Suggested solutions for Chapter 4

PROBLEM 1
Textbooks sometimes describe the structure of sodium chloride like this ‘an electron is transferred from the valence shell of a sodium atom to the valence shell of a chlorine atom.’ Why would this not be a sensible way to make sodium chloride?

Purpose of the problem
To make you think about genuine ways to make compounds rather than theoretical ways.

Suggested solution
Of course sodium chloride consists of arrays of sodium cations without their 2s electron and chloride anions that have eight electrons in the 2s and 2p orbitals, but that is not how sodium chloride is made. Sodium atoms are present in sodium metal but where would you get the chlorine atoms? Mixing sodium and chlorine (Cl₂) would undoubtedly give sodium chloride but these are two aggressive reagents that would probably explode. Indeed, you would be more likely to make sodium and chlorine by the electrolysis of sodium chloride than the other way round. In any case, why make sodium chloride? Salt mines and the oceans are full of it.

PROBLEM 2
The H–C–H bond angle in methane is 109.5°. The H–O–H bond angle of water is close to this number but the H–S–H bond angle of H₂S is near 90°. What does this tell us about the bonding in water and H₂S? Draw an diagram of the molecular orbitals in H₂S.

Purpose of the problem
An exploration of hybridization.

Suggested solution
If the bond angle in water is close to the tetrahedral angle of perfectly symmetrical methane, water must be more or less tetrahedral (with respect to the arrangement of its electrons) too. We can think of the 2s
and 2p electrons in water as hybridized into four pairs of electrons, two in H–O bonds and two as lone pairs on the oxygen atom. But H₂S has a near right angle for its H–S–H bond. This suggests that the bonds are formed with p orbitals on the sulfur atom and that H₂S is not hybridized. Orbital diagram of H₂S: you might have drawn something like this:

PROBLEM 3

Though the helium molecule He₂ does not exist (p. 91 of the textbook explains why), the cation He₂⁺ does exist. Why?

Purpose of the problem

To encourage you to think about the filling of molecular orbitals and to accept surprising conclusions.

Suggested solution

He₂ does not exist because the number of anti-bonding electrons is the same as the number of bonding electrons. The bond order is zero. But if we remove an electron from the diagram on p. 91 of the textbook we have He₂⁺, with two bonding electrons and only one anti-bonding electron. The bond order is one half. He₂⁺ does exist.
**PROBLEM 4**
Construct an MO diagram for LiH and suggest what type of bond it might have.

**Purpose of the problem**
To demonstrate that a simple MO treatment can be applied to ionic as well as covalent structures.

**Suggested solution**
H has of course only one electron in a 1s orbital. Li has three – a full 1s shell and one electron in the 2s orbital. Li is very electropositive so its 2s orbital is high in energy—much higher than that of the 1s orbital of H. An electron is more stable in the 1s orbital of H than in the 2s orbital of Li, and the molecule is ionic. Both ions have the same electronic configuration: 1s².

![Diagram of Li and H atoms](image1)

LiH: the molecule

![Diagram of Li⁺ and H⁻](image2)

**PROBLEM 5**
What is the hybridization and shape of each carbon atom in these molecules?

![Molecules](image3)

**Purpose of the problem**
To give you practice at selecting the correct hybridization of carbon atoms.
Suggested solution

Simply count the number of $\sigma$-bonds at each carbon atom (not forgetting the hydrogens that may not be shown). Two bonds means sp and linear, three means sp$^2$ and trigonal, and four means sp$^3$ and tetrahedral. In each case the bonds stay as far from each other as they can.

![Diagram showing the sp, sp$^2$, and sp$^3$ hybridization and bond angles.]