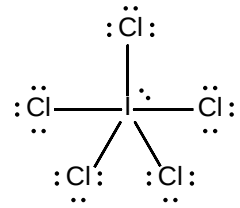
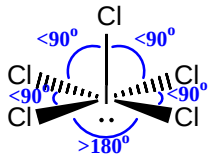
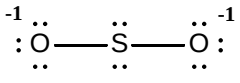
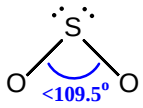
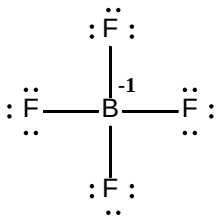
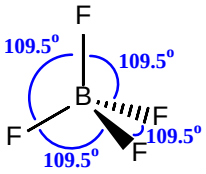
Answers to Practice Test Questions 8
Lewis Diagrams and VSEPR Geometries

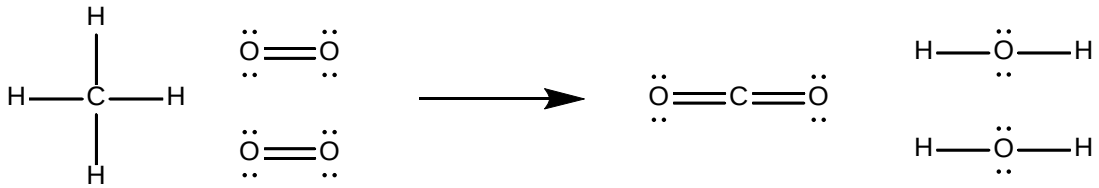
1.

Chemical Formula	Lewis Diagram	Electron Group Geometry (name)	Molecular Geometry (name)	Drawing of Molecule's Geometry with Bond Angles Labeled
ICl_5		octahedral	square pyramidal	
SO_2^{2-}		tetrahedral	bent	
BF_4^-		tetrahedral	tetrahedral	

Note that an angle must involve three atoms. Two points gives a straight line – not an angle!

2. The enthalpy change for a reaction can be estimated by subtracting the heat released by forming bonds from the heat consumed by breaking bonds

Step 1: Draw Lewis diagrams for all molecules in the balanced reaction equation



Step 2: Identify which bonds will be formed and which bonds will be broken

Bonds formed: two C=O and four O-H

Bonds broken: two O=O and four C-H

Step 3: Calculate the enthalpy change from forming and breaking bonds

$\Delta H \approx \sum \Delta_{bd}H(bonds\ broken) - \sum \Delta_{bd}H(bonds\ formed)$

$\Delta H \approx [4\Delta_{bd}H(C-H) + 2\Delta_{bd}H(O=O)] - [4\Delta_{bd}H(O-H) + 2\Delta_{bd}H(C=O)]$

$\Delta H \approx \left[4\left(413\frac{kJ}{mol}\right) + 2\left(498\frac{kJ}{mol}\right)\right] - \left[4\left(467\frac{kJ}{mol}\right) + 2\left(799\frac{kJ}{mol}\right)\right]$

$\Delta H \approx -818\frac{kJ}{mol}$

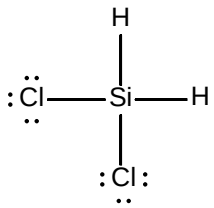
Step 4: Check your work

Does your answer seem reasonable? Are sig. fig. correct?

We know that burning natural gas (methane) is exothermic, so it makes sense that the enthalpy change for this reaction is negative.

3.

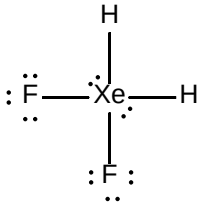
(a)



tetrahedral electron group geometry

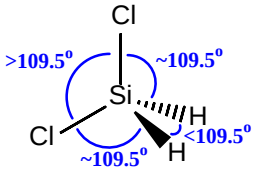
tetrahedral molecular geometry

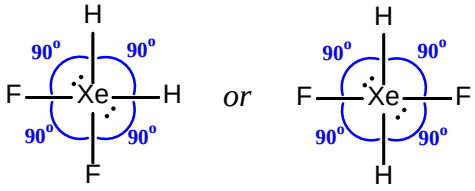
(b)



octahedral electron group geometry

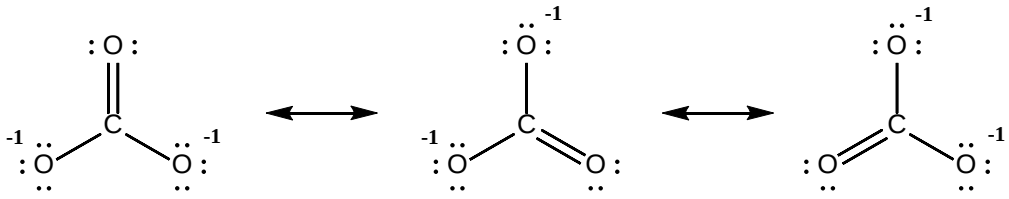
square planar molecular geometry





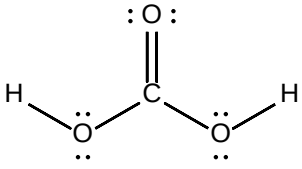
4.

(a)



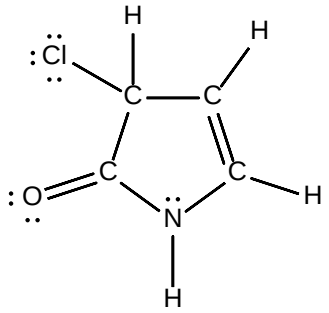
(b) bond order $= \frac{2+1+1}{3} = 1\frac{1}{3}$

(c)



The hydrogen atoms MUST be bonded to oxygen – not to carbon. It is logical that the H^+ would attach themselves to the O^- of the CO_3^{2-} anion.

5.



6. Fluorine is a second period element, so it cannot exceed an octet of valence electrons. As such, a fluorine bonded to two different atoms could have no more than two lone pairs of electrons. This would result in a fluorine atom with a charge of +1:

$$Q_f = 7 - [\tfrac{1}{2}(4) + (4)] = +1$$

Fluorine is the most electronegative element and therefore cannot exist in molecules which would force it be assigned a positive formal charge.

Thus, molecules containing X-F-Y are highly unusual.

7.

(a) *Your answer must have two atoms connected by a single bond, and both atoms must obey the octet rule (or be hydrogen with one bond and no lone pairs).*

e.g. H_2, F_2, Cl_2, Br_2, I_2, HF, HCl, HBr, HI

(b) *Your answer must have two atoms connected by a double bond, and both atoms must obey the octet rule.*

e.g. O_2

(c) *Your answer must have two atoms connected by a triple bond, and both atoms must obey the octet rule.*

e.g. N_2

8.

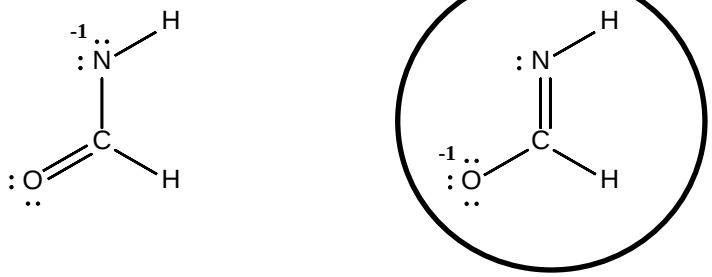
(a) *Your answer must have a central atom with one lone pair and three atoms attached.*

e.g. NH_3, NF_3, PF_3, PCl_3

(b) *Your answer must have a central atom with no lone pairs and three atoms attached.*

e.g. BH_3, BF_3, BCl_3, BBr_3, CH_2O

9.



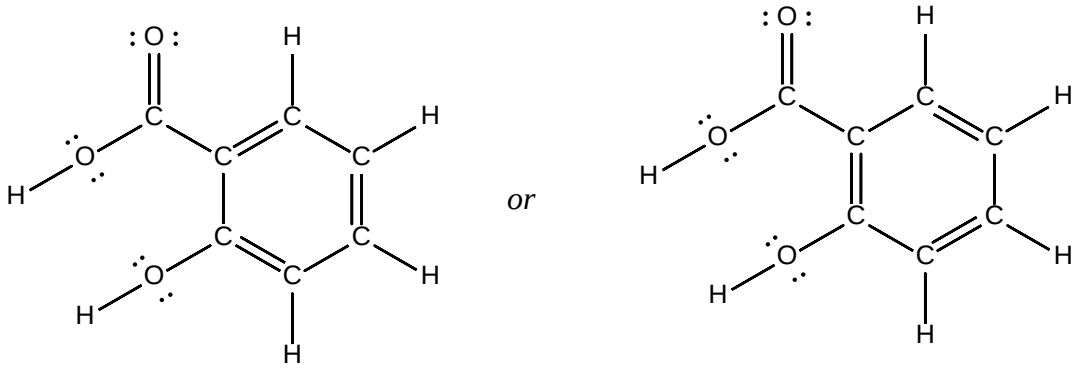
(a) *see above*

(b) *see above*

O is more electronegative than N and they are similarly sized atoms, so the resonance structure with the negative charge on O will dominate.

10.

(a)



(b) 2

the two resonance structures are the two options shown above

(c) O-H

H is a very small atom, so the C-H and O-H bonds should be shortest.
O is smaller than C, so expect O-H to be shorter than C-H.

(d) C-C (the one that is a single bond in both resonance structures)

C and O have similar radii (both in period 2), so we look at bond order. Assuming atoms with similar radii, bond length decreases with increasing bond order so the C-C and C-O single bonds should be longest. C has a larger radius than O, so C-C is expected to be longer than C-O.

(e) (i) trigonal planar

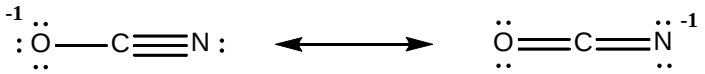
(ii) trigonal planar

(iii) tetrahedral

(iv) bent

11.

(a)



(b) The major resonance structure will be the one with the negative charge on O. This is because O is more electronegative than N (and the sizes of N and O are similar).

(c) The C-O bond order is between 1 and 1.5.

The C-N bond order is between 2.5 and 3.

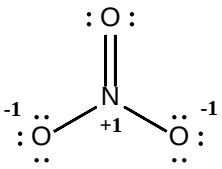
The charge on O is between -0.5 and -1.

The charge on N is between 0 and -0.5.

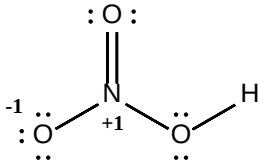
We could use molecular modeling (computer calculations) to obtain exact answers if we needed them. For most purposes, knowing that the charge on O is "a bit bigger than -0.5" is good enough. You will discuss computational chemistry in CHEM 2000. It's useful for much more than just calculating these types of values!

12.

(a)



(b)



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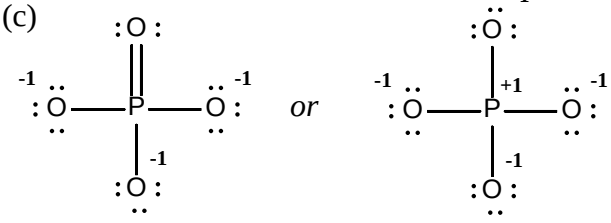
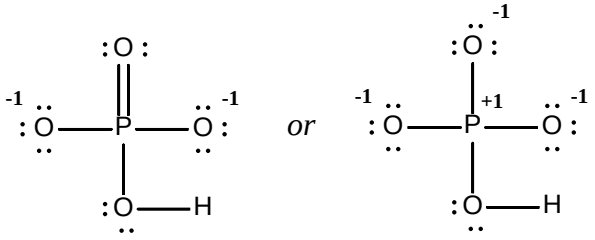
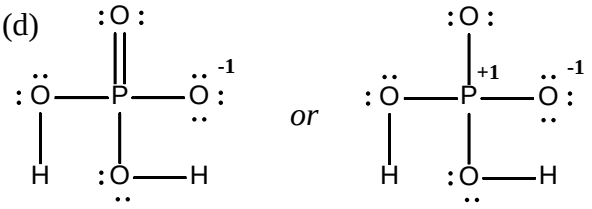


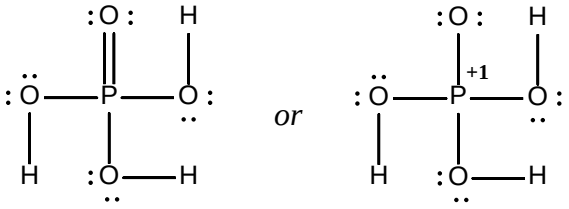
(d)



(d)



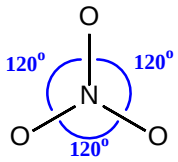
(e)



*When attaching H^+ to an oxoanion (e.g. making HNO_3 from NO_3^-), attach it to an O^-! There are *very* few oxoacids in which H is attached to the central atom. It's almost always attached to O!*

13.

(a)



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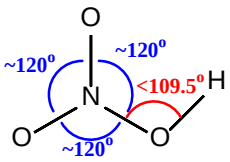

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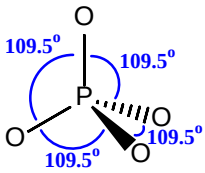
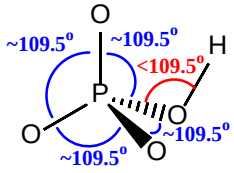


(d)



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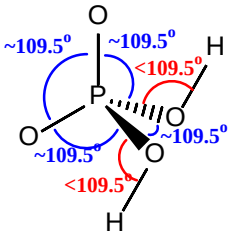

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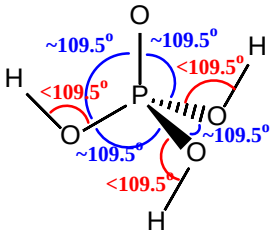


Note that:

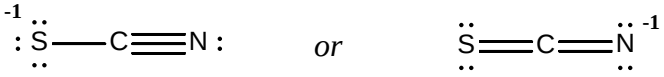
- *It does not matter which convention you used to draw your Lewis diagrams. VSEPR predicts the same geometry from either Lewis diagram.*
- *There are three resonance structures for NO_3^-. As such, all three N-O bonds have the same bond order and the same bond length. The molecule is therefore symmetrical and all three bond angles are exactly 120˚.*
- *If you use the "minimize formal charge" convention for drawing PO_4^{3-}, there are four resonance structures. As such, all four P-O bonds have the same bond order and the same bond length. The molecule is therefore symmetrical and all four bond angles are exactly the same.*
- *When H^+ is attached to one of the oxygen atoms (which was formerly O^-), that breaks the symmetry. As soon as that happens, the bond angles no longer perfectly match the official values for trigonal planar (120˚) or tetrahedral (109.5˚).*
- *The two lone pairs on the bent oxygen atoms make the X-O-H bond angles less than 109.5˚.*

14. *All three atoms should obey the octet rule.*

None of these elements can make a quadruple bond!

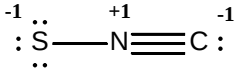
Formal charges must be shown.. You can't choose between the options without looking at formal charge!

(a)



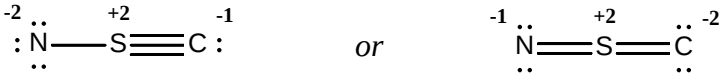
Both are good Lewis structures. N is slightly more electronegative than S (3.0 vs. 2.5) so you might prefer to put the negative charge on N for that reason; however, S is larger than N so it will have a lower charge density and it can be argued that it should therefore have the negative charge. In reality, the two atoms share the negative charge and the real structure for this ion is an average of these two resonance structures.

(b)



When you must have formal charge, it is best not to concentrate it all on one atom. So, this resonance structure is favoured over alternatives in which N has a charge of +1 and one of the other atoms has a charge of -2. (two atoms with -1 is better than one with -2)

(c) *This was the version for which none of the options looks great. Several answers are acceptable including:*



The diagram on the left is better as it puts more negative charge on N than C (and N is more electronegative than C), but they're both pretty bad from a formal charge perspective. They are, however, as good as it gets for this connectivity.

(d) Carbon is the central atom.

This gives the Lewis diagram with the least formal charge.