1. Give the expected bond angles and predict the hybrid orbitals expected for the following compounds: (4 pts)

   a) CH₂O
   
   \[
   \begin{array}{c}
   O \\
   H \quad C \quad H
   \end{array}
   \]
   
   electron geometry: trigonal planar
   
   hybrid orbitals: \(sp^2\)
   
   bond angle: 120°

   b) \(O_2\)
   
   \[
   \begin{array}{c}
   O = O
   \end{array}
   \]
   
   electron geometry: trigonal planar
   
   hybrid orbitals: \(sp^2\)
   
   bond angle: 120°

2. Use partial orbital diagrams (Valence Bond Theory) to show how the atomic orbitals of the central atom leads to hybridization for the following molecules: (4 pts)

   a) \(CO_3^{2-}\)
   
   Trigonal Planar
   
   \[
   \begin{array}{cccccc}
   2p & \uparrow & \uparrow & \uparrow & \uparrow & \uparrow \\
   2s & \uparrow & \uparrow & \uparrow & \uparrow & \uparrow \\
   \end{array}
   \]
   
   \(sp^3\)

   b) \(SF_4\)
   
   Trigonal bipyramidal (See saw)
   
   \[
   \begin{array}{cccccc}
   3d & \_ & \_ & \_ & \_ & \_ \\
   3p & \_ & \_ & \_ & \_ & \_ \\
   \end{array}
   \]
   
   \(sp^3d\)

3. Using a Molecular Orbital Diagram, calculate the bond order for carbon CO. (2 pts)

   Bond Order = \(\frac{1}{2}(8 - 2) = 3\)