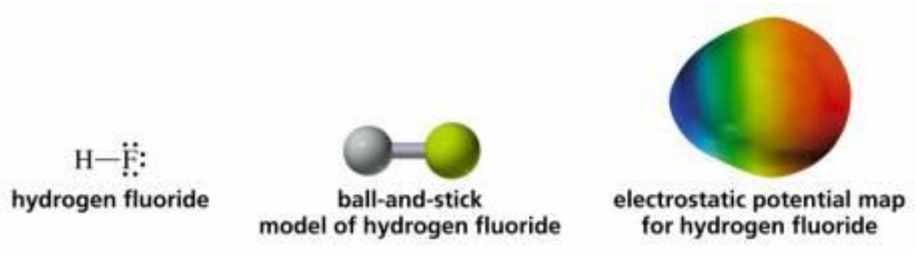
Question #83953, Chemistry / Inorganic Chemistry | for completion

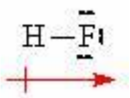
With three named examples explain how hybridization of atomic orbitals May influence the shape and bond angles of chemical entities

Answer:

We can use Lewis dot structures to determine bonding patterns in molecules. We can then use VSEPR to predict molecular shapes, based on the valence electron pairs of the Lewis structures. Once we know a molecular shape, we can start to look at the **physical properties** of compounds. For example, we should now be able to predict which molecules will be **polar**. **Polarity** exists when there is a **separation ofcharge** within a molecule. This will arise from **polar bonds** within the molecule, due to differences in **electronegativity** values between bonded atoms. For example, HF is a **polar compound**. Fluorine is much more electronegative than hydrogen and the shared pair of bonding electrons will spend more time near the F nucleus, than near the H nucleus.



The direction of a **dipole moment** (charge imbalance) is usually indicated by the presence of an arrow, as shown below for HF.



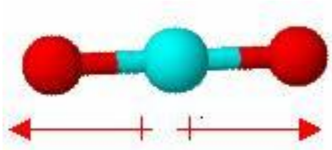
This indicates that the H will carry a partial positive charge (δ+) and the F will carry a partial negative charge (δ -). All **diatomic** molecules containing atoms of different electronegativities will be **polar** molecules. This will affect their physical properties (melting and boiling points, solubilities, etc.).

In larger molecules (more than two atoms), the polarity of the overall compound will be determined by the presence of **polar bonds and** the **molecular shape**.

For example, we can compare carbon dioxide, CO_2 to sulfur dioxide, SO_2. Their Lewis structures are shown below.

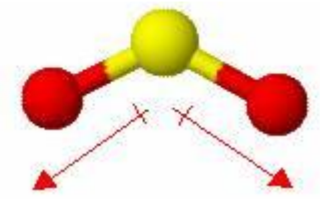
Carbon and sulfur have the same electronegativity, much less than that of oxygen. So, in both compounds the bonds will be equally polar. However, they have very different physical properties, CO_2 boils at **-78ºC**, and SO_2 boils at **+22.8ºC**, a **100º** difference. This must depend on more than just the presence of the two polar bond in each molecule. What makes the difference is the **molecular shape**.

CO_2 will be a **linear** molecule, because there are only two electron pairs on the central carbon atom. It will have the shape shown below.



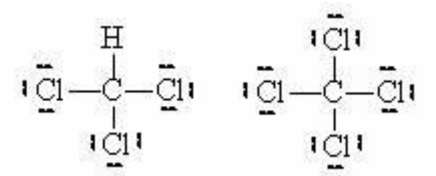
As indicated by the arrows, there are two very polar bonds in this molecule. However, because of the molecular shape of CO_2, they are pointing in opposite directions, and will **cancel out**. CO_2 is a **non-polar** compound, due to its molecular shape.

Shown below is the shape of an SO_2 molecule. Its **molecular geometry** will be **trigonal** because of the **three valence electron pairs** on sulfur, two bonding pairs and a lone pair. This will give it a **bent**molecular shape. In this molecule, the dipoles are **not** pointing in opposite directions, and will not cancel out. They will, in fact add, and give a **net dipole moment**. SO_2 is a **polar** compound, which explains (as we will soon see) its elevated boiling point.



(Quiz For a similar example, draw the Lewis structures of BF_3 (-99) and NH_3 (-33)

Another indication of the importance of molecular shape can be seen by comparing the physical properties of $CHCl_3$ and CCl_4. $CHCl_3$ dissolves in water, and CCl_4 does not. Why?



They will both have **tetrahedral** geometries, with 4 valence pairs of electrons on each C. The C-Cl bonds will all be **polar**. Their shapes are shown below.

In the $CHCl_3$ molecule, the three polar C-Cl bonds add (vector addition) to give a net dipole moment to the molecule. In CCl_4, the four polar C-Cl bonds will cancel out, making this a non-polar molecule. Water is a polar solvent, which only interacts with other polar species, "likes dissolve likes".