6.5 Review Questions (p. 360)

1. (a) The letters stand for: **V**alence **S**hell **E**lectron **P**air **R**epulsion.

(b) VSEPR theory allows us to use two-dimensional Lewis diagrams for molecules to quite accurately predict the three-dimensional shapes of those molecules.

2. As lone-pair electrons are attracted to only one atomic nucleus, they are held less tightly than bonding electron groups. Their electron clouds therefore occupy more space and exert more repulsive force on bonding electron groups than those groups exert on each other.

3. Both BF_3 and CH_2O have the same shape because both of them are AX_3 molecules. Even though the peripheral atoms in each molecule possess different numbers of non-bonding electrons and one of the AX bonds on CH_2O is a double bond, the shapes are identical.

4.

AX_mE_n Notation	Molecular Shape	AX_mE_n Notation	Molecular Shape
AX_3	angular	AX_4E	T-shaped
AX_2E_3	trigonal bipyramidal	AX_2E	octahedral
AX_4	trigonal pyramidal	AX_3E_2	square pyramidal
AX_3E	trigonal planar	AX_6	square planar
AX_2E_2	tetrahedral	AX_5E	angular
AX_5	linear	AX_4E_2	seesaw

5. (a) The X – A – X bond angles in ammonia will be smaller than in methane (107° vs. 109.5°) because ammonia is an AX_3E molecule and methane is an AX_4 molecule. The lone pair on the central nitrogen in ammonia occupies more space than the bonded pairs on the central carbon in methane and will therefore force the bonded electron pairs in ammonia closer together.

(b) Methane is a symmetrical AX_4 molecule and is non-polar. The intermolecular forces acting are therefore weak London dispersion forces. Ammonia is an asymmetrical AX_3E polar molecule that exhibits much stronger hydrogen bonding.

6.

AX_mE_n Category	**AX_2**	**AX_3**	**AX_4**	**AX_5**		**AX_2E_3**	**AX_6**	**AX_4E_2**
X–A–X Bond Angle	**180°**	**120°**	**109.5°**	**120°**	**180°**	**180°**	**90°**	**90°**

7. The molecules of each compound are asymmetric AX_3E molecules. If we consider the ΔEN values as we move up the family from bottom to top, we see the following:

Compound	SbH_3	AsH_3	PH_3	NH_3
ΔEN Value	0.2	0.1	0.0	0.9

Even though asymmetry exists in each of the molecules, because the ΔEN values for the first three compounds are so small, all of the bonds in those molecules are non-polar and so the molecules themselves are non-polar. As a result, the intermolecular forces acting between them are weak London dispersion forces. However, the bonds in ammonia are quite polar and so the molecule itself is also polar. As nitrogen is bonded to hydrogen in a polar molecule, hydrogen bonds exist between the ammonia molecules. We would therefore expect ammonia to have the highest boiling point.

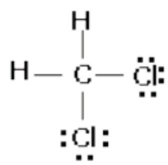
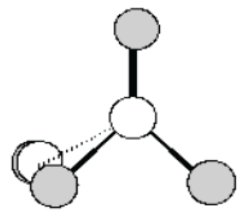
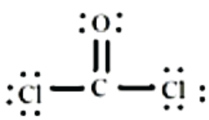
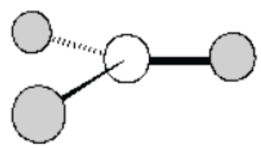
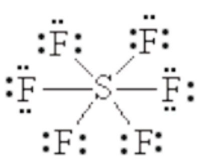
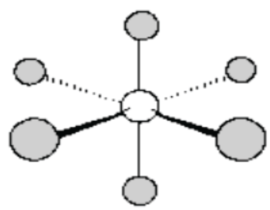
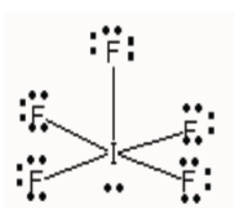
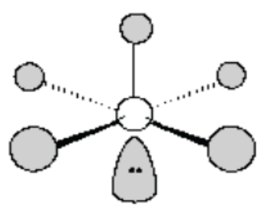
8. Because hydrogen bonds exist between ammonia molecules but only London dispersion forces exist between molecules of nitrogen and hydrogen, the ammonia has a higher boiling point and thus liquefies (condenses) at a higher temperature than nitrogen and hydrogen, which remain in the gaseous phase upon cooling.

9. As a result of hydrogen bonding, water remains liquid until 100°C. This is very important given the fact that most of our Earth and most of our bodies are composed of and require liquid water to survive. The three-dimensional structure of proteins and the base-pairing in the double helix of DNA molecules depend on hydrogen bonds. Hydrogen bond formation in ice makes it less dense than liquid water, which is crucial to aquatic life when bodies of water freeze.

10. I_2 is a relatively large and massive diatomic molecule and the strength of London dispersion forces increase as the size of the molecules involved increases. This is because large electron clouds are more loosely held than smaller clouds and thus more easily deformed or polarized by a nearby dipole than compact tightly held clouds. In addition, large molecules with more surface area have electron clouds that are spread out and so are more easily distorted by neighboring dipoles. As a result, the dispersion forces are strong enough to keep the molecules of I_2 attached to each other even at room temperature.

11. Although the ionic bonds holding a crystal lattice together are strong, when the surface of that lattice is in contact with water, each ion on that surface will attract the oppositely charged end of polar water molecules near them. That attraction between an ion and a polar molecule is called an **ion-dipole force**. These attractive forces soon overcome those existing between the ions themselves and so the crystal structure begins to break down and the ionic compound dissolves. As the ions move away from the lattice surface, they immediately become surrounded or enclosed in what chemists call a hydration shell. Ion-dipole forces are the primary force responsible for the solubility of ionic compounds in water.

12.

Lewis Structure	**AX_mE_n Notation**	**Shape of Molecule (Name and Diagram)**	**Type of Intermolecular Force Acting Between Molecules**
(a)  dichloromethane	**AX_4**	 tetrahedral	dipole-dipole
(b)  phosgene (a gas used in the trenches in WWI)	**AX_3**	 trigonal planar	dipole-dipole
(c)  sulphur hexafluoride	**AX_6**	 octahedral	London dispersion
(d)  iodine pentafluoride	**AX_5E**	 square pyramidal	dipole-dipole