

## Chemical Bonding

### Answers and Explanations

#### 1. B

A covalent bond involves the sharing of electron pairs between atoms leading to a stable balance of attractive and repulsive forces between atoms. In many cases, the sharing of electrons allows each atom to attain the equivalent of a full outer shell, which is a stable electronic configuration.

#### 2. C

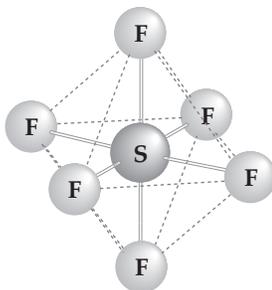
As with Mg (electronegativity 1.3) and Cl (electronegativity 3.1), a bond will be ionic (> 50% ionic character) when the electronegativity difference between bonded atoms is 1.7 or greater. Ionic bonding occurs when elements from the far left of the periodic table (alkali and alkaline earth metals) form compounds with the more electronegative nonmetals from the right side of the periodic table.

#### 3. D

$O^{2-}$ ,  $F^-$  and  $Na^+$  possess the same configuration as neon, not argon.

#### 4. C

The premise of valence shell electron pair repulsion theory (VSEPR) is that the valence electron pairs surrounding an atom tend to repel each other and adopt an arrangement that minimizes this repulsion. In  $SF_6$  there are six electron pairs in the covalent bonds between sulfur and fluorine. The geometry that results is octahedral.



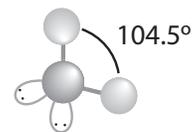
#### 5. D

Sulfur dioxide has two resonance structures which contribute equally to the overall hybrid structure of the molecule.



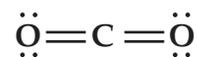
#### 6. A

The H–O–H bond angle in water equals  $104.5^\circ$ . With four regions of electron density, ie. the two electron pairs in the covalent bonds plus the two nonbonded pairs, the geometry is based on the tetrahedron, but the angle is not the tetrahedral angle of  $109.5^\circ$  because the nonbonded pair repulsions are stronger.



#### 7. D

With all of its electrons participating in the double bonds to the two oxygens, the carbon of  $CO_2$  has two regions of electron density. A double bond is considered a single region of electron density in VSEPR. This produces a linear geometry. In other words, the O–C–O bond angle is  $180^\circ$ .



#### 8. B

Molecules with three regions of electron density have a geometry that is trigonal planar, so one might expect the O–O–O bond angle to be  $120^\circ$ . However, lone pairs of electrons are slightly more repulsive than bonding pairs of electrons. Therefore the O–O–O angle is a few degrees less than  $120^\circ$ .

#### 9. A

The energy released when the bond is formed is the bond dissociation energy. The respective bond dissociation energies for the respective hydrogen-halogen bonds is as follows:

Bond dissociation energy – kJ mol<sup>-1</sup>

HF	570
HCl	432
HBr	366
HI	298

As a general rule, bonds with greater electronegativity difference will involve greater bond dissociation energy. As the bond forms there is not only the decrease in energy as the atoms fall into the well of binding energy associated with electrons in a molecular orbital between the nuclei pulling them together. There is also the additional decrease in energy associated with the electronegative element pulling the bonding electrons in towards its nucleus.

### 10. A

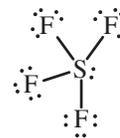
Bond length or bond distance is defined as the average distance between nuclei of two bonded atoms in a molecule. In the energy diagram, there is an energy minimum at 0.74 Å. It would energy to pull the two atoms apart from this distance due to the attractive force exerted on nuclei by the electrons within the sigma bonding orbital between them. It would also take energy to push the two atoms any nearer to each other due to the repulsive force of the two nuclei. Due to the thermal energy they possess being at some temperature, the two atoms will be vibrating much like a mass-spring oscillator with the bond distance analogous to the the equilibrium position of the oscillator.

### 11. B

Because it is a triple bond, the bond dissociation energy of N<sub>2</sub> (945 kJ/mol) is greater than the bond dissociation energy of H<sub>2</sub>. In forming the bond, the nitrogen atoms fall together into a deeper well of electrostatic potential energy.

### 12. C

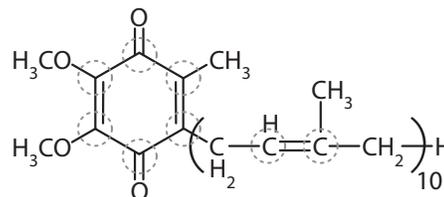
Sulfur has six valence electrons. Fluorine has seven. Constructing the Lewis dot structure of SF<sub>4</sub>, after completing the octets of the fluorine atoms in the molecule, sulfur has one pair of electrons remaining.



Along with phosphorus, having access to *d* subshell orbitals, sulfur is a prominent example of an element that can form an 'expanded octet.' Here, to accommodate five pairs of electrons, sulfur undergoes *sp*<sup>3</sup>*d* hybridization. This is the only choice available with five orbitals.

### 13. C

There are 26 *sp*<sup>2</sup> hybridized carbons depicted – 6 in the benzoquinone ring and 20 in the isoprenoid tail.



### 14. C

In F<sub>2</sub> there is one more pair of electrons in bonding orbitals than in antibonding orbitals, corresponding to a bond order of 1.

### 15. D

In the Ne<sub>2</sub> molecular orbital diagram there are equal numbers of electrons in bonding and antibonding orbitals. The bond order is zero. Any covalent bonding interaction would be transitory and unstable.

### 16. B

The O<sub>2</sub> molecular orbital diagram provides the explanation for why the O<sub>2</sub> molecule is paramagnetic. There are two singlet electrons in antibonding orbitals. With O<sub>2</sub> you have a triple bond with two additional singlet electrons in antibonding orbitals. The bond order is 2.