ ] electrons.



The structure below [ ] showing the correct number of valence electrons. All the O atoms are surrounded by [ ]. The S atom is surrounded by [ ] valence electrons; this is [ ] because S is in the [ ] row of the periodic table; it [ ] be surrounded by more than an octet.



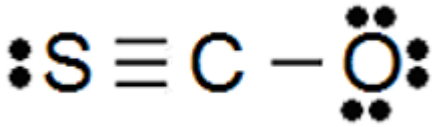
Question **4**

Not yet answered

Marked out of 10.00

Formal charges. Structures with zero formal charges are more likely to better represent reality; structures with formal charges other than 0, +1, or -1 are very unlikely to represent reality. Structures with formal charges of the same sign on neighboring atoms are also very unlikely.

To calculate formal charge (F.C.), we subtract the number of valence electrons that an atom appears to own in the structure (A) from the number of valence electrons in the free, unbonded atom (F).
F.C. = F - A
The number of valence electrons that an atom appears to own is the number of electrons in the lone pairs plus half of the electrons in the bonding pairs. What are the formal charges of the atoms in the structure drawn below. You may find it useful to first watch this video:



The S atom appears to own _____ valence electrons. By itself, an S atom has _____ valence electrons. Therefore, the formal charge on the S atom is _____.

The C atom appears to own _____ valence electrons. By itself, a C atom has _____ valence electrons. Therefore, the formal charge on the C atom is _____.

The O atom appears to own _____ valence electrons. By itself, an O atom has _____ valence electrons. Therefore, the formal charge on the O atom is _____.

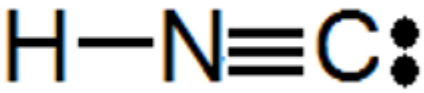
The formal charges of all atoms of a molecule should add up to [ ] since a molecule is electrically neutral. For a polyatomic ion, the sum should be equal to the charge of the ion.

Question **5**

Not yet answered

Marked out of 7.00

Isomers. In general, there is more than one way of connecting the same set of atoms to make a molecule. Each unique way is called a structural isomer; each isomer will have a unique set of properties and is considered a different molecule. For example, a molecule consisting of one H, one C, and one N atom can have two possible isomers: hydrogen cyanide (HCN) and hydrogen isocyanide (HNC).



Both molecules exist. Using formal charges, we can predict which one is more stable.

The formal charge of H in both isomers is [ ].

The formal charges of C and N in the structure shown on the left are [ ] and [ ].

For the structure shown on the right, the formal charges are [ ] for N and [ ] for C.

Therefore, we expect [ ] to be more stable than [ ]. In fact, HNC is only found in significant amounts in very cold interstellar space.

Question **6**

Not yet answered

Marked out of 22.00

Resonance Structures. For a given molecule, there may be more than one way of distributing valence electrons; each way is called a resonance structure. If there are two or more plausible resonance structures (based on formal charges), it means that none of these resonance structures is an adequate representation of reality; the real distribution of valence electrons is best thought of as a hybrid of the plausible resonance structures and structures with better formal charges are expected to have a more significant "contribution" to the hybrid. We can represent the resonance hybrid by drawing a double-headed arrow (↔) between the resonance structures.

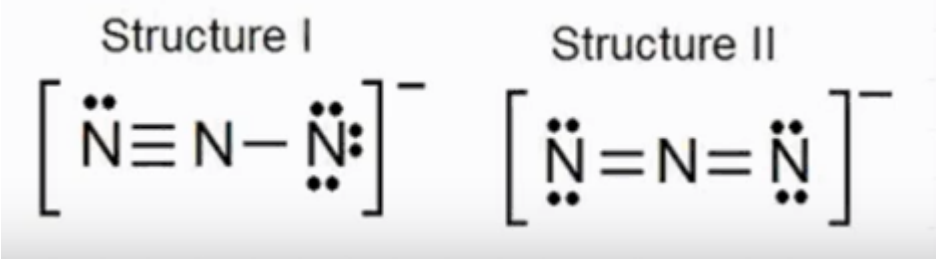
Draw resonance structures for nitrite ion (NO_2^-) and answer the questions below. You may find it useful to first watch this video:

The total number of valence electrons in nitrite ion is [ ].

With the nitrogen atom between the two oxygen atoms, it is possible to draw

- [ ] resonance structure where the N is surrounded by only [ ] valence electrons and each O is surrounded by an octet. The formal charges for this structure are: [ ] for N and [ ] for each O.
- [ ] resonance structures where all the atoms are surrounded by an octet. Both of these have [ ] single bond, [ ] double bond, [ ] lone pair on the N, [ ] lone pairs on the singly bonded O, and [ ] lone pairs on the doubly bonded O. By symmetry, both of these structures are equally plausible suggesting that [ ] one is an adequate representation of reality. However, these two structures are better than the one where the N is only surrounded by six valence electrons if we examine the formal charges. For these two, the formal charges are: [ ] for N, [ ] for the doubly bonded oxygen, and [ ] for t singly bonded oxygen.

Use formal charges to determine which of the two resonance structures of azide drawn below is better. You may find it useful to watch this video first:



For structure I, the formal charges on N (from left to right) are [ ], [ ], and [ ].

For structure II, the formal charges on N (from left to right) are [ ], [ ], and [ ].

The better structure is [ ].

Question **7**

Not yet answered

Marked out of 9.00

Strategies for Drawing Lewis Structures. H and F can only form one single bond. By forming a single bond, H acquires [ ] and F acquires [ ]. When either H or F forms a single bond, the atom gets a formal charge of [ ]. Draw Lewis structures for H_3O^+ and OF_2 and answer the questions below. You may find it useful to watch this video:

In the Lewis structure for hydronium ion, how many H atoms are bonded to the O? [ ]

In the Lewis structure for oxygen difluoride molecule, how many F atoms are bonded to the O? [ ]

When drawing Lewis structures, H and F should be drawn as *terminal* atoms; this means that they must be on the edge of the structure as each one is bonded to [ ] atom. H and F are said to be univalent.

The other halogens (Cl, Br, and I) because they have [ ] valence electrons (just like F), also *tend* to form just [ ] single bond (just like F). However, they can have more than one bonding pair because they are in rows 3, 4, and 5 of the periodic table; their outermost shell [ ] accommodate more than 8 electrons.

Question **8**

Not yet answered

Marked out of 41.00

Strategies for Drawing Lewis Structures. You need to be familiar with the typical bonding patterns of C, N, and O, which [ ] be surrounded by more than an octet. Whenever possible, draw a structure that gives these atoms a zero formal charge. You also need to be familiar with the common ways that N and O can have formal charges of +1 and -1 in polyatomic ions. Answer the questions below. You may find it useful to watch this video first:

Because a free C atom has [ ] valence electrons, it gets a zero formal charge if it is sharing [ ] bonding pairs and has no lone pair. A C atom acquires a zero formal charge if we surround it with

- [ ] single bonds
- [ ] double bonds
- [ ] single bonds and [ ] double bond
- [ ] single bond and [ ] triple bond

Oxygen acquires a zero formal charge if we surround the O atom with

- [ ] single bonds and [ ] lone pairs, as in [ ]
- [ ] double bond and [ ] lone pairs, as in [ ]

In a polyatomic ion, O acquires a +1 formal charge if it is surrounded by [ ] bonding pairs and [ ] lone pair. It acquires a -1 formal charge if it is surrounded by [ ] bonding pair and [ ] lone pairs.

Nitrogen acquires a zero formal charge if we surround the N atom with

- [ ] single bonds and [ ] lone pair, as in [ ]
- [ ] single bond, [ ] double bond, and [ ] lone pair, as in as in [ ]
- [ ] triple bond, and [ ] lone pair, as in as in [ ]

In a polyatomic ion, N acquires a +1 formal charge if it is surrounded by [ ] bonding pairs and no lone pair, as in

- [ ] single bonds
- [ ] double bonds
- [ ] single bonds and [ ] double bond
- [ ] single bond and [ ] triple bond

N acquires a -1 formal charge if it is surrounded by [ ] bonding pairs and [ ] lone pairs.

Atoms in the same column of the periodic table have a similar bonding pattern. For example, Si will behave like [ ], P will be have like [ ], and S will behave like [ ]. However, since these atoms are in third row, it is also possible for them to be surrounded by more than an octet.

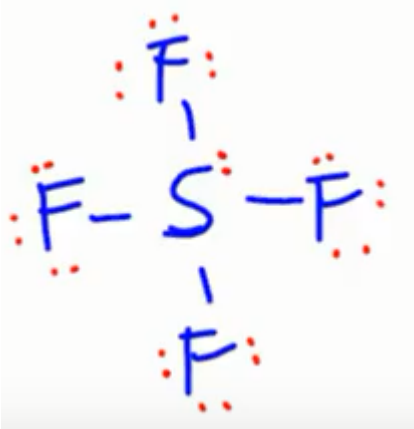
Question **9**

Not yet answered

Marked out of 42.00

Type AX_nE_m molecules and ions. A molecule or ion is said to be of type AX_nE_m if it has a central atom (A, usually written first) bonded to n terminal atoms (X; also called ligands), where n>1, and A has m lone pairs. If m=1, we can omit it; AX_nE_1 can also be referred to as AX_nE. If A has no lone pair (m=0), we can just say it is of the type AX_n instead of AX_nE_0.

Example: consider the Lewis structure of SF_4 as drawn below:



As expected, the structure shows a total of [ ] valence electrons, with the S atom contributing [ ] and each of the F atoms contributing [ ].

The central atom in this structure is [ ], which is shown to have [ ] lone pair. The [ ] atoms are the ligands. We say that this molecule is of type $AX_{\ }E_{\ }$.

Since S belongs to period 3, we [ ] allowed to surround it with more than an octet; as drawn above, the S atom has is surrounded by [ ] valence electrons. None of the F atoms can accommodate an additional lone pair since each F atom, as drawn, is already surrounded by an octet and F, which belongs to period 2, [ ] be surrounded by more than an octet.

In the structure, S appears to own [ ] valence electrons; since a free S atom has [ ] valence electrons, the formal charge for S is [ ].

In the structure, F appears to own [ ] valence electrons; since a free F atom has [ ] valence electrons, the formal charge for F is [ ].

For structures of type AX_nE_m, it is straightforward to determine n (which is also known as the *coordination number* of the central atom; the number of atoms bonded to the central atom). For example, we know that for SF_4, the coordination number n is [ ]. But how can we figure out the number of lone pairs (m)? Here's a general procedure:

- Connect the central atom to the ligands with single bonds.
- Add lone pairs to the ligands to give each one an octet (except if the ligand is H, which can only have a duet). If there are electrons left over, add these to the central atom as lone pairs.
- Check the formal charge of the central atom. If it's equal to the charge of the entire structure, is zero, or negative, you're done. If not, try lowering the formal charge by removing a lone pair from a ligand and adding a bonding pair between that ligand and the central atom. Repeat, if necessary. Don't forget that a central atom belonging to period 2 cannot be surrounded by more than an octet. The central atom does not need to have an octet either.

Answer the following regarding the Lewis structures for nitrite. You may find it useful to watch this video first.

In nitrite, the central atom is [ ], the ligands are the [ ] atoms. The total number of valence electrons is [ ]. After assigning an octet to each O atom and assigning a single bond between the N and each of the O atoms, we find that the N atom would only be surrounded [ ] valence electrons. That would give N a formal charge of [ ], while the two O atoms each have a formal charge of [ ]. By removing a lone pair from one of the O atoms and adding a second bonding pair between that O atom and the N atom, both the N and O end up with a formal charge of [ ], which is better; the formal charge on the other O atom remains as [ ]. The structure, so far, which can be described as having [ ] single bond and [ ] double bond is just one of two possible resonance structures. There is no reason why a double bond would be preferably formed with either one of the two O atoms. The structure of nitrite is of type [ ].

Answer the following questions regarding the Lewis structure of BeH_2 and BF_3. You might find it us to first watch this video:

The structure of BeH_2 is of type [ ]. The central atom [ ] surrounded by an octet, but has a formal charge of [ ].

The structure of BF_3 is of type [ ]. Although it is possible for BF_3 to be drawn with an octet around all the atoms, the B in this case would have a formal charge of [ ]. In the structure where B is surrounded by [ ] valence electrons, the B would have a formal charge of zero, which is better.

Answer the following questions regarding the Lewis structure of sulfate ion, SO_4^{2-}. You may find it useful to watch this video first:

We can account for all [ ] valence electrons by drawing single bonds between S and each of the four O atoms while putting [ ] lone pairs on each of the four O atoms. However, this structure would give S a formal charge of [ ], which is unacceptable. We can lower the formal charge on the S to zero by removing one lone pair each from [ ] O atoms, and converting the single bonds between these O atoms and S to [ ] bonds. The two O atoms also end up with a formal charge of [ ], while the other two each have a formal charge of [ ],

Question **10**

Not yet answered

Marked out of 55.00

A Lewis structure is a two-dimensional representation of a molecule or ion; its intent is to show how the atoms are connected and how valence electrons are distributed in the molecule. A Lewis structure does not necessarily have to depict the actual shape of the molecule or ion but we can use it to deduce what the actual shape is. Since atoms are _____ in motion, when we talk about the shape of a molecule or polyatomic ion, we mean the shape based on the average location of the atoms.

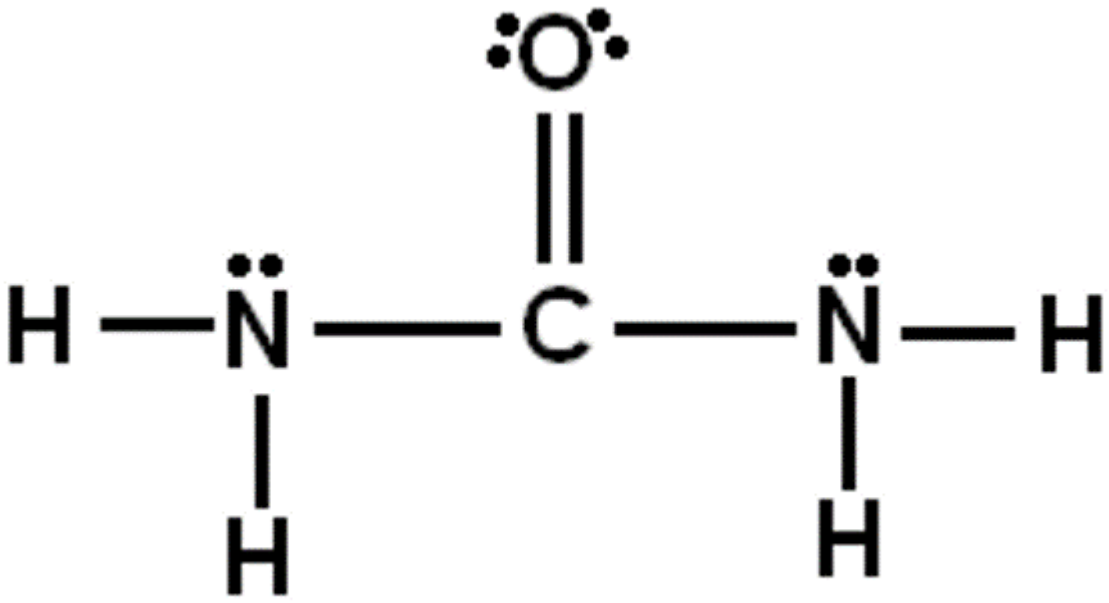
Answer the questions below. You my find it useful to watch this video:

To determine the shape of a molecule or ion we use a simple theory called VSEPR, which stands for Valence Shell Electron Pair Repulsion. The basic idea behind VSEPR is that regions of high electron density around each atom in the molecule tend to stay as far away from one another as possible. These regions of high electron density are often referred to as *electron domains* or *electron groups*.

The justification for the idea is simple: electrons _____ each other.

The term *steric number* refers to the number of electron domains around an atom. The number of atoms bonded to an atom (n) is called the atom's *coordination number*. If atom A, with coordination number n, has m lone pairs, then its steric number is n+m. In other words, each set of bonding pairs (whether single, double, or triple) counts as one domain (or group) and each lone pair counts as one domain.

Consider the Lewis structure below.



There are [ ] atoms bonded to the C atom; therefore, its coordination number is [ ]. The number of lone pairs on the C atom is [ ]. Therefore the number of electron domains around C is [ ]; its steric number is is [ ].

For each of the two N atoms, the coordination number is [ ] and the number of lone pairs is [ ]; the steric number is [ ].

Enter the steric number for atom A in molecules or ions of the type AX_nE_m listed below; these are molecules or ions with one central atom (A), with m lone pairs, surrounded by n terminal atoms or ligands (X) Note that if E_m is not specified, it is implied to be E_0; if m is not specified, it is implied to be 1.

Type	Coordination number (n)	Number of Lone Pairs (m)	Steric Number (n+m)
AX_2			
AX3			

AX_2			
AX_4			
AX_3E			
AX_2E_2			
AX_5			
AX_4E			
AX_3E_2			
AX_2E_3			
AX_6			
AX_5E			
AX_4E_2			
AX_3E_3			
AX_2E_4			

Question **11**

Not yet answered

Marked out of 28.00

Answer the questions below. You may find it useful to watch this video first:

If the steric number of an atom is 2, then the two electron domains will be as far away from each other if they are sticking out of the atom in opposite directions. We say that the angle between the electron domains is [ ] degrees. The atom is said to be a *linear* center. The *electron domain geometry* (also called electron pair geometry, or VSEP geometry, or just electron geometry) around the atom is said to be linear.

If the steric number of an atom is 3, then the three electron domains will be as far away from each other if they are sticking out of the atom towards the corners of a triangle. The atom is said to be a *trigonal planar* center. The electron domain geometry around the atom is said to be trigonal planar. If the electron domains are identical, we expect the electron domains to be directed towards the corners of an equilateral triangle and the angle between any pair of electron domains is [ ] degrees. If the electron domains are not identical, we observe a slight deviations from this angle.

The central atom of a molecule or ion of type AX_2 is a [ ] center and the *molecular shape* is described as linear.

The central atoms of a molecule or ion of type AX_3 and AX_2E are [ ] centers. However, the molecular shape is described as *trigonal planar* for type AX_3, but *bent* for type AX_2E. The molecular shape description is based on the locations of the atoms.

If a central atom has two single bonds and no lone pair, it is of type [ ], its steric number is [ ], it is a [ ] center. The angle between the two bonds is [ ] degrees.

If a central atom has two double bonds and no lone pair, it is of type [ ], its steric number is [ ], it is a [ ] center. The angle between the two bonds is [ ] degrees.

If a central atom has one single bond, one triple bond, and no lone pair, it is of type [ ], its steric number is [ ], it is a [ ] center. The angle between the two bonds is [ ] degrees.

If a central atom has three single bonds and no lone pair, it is of type [ ], its steric number is [ ], it is a [ ] center. The angle between any two bonds is closest to [ ] degrees.

If a central atom has two single bonds, a double bond, and no lone pair, it is of type [ ], its steric number is [ ], it is a [ ] center. The angle between any two bonds is closest to [ ] degrees.

If a central atom has a double bond, a single bond and one lone pair, it is of type [ ], its steric number is [ ], it is a [ ] center. The angle between the two bonds is closest to [ ] degrees. The shape of this molecule is [ ]

Question **12**

Not yet answered

Marked out of 21.00

Answer the questions below. You may find it useful to watch this video first.

If there are four electron domains around an atom, we say its steric number is [ ] and the atom is a *tetrahedral* center or that the electron domain geometry around the atom is tetrahedral. In Greek, *tetra* means [ ] and *hedron* means side; a tetrahedron is a 4-sided solid; imagine four equilateral triangles put together to make a pyramid. If the electron domain geometry is tetrahedral, then the electron domains are directed towards the corners of a tetrahedron and the angle between any pair of domains is about 109.5 degrees; slight deviations from this angle is observed if the electron domains are not identical.

Molecules or ions with a steric number of 4 can be type AX_4, AX_3E, and AX_2E_2. All of these types have a [ ] center, but the *molecular shapes* are said to be *tetrahedral* for AX_4, *trigonal pyramidal* for AX_3E, and *bent* for AX_2E_2. The molecular shape description is based on the locations of the atoms.

A methane molecule (CH_4) has [ ] C-H single bonds and no lone pair. This would be of type [ ]. The C atom can be classified as a [ ] center. Since there is an atom (H) at the other end of all four electron domains, the shape of the molecule is [ ].

An ammonia molecule (NH_3) has [ ] N-H single bonds and [ ] lone pair. This would be of type [ ]. The N atom can be classified as a [ ] center. However, since three domains have an atom on the other end, the shape of the molecule is [ ]

A water molecule (H_2O) has [ ] O-H single bonds and [ ] lone pairs. This would be of type [ ]. The O atom can be classified as a [ ] center. However, since only two domains have an atom on the other end, the shape of the molecule is [ ].

The bond angles in NH_3 and H_2O are 107 and 104.5 degrees, respectively. These are slightly [ ] than the ideal tetrahedral bond angle of [ ]. An explanation for this is that the lone pairs are [ ] than the single bonds and will push them away more strongly; the single bonds, therefore, end up being [ ].

Question **13**

Not yet answered

Marked out of 20.00

Answer the questions below. You may find it useful to watch this video first.

For molecules or ions of type AX_5, AX_4E, AX_3E_2, and AX_2E_3, the steric number of central atom A is [ ]. The central atom is said to be a *trigonal bipyramidal* center; imagine it being at the center of two pyarmids with a common triangular base (one pyramid pointing upward, the other pointing downward). Three of the five electron domains are directed towards the corners of a triangle (at the common base of the two pyramids) and are called the *equatorial* groups; the other two are directed in opposite directions perpendicular to the plane of the triangle and are called the *axial* groups.

If all electron five electron domains are identical, then

- the angle between any pair of equatorial groups is [ ] degrees
- the angle between the two axial groups is [ ] degrees.
- the angle between an axial group and an equatorial group is [ ] degrees.

In general, there will be slight deviations from these angles.

Since molecular shape is based on the locations of atoms in a molecule,a molecule with a trigonal bipyramidal center also has a trigonal bipyramidal shape only if there is an atom at the other end of all five electron domains, which means that it must be of type [ ]. If there are lone pairs, the lone pairs prefer to take the equatorial positions and the molecular shape can be described as *seesaw* (AX_4E), *T-shape* (AX_3E_2), or *linear* (AX_2E_3).

PCl_5 has [ ] single bonds and no lone pair. The electron domain geometry around the P atom is [ ], the shape of the molecule is [ ].

SF_4 has [ ] single bonds and [ ] lone pair. The electron domain geometry around the S atom is [ ], the shape of the molecule is [ ].

BrF_3 has [ ] single bonds and [ ] lone pairs. The electron domain geometry around the Br atom is [ ], the shape of the molecule is [ ].

ICl_2^- has [ ] single bonds and [ ] lone pairs. The electron domain geometry around the I atom is [ ], the shape of the molecule is [ ].

Question **14**

Not yet answered

Marked out of 15.00

Answer the questions below. You may find it useful to watch this video first.

For molecules or ions of type AX_6, AX_5E, AX_4E_2, AX_3E_3, and AX_2E_4, the steric number of central atom A is [ ]. The central atom is said to be an *octahedral* center; imagine it being at the center of two pyramids sharing a common square base (one pyramid pointing upward, the other pointing downward); the two pyramids together form a 8-sided figure (an octahedron). If all six electron domains are identical, then each electron domain is directed opposite one other domain and is perpendicular to four other domains. In general, there will be slight deviations from these angles.

The angle between one domain and any one of its four nearest neighboring domains is [ ] degrees.

The angle between one domain and the domain pointing in the opposite direction it is [ ] degrees.

Since molecular shape is based on the locations of atoms in a molecule,a molecule with an octahedral center also has a octahedral shape only if there is an atom at the other end of all six electron domains, which means that it must be of type [ ]. If there are lone pairs, the molecular shape can be described as *square pyramidal* (AX_5E), *square planar* (AX_4E_2), *T-Shape* (AX_3E_3), and *linear* (AX_4E_2).

SF_6 has [ ] single bonds and no lone pair. The electron domain geometry around the S atom is [ ], the shape of the molecule is [ ].

BrF_5 has [ ] single bonds and [ ] lone pair. The electron domain geometry around the Br atom is [ ], the shape of the molecule is [ ].

ClF_4^- has [ ] single bonds and [ ] lone pairs. The electron domain geometry around the Cl atom is [ ], the shape of the molecule is [ ].

Question **15**

Not yet answered

Marked out of 29.00

Valence Bond Theory explains bonding as being due to the overlapping of atomic orbitals of the bonded atoms. To explain the orientations of electron groups sticking out of an atom, the valence orbitals of the atom are mathematically combined (a process called *linear combination*). Consider a C atom that has four single bonds. According to valence bond theory, the C atom in the molecule no longer has one 2s and three 2p orbitals; instead it has four hybrid orbitals (called sp^3) which describe regions of high electron density that are oriented 109.5 degrees apart. The number of hybrid orbitals obtained is equal to the number of atomic orbitals that are used to construct the hybrids.

Complete the table and answer the questions below. You may find it useful to watch this video:

Steric Number	Hybridization	Number of s orbital used	Number of p orbitals used	Number of d orbitals used	Total Number of Hybrid Orbitals
2	sp				
3	sp^2				
4	sp^3				
5	sp^3d				
6	sp^3d^2				

The angle between two sp hybrids is _____ degrees.

The angle between any two sp^2 hybrids is _____ degrees.

The angle between any two sp^3 hybrids is _____ degrees.

If an sp^3d hybrid orbital has a 120-degree angle with 2 other sp^3d hybrids, it also has a [ ] degree angle with [ ] other sp^3d hybrids.

If an sp^3d hybrid orbital has a 180-degree angle with another sp^3d hybrid, it also has a [ ] degree angle with [ ] other sp^3d hybrids.

Each sp^3d^2 hybrid orbital has a [ ] degree angle with four other sp^3d^2 hybrids and a [ ] degree angle with one other sp^3d^2 hybrid.

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