To satisfy the minimum requirements for this course, you should be able to:

1. Use valence bond theory to explain how a covalent bond forms and to account for molecular geometry. You should be able to:
   - explain why bond formation is an exothermic process and bond breaking is an endothermic process
   - understand a potential energy curve of a diatomic molecule and determine the bond energy and equilibrium bond distance
   - explain the relationship between atomic orbitals and hybrid orbitals up to sp³
   - recognize names, shapes, and orientation of hybrid orbitals appropriate for central atoms surrounded by up to 4 electron pairs. (Note: In the textbook’s (Kotz et al) terminology, an “electron pair” can be a lone pair, a single bond, a double bond, or a triple bond even though a double bond consists of two electron pairs and a triple bond consists of three electron pairs. To avoid this confusion, other textbooks use the term “electron domain” rather than “electron pair”.)
   - use Lewis structures to predict the hybridization state of each central atom in a molecule and the geometry around each atom
   - distinguish between σ bonds and π bonds and be able to determine the number of sigma and pi bonds in a molecule
   - describe the delocalized pi bonding found in species such as benzene and carbonate ion and draw Lewis structures to depict the delocalized bonding.

2. Explain bonding from a molecular orbital theory point of view, and be able to:
   - recognize a molecular orbital (MO) diagram of H₂, He₂, and Li₂
   - distinguish bonding and antibonding orbitals and the effect of orbital occupation by electrons
   - explain the band theory of metals (12.4)
   - compare energy band diagrams of metallic conductors, semiconductors, and insulators (12.4)

3. NavApp: Semiconductors
   - understand that electronic devices are made from semiconductors (integrated under LO 2)