

## REVIEW QUESTIONS

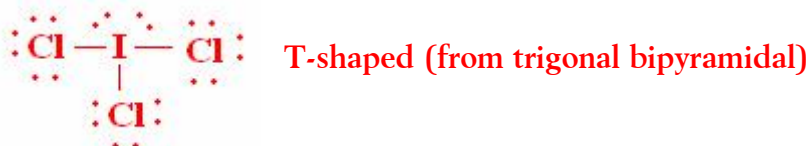
## Chapter 10

1. Draw Lewis structures and determine the molecular geometry of each molecule or ion shown below:

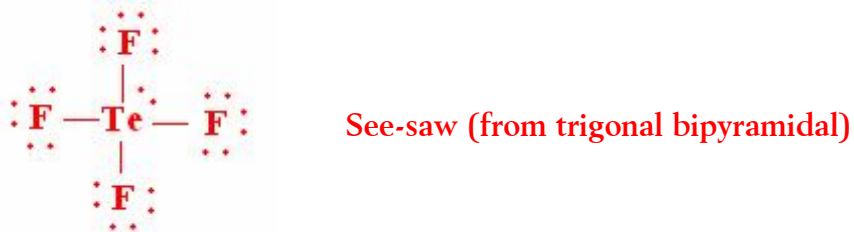
A)  $\text{ClO}_2^-$       **20 electrons**



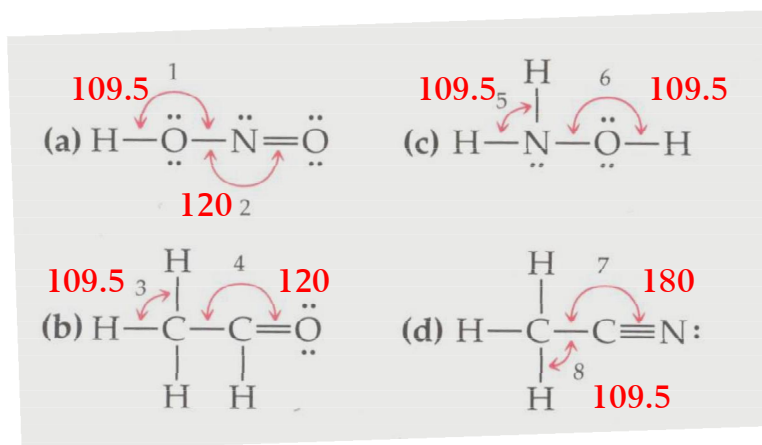
B)  $\text{ICl}_3$       **28 electrons**



C)  $\text{TeF}_4$       **34 electrons**



2. Give approximate value for each bond indicated in the molecules shown below:



3. Determine if each molecule below would be polar or non-polar. Give a brief explanation for your choices.

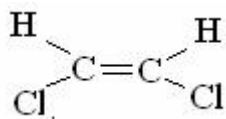
A)  $\text{CS}_2$       **Non-polar**      **(linear)**

B)  $\text{SO}_3$       **Non-polar**      **(trigonal planar)**

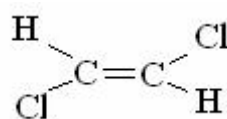
C)  $\text{SF}_4$       **Polar**      **(see-saw)**

D)  $\text{IF}_5$       **Polar**      **(square pyramidal)**

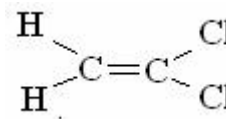
4. Dichloroethylene,  $\text{C}_2\text{H}_2\text{Cl}_2$ , can have any one of the geometries shown below, each of which is an individual substance.



(A)



(B)



(C)

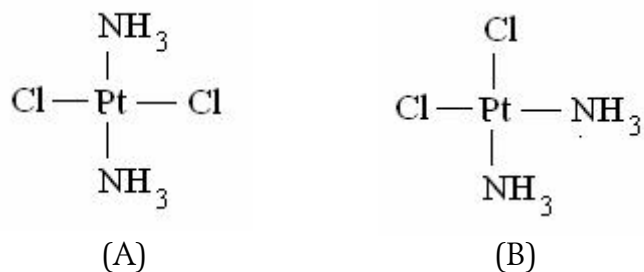
a) What is the geometry around each carbon atom in any of these molecules?

**Since each carbon has 3 groups around it, its geometry would be trigonal planar.**

b) Which of these molecules would you expect to be non-polar?

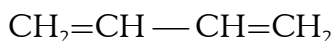
**(B) would be expected to be non-polar, since the opposite polarities of C—Cl bonds and C—H bonds cancel one another.**

5. There are two compounds with the formula  $\text{Pt}(\text{NH}_3)_2\text{Cl}_2$ . Structure A is called cisplatin and is used in cancer therapy. Both molecules have square-planar geometry. Which structure do you expect to be polar? Explain.

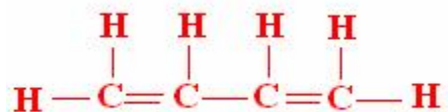


Since both molecules have square planar geometry, the opposite polarities of Pt—Cl bonds and Pt—NH<sub>3</sub> bonds cancel one another in A, but cannot cancel in B. Therefore A would be expected to be non-polar, and B would be polar.

6. Butadiene,  $\text{C}_4\text{H}_6$ , is an important molecule found in natural rubber, and has the following structural formula:



Determine the bond angle around each carbon and sketch the molecule showing its actual structure. Is this molecule planar or not? (Hint: build a model).



Based on the Lewis structure (above), each carbon has a trigonal planar geometry and an angle of  $120^\circ$ . The actual structure would be as shown below:



Since all the carbons are trigonal planar geometry, the entire molecule is therefore planar.