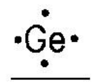
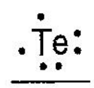
SOLUTIONS TO THE PRACTICE PROBLEMS FOR MODULE #8

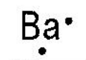
1. a. Ge is in group 4A, so it has 4 valence electrons:



b. Te is in group 6A, so it has 6 valence electrons:



c. Ba is in group 2A, so it has 2 valence electrons:



2. a. Al is in group 3A, so it wants a charge of 3+. Sulfur is in group 6A, so sulfide will have a charge of 2-. Ignoring the signs and switching the numbers gives us Al_2S_3.

b. Cs is in group 1A, so it wants a charge of 1+. Nitrogen is in group 5A, so nitride will have a charge of 3-. Ignoring the signs and switching the numbers gives us Cs_3N.

c. Mg is in group 2A, so it wants a charge of 2+. Oxygen is in group 6A, so oxide will have a charge of 2-. The numbers are the same so we ignore them: MgO.

d. Cr is an exception, because there is a Roman numeral in the name. The numeral tells us that Cr want a charge of 3+. Oxygen is in group 6A, so oxide will have a charge of 2-. Ignoring the signs and switching the numbers gives us Cr_2O_3.

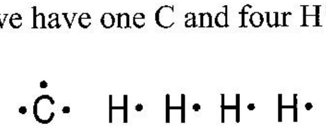
3. Ionization potential decreases as you go down the chart. The atom with the lowest ionization potential will give up its electrons easiest. That would be In.

4. Ionization potential increases as you go from left to right on the chart, so the order is Sr < Sb < I.

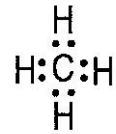
5. Electronegativity decreases as you go down the chart, so N has the greatest desire for extra electrons.

6. Atomic radius decreases as you go from left to right on the chart, so the order is Br < Se < As < K.

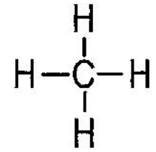
7. The chemical formula tells us that we have one C and four H's to work with:



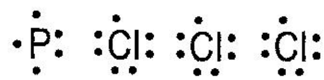
Because C has the most unpaired electrons, it goes in the center and we try to attach the H's to it. This is easy since each H has a space for an unpaired electron, and the C has four unpaired electrons. The Lewis structure, then, looks like this:



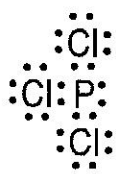
All atoms have their ideal electron configuration, so we are all set. Now we just have to replace the shared electron pairs with dashes:



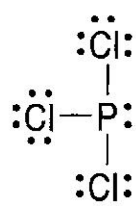
8. The chemical formula tells us that we have one P and three Cl's to work with:



Because P has the most unpaired electrons, it goes in the center and we try to attach the Cl's to it. This is easy since each Cl has a space for an unpaired electron, and the P has three unpaired electrons. The Lewis structure, then, looks like this:

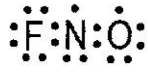


All atoms have their ideal electron configuration, so we are all set. Now we just have to replace the shared electron pairs with dashes:



9. The chemical formula tells us that we have one F, one N, and one O to work with:

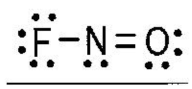
The N has the most unpaired electrons, so it goes in the middle. We attach the others to it:



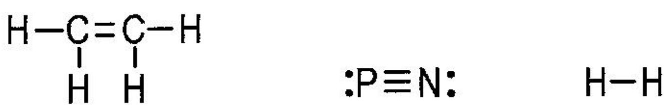
The F now has eight electrons, so its all set. The N and O, however, have only seven each. We will give the O eight by taking the unpaired electron on the N and putting it in between the N and the O:

We can also give the N its eight by taking the unpaired electron on the oxygen and moving it in between the N and the O.

Now all atoms have eight valence electrons. All we have to do is replace the shared electron pairs with dashes:



10. To solve this, we have to determine the Lewis structures of each molecule. They turn out to be:



Based on these three Lewis structures, the H_2 will be easiest to break apart because it is held together by a single bond, while the C_2H_4 is held together by a double bond, and the PN is held together by a triple bond.