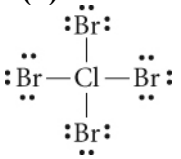
Chapter 4 Review, pages 262–267

Knowledge

1. (d)
2. (c)
3. (d)
4. (c)
5. (d)
6. (b)
7. (c)
8. (a)
9. (c)
10. (d)
11. (a)
12. (b)
13. True
14. True
15. False. A molecule that has two equal but oppositely pulling dipoles, such as carbon dioxide, will *not* orient in an electric field.
16. True
17. True
18. False. Linear molecules with three different types of atoms *may be* polar molecules, *depending on the electronegativity differences between the bonded atoms.*
19. False. A molecule of difluoroethyne, C_2F_2, contains *sigma and pi* bonds.
20. False. The smaller the molecular mass of a substance, the *smaller* the London dispersion forces within it.
21. True
22. True
23. False. The electrons in covalent network crystals *do not* move freely through the crystals.

Understanding

24. Answers may vary. Sample answer: Both ionic and covalent bonds are electrical attractions that hold atoms together. An ionic bond occurs between ions of opposite charges. This situation arises when the electronegativity difference between two atoms is large enough for one atom to completely lose an electron to another atom. This often occurs between a metal and a non-metal. A covalent bond involves sharing of electrons between atoms. This sharing may be unequal, creating a dipole. Covalent bonds typically form between non-metal atoms.
25. (a) Lewis structure of tetrabromomethane, CBr_4.



(b) In a molecule of tetrabromomethane, four fluorine atoms are bonded to a central carbon atom, so there are four C–F bonds. Carbon shares one of its four valence electrons with each bromine, and each bromine shares one of its seven valence electrons with carbon.

26. Hydrogen forms a covalent bond. It will share its single electron with another electron from the atom it bonds with. Covalent bonds typically form between atoms of non-metal elements. When the atoms are a certain optimum distance from each other, the attractive proton–electron force balances the repulsive proton–proton and electron–electron forces, and a covalent bond forms. The resulting molecule is more stable and has lower energy than the individual atoms that comprise it.

Carbon also forms a covalent bond. It will share its 4 valence electrons with electrons from other atoms it bonds with. If carbon were to gain 4 electrons, it would be very difficult for carbon's 6 protons to hold onto 10 electrons. For carbon to lose 4 electrons would require a large quantity of energy and would leave behind 6 protons with only 2 electrons. Therefore, carbon atom completes its octet only by sharing its valence electrons with other atoms.

Sodium forms an ionic bond. Sodium loses its one valence electron relatively easily. In the reaction of sodium with a non-metallic element, each sodium atom's valence electron is transferred to an atom of the non-metallic element. The resulting sodium cations attract the anions produced in the reaction to form ionic bonds. Ionic bonds usually form between a metal and a non-metal.

Helium does not form bonds because its valence shell is already full.

27. (a) Answers may vary. Sample answer: Two cases where an atom in a molecule has more or fewer than 8 valence electrons are $BeCl_2$ and PCl_5. The beryllium atom in beryllium chloride has only 4 valence electrons. The phosphorus atom in phosphorus pentachloride has 10 valence electrons.

(b) Other than hydrogen, atoms with fewer than 8 valence electrons are unstable. These atoms tend to react readily in order to complete their valence shell.

(c) If a molecule contains an atom (other than hydrogen) with fewer than 8 valence electrons, I would expect that atom to form a coordinate covalent bond, with both electrons involved in the bond coming from another atom.

28. (a) The three-dimensional bonding structure of ammonia is trigonal pyramidal. The structure of methane is tetrahedral. In both cases, the central atom has 4 electron pairs, but in methane they are all bonding pairs, whereas in ammonia, one pair is a lone pair.

(b) The angles in ammonia are 107º, and the angles in methane are 109.5º.

(c) The tetrahedral structure of methane with a 109.5º angle between bonds allows the maximum possible separation of the 4 pairs of bonding electrons around the central carbon atom. The bond angles in ammonia are 107º because the lone electron pair on the central nitrogen atom repels a little more than the 3 bonding pairs, and thus pushes the bonding pairs closer together. As a result, the bond angles decrease from 109.5° to 107°.

29. Boron tribromide, BBr_3, has a trigonal planar structure. Each of the three valence electrons in boron is shared with a valence electron from each of the bromines. Since there are three identical electron pairs around boron, a symmetrical arrangement (trigonal planar) is formed, with an angle of 120º between bonds.

30.

Molecule	Electronegativity difference		Type of bond
Br_2	$\Delta EN_{Br-Br} = 0$	$\Delta EN < 0.5$	non-polar
CaO	$\Delta EN_{O-Ca} = EN_O - EN_{Ca}$ $= 3.5 - 1.0$ $\Delta EN_{O-Ca} = 2.5$	$\Delta EN > 1.7$	ionic bond
CO	$\Delta EN_{O-C} = EN_O - EN_C$ $= 3.5 - 2.5$ $\Delta EN_{O-C} = 1.0$	$0.5 \le \Delta EN \le 1.7$	polar covalent

The greater the difference in electronegativity, the more ionic the character of the bond will be. The type of bonding and differences in electronegativity are as follows: non-polar, < 0.5; polar covalent, 0.5–1.7; and ionic, > 1.7.

31. (a)

Molecule	Electronegativity difference		Type of bond
HCl	$\Delta EN_{O-Ca} = EN_O - EN_{Ca}$ $= 3.5 - 1.0$ $\Delta EN_{O-Ca} = 2.5$	$0.5 \le \Delta EN \le 1.7$	polar covalent
Cl_2	$\Delta EN_{Cl-Cl} = 0$	$\Delta EN < 0.5$	non-polar

HCl has polar covalent bonding and Cl_2 has non-polar bonding.

(b) If these 2 molecules were placed in an electric field, there would be no effect on Cl_2, and HCl would align its dipoles with the electric field.

32. $\Delta EN_{C-S} = 0.0$, $\Delta EN_{S-Cl} = 0.5$, $\Delta EN_{N-S} = 0.5$, $\Delta EN_{S-P} = 0.4$, and $\Delta EN_{O-S} = 1.0$.

From greatest electronegativity difference to least:
S–O > N–S = S–Cl > P–S > C–S

33. (a) Of the molecules of water, carbon monoxide, carbon tetrachloride, boron trifluoride, and nitrogen tribromide, the boron trifluoride and carbon tetrachloride molecules would not have a net dipole.

(b) Water is a polar molecule because of its shape and the polarity of its two O–H bonds. Since the water molecule is bent, the bond dipoles add to give a net molecular dipole. Carbon monoxide is polar because of the electronegativity difference between carbon and oxygen. Nitrogen tribromide is polar because of its trigonal pyramidal structure and the polarity of its three N–H bonds. Since the bond dipoles in nitrogen tribromide are oriented in a trigonal pyramidal geometry, they add to produce a net molecular dipole. Boron trifluoride has identical bonds in trigonal planar geometry and carbon tetrachloride has identical bonds in a tetrahedral geometry. Because the structures are symmetrical and the bonds are identical, the individual bond dipoles cancel each other. As a result there is no net dipole in these molecules.

(c) In water and carbon monoxide, the dipole direction is toward the oxygen of the molecule. In nitrogen tribromide, the direction of the dipole is toward the bromines of the molecule.

34. Ammonia is a polar molecule because its lone pair of electrons gives it a trigonal pyramidal geometry, so that the bond dipoles add together to produce a net dipole. Methane is a non-polar molecule because of its symmetrical tetrahedral structure, in which opposing dipoles cancel each other out.

35. CCl_2H_2, CCl_4, CF_2H_2, and CF_2Cl_2 are all tetrahedral molecules. To compare the polarities of these molecules, look at the electronegativity differences between the atoms.

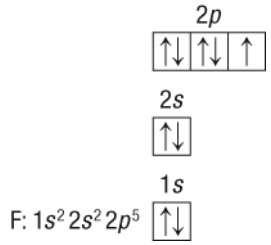
$$\Delta EN_{F-C} = EN_F - EN_C \qquad \Delta EN_{Cl-H} = EN_{Cl} - EN_H$$
$$= 4.0 - 2.5 \qquad = 3.0 - 2.1$$
$$\Delta EN_{F-C} = 1.5 \qquad \Delta EN_{Cl-H} = 1.1$$
$$\Delta EN_{F-Cl} = EN_F - EN_{Cl} \qquad \Delta EN_{Cl-C} = EN_{Cl} - EN_C$$
$$= 4.0 - 3.0 \qquad = 3.0 - 2.5$$
$$\Delta EN_{F-Cl} = 1.0 \qquad \Delta EN_{Cl-C} = 0.5$$

The differences in electronegativity between F and H are the greatest, then between Cl and H, then between Cl and F, and lastly between C and Cl. Thus, the arrangement of the molecules from most polar to least polar is $CF_2H_2 > CCl_2H_2 > CF_2Cl_2 > CCl_4$.

36. (a) Description of the bonding in a molecule of carbon tetrafluoride, CF_4, using valence bond theory:

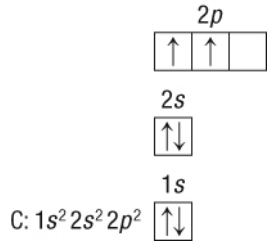
There are 4 C–F bonds in carbon tetrafluoride.

The electron configuration for fluorine is $1s^2 2s^2 2p^5$. The energy-level diagram is

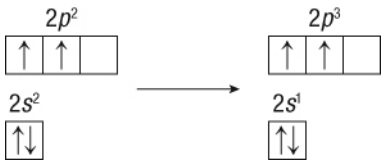


The valence electron diagram of the fluorine atom has 1 unpaired electron in the *p* orbital. Each of the 4 fluoride atoms contributes one 2*p* electron to the bonds in CF_4.

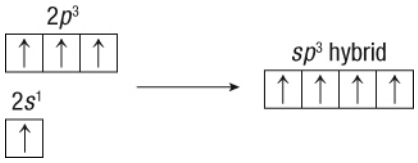
The electron configuration for carbon is $1s^2 2s^2 2p^2$. The energy-level diagram is



A covalent bond forms when 2 atomic orbitals, each with an unpaired electron, overlap; thus, of the 4 valence electrons for carbon, 1 of the 2 electrons in the 2*s* orbital must go into the empty 2*p* orbital.



The C–F bonds in carbon tetrafluoride are all of equal length and are arranged symmetrically in a tetrahedral geometry. The 3 *s* orbitals and 1 *p* orbital are combined to form 4 equivalent hybrid sp^3 orbitals, each with 1 unpaired electron.



The 2*p* orbitals from the fluorine atoms overlap with the sp^3 orbitals of the carbon atoms to form the covalent bonds in carbon tetrafluoride.

(b) The bonds formed in carbon tetrafluoride are sigma bonds, because *p* orbitals and hybrid orbitals overlap.

37. (a) Hybrid orbitals are used to describe bonding in molecules that have equivalent chemical bonds, yet the sharing of electrons occurs in multiple orbitals. Since different orbitals have different shapes, atoms in different orbitals would form different types of chemical bonds. Since all the bonds are equivalent, the orbitals must be hybridized. The hybrid orbitals are identical to each other and result in equivalent chemical bonds.

(b) Hybridization of carbon atoms in a molecule of dibromoethyne, C_2Br_2:

$$Br-C\equiv C-Br$$

There are 2 single Br–C bonds and 1 triple C≡C bond in carbon tetrafluoride. Each bromine atom and each carbon atom contributes one electron to the Br–C bonds. Each carbon atom undergoes partial hybridization to form two *sp* orbitals. The remaining two *p* orbitals of each carbon atom each hold a single electron.

(c) Types of bonds formed between all atoms in C_2Br_2:

Using their *sp* orbitals, the carbon atoms form 1 sigma bond with a bromine atom and 1 sigma bond with the other carbon atom. Each carbon atom is left with 2 additional unpaired electrons, each in a *p* orbital. These unpaired electrons form pi bonds with the other carbon atom. Thus, the triple bond consists of 1 sigma bond and 2 pi bonds.

38. Ethanol is a good solvent because of its small size and polarity.

39. (a) Of PCl_3, NaCl, SiC, and CH_3OH, $SiCl_4$ has physical properties primarily determined by London dispersion forces because it consists of non-polar molecules. PCl_3 and CH_3OH consist of polar molecules, so their properties are determined by dipole–dipole forces and hydrogen bonding, respectively. Silicon carbide, SiC, is a network solid in which each silicon atom forms four strong covalent bonds with neighbouring carbon atoms.

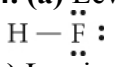
(b) Of the compounds listed in (a), SiCl has the strongest intermolecular forces of attraction, becasue of the strength of its covalent bond.

40. The shape of a meniscus of water in a glass cylinder is different from the shape of a meniscus of mercury in a glass cylinder because the physical properties of water and mercury—such as viscosity, density, and surface tension—are different as result of the different intermolecular forces in water and mercury. In water, the adhesive forces between water molecules and glass molecules are stronger than the adhesive forces among water molecules and a concave meniscus forms. In mercury, the cohesive forces are stronger than adhesive forces and a convex meniscus forms.

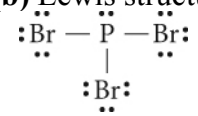
41. (a) KBr(s) is an ionic crystal, where ionic bonds result in the following properties: hard, brittle, does not conduct electricity, soluble, conducts electricity in solution, and has a high melting point.
(b) $H_2O(s)$ is a molecular crystal, where intermolecular forces result in the following properties: weak, brittle, does not conduct electricity, and has a low melting point.
(c) C(s) (diamond) is a covalent network crystal, where covalent bonds result in the following properties: hard, usually does not conduct electricity, insoluble, and has a very high melting point.
(d) Cu(s) is a metallic crystal, where metallic bonding results in the following properties: hard, malleable, conducts electricity, insoluble, has a high melting point.
42. Doping silicon with either phosphorus or gallium increases the electrical conductivity of pure silicon because by adding an atom with 5 valence electrons such as phosphorus, the atom still binds to silicon but has a free electron that undergoes movement and allows conductivity through the solid. An atom with 3 valence electrons, such as gallium, has 1 fewer valence electron than silicon, which creates a "hole" that electrons from surrounding atoms can move in to fill.
43. (a) Metals are malleable, shiny, and conduct electricity.
(b) Ionic crystals are hard and brittle, conduct electricity in solution, and have high melting points.
(c) Molecular crystals of polar compounds are less hard than ionic crystals, have a low melting point, and do not conduct an electric current well.
(d) Molecular crystals of non-polar compounds are less hard than ionic crystals, have a very low melting point, and do not conduct an electric current well.
(e) Network crystals are hard, have a very high melting point, and are poor conductors of electricity.

Analysis and Application

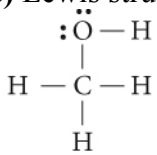
44. (a) Lewis structure for HF:



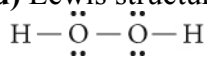
(b) Lewis structure for PBr_3:



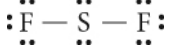
(c) Lewis structure for CH_3OH:



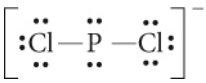
(d) Lewis structure for H_2O_2:



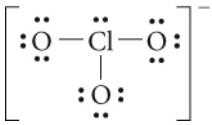
(e) Lewis structure for SF_2:



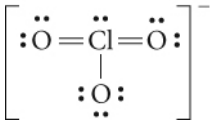
45. (a) Lewis structure for PCl_2^-:



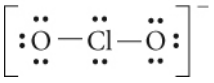
(b) Lewis structure for ClO_3^-:
Following the procedure outlined in this text, the structure of the chlorate ion is



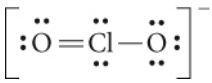
However, evidence suggests that chlorate ions contain two double bonds. This is thought to occur as a result of the overlap of two oxygen orbitals containing lone pairs with two empty 3d orbitals. The Lewis structure consistent with this evidence is



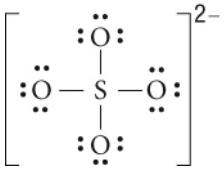
(c) Lewis structure for ClO_2^-:



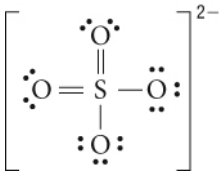
However, evidence suggests that chlorate ions contain one double bond. This is thought to occur as a result of the overlap of one oxygen orbital containing a lone pair with an empty 3d orbital. The Lewis structure consistent with this evidence is



46. (a) There is 1 lone pair of electrons on the central atom in SO_2.
(b) There is 1 lone pair of electrons on the central atom in NBr_3.
(c) There are no lone pairs of electrons on the central atom in BF_3.
(d) There are 2 lone pairs of electrons on the central atom in H_2O.
47. (a) Lewis structure of the sulfate ion, SO_4^{2-}:



However, experimental evidence shows that sulfur tends to form six bonds rather than four. Further evidence shows that two of the sulfur–oxygen bonds are actually double bonds. The Lewis structure that agrees with this evidence is



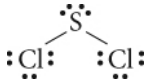
(b) In the sulfate ion, the bonds between sulfur and the two oxygen atoms are covalent, and the bonds between sulfur and the two oxygen anions are ionic.

48.

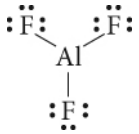
	Molecule	Number of electrons around central atom	Number of bonding pairs on central atom	Number of lone pairs on central atom
(a)	SO_2	8	3	1
(b)	NBr_3	8	3	1
(c)	BF_3	6	3	0
(d)	CO_2	8	4	0
(e)	SCl_6	12	6	0
(f)	$SiCl_4$	8	4	0
(g)	H_2	2	1	0
(h)	$AlBr_3$	6	3	0

49. Using the VSEPR theory to predict molecular shape**:**

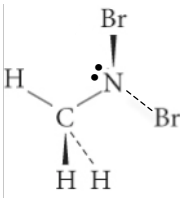
(a) In the SCl_2 molecule, the central S atom has 4 electron pairs, including 2 lone pairs, so SCl_2 has a bent structure.



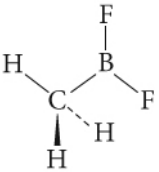
(b) In the AlF_3 molecule, the central Al atom has 3 electron groups and no lone pairs, so AlF_3 has a trigonal planar structure.



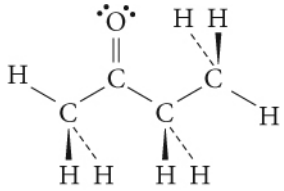
(c) In the CH_3NBr_2 molecule, the C atom has 4 electron pairs and no lone pairs, so the geometry around the C atom is tetrahedral; the N atom has 4 electron pairs, including 1 lone pair, so the geometry around the N atom is trigonal pyramidal.



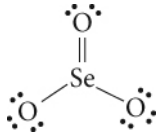
(d) In the CH_3BF_2 molecule, the C atom has 4 electron pairs and no lone pairs, so the geometry around the C atom is tetrahedral; the B atom has 3 electron pairs and no lone pairs, so the geometry around the B atom is trigonal planar.



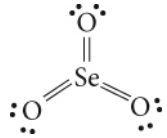
(e) In the $CH_3COCH_2CH_3$ molecule, the C 1, 3, and 4 each have 4 electron pairs and no lone pairs, so the geometry around these C atoms is tetrahedral; C 2 has 3 electron groups and no lone pairs, so the geometry around C 2 is trigonal planar.



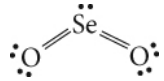
50. Three-dimensional structure if molecule, including bond angles:
(a) In SeO_3, the central Se atom has 3 electron groups and no lone pairs, so SeO_3 has a trigonal planar structure. The bond angles are 120º.



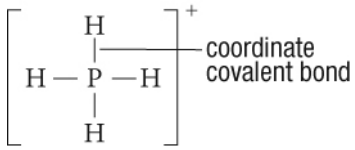
The three dimensional structure of this molecule is trigonal planar with 120º bond angles. However, experimental evidence shows that selenium forms three double bonds in this molecule. The Lewis structure that agrees with this evidence is:



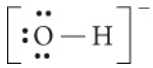
(b) In SeO_2, the central Se atom has 3 electron groups, including 1 lone pair, so SeO_2 has a bent, or V-shaped, structure. The bond angle is 104º.



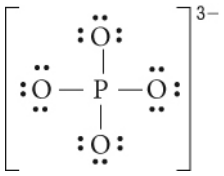
51. (a) CS_2 is linear.
(b) CH_3SH is tetrahedral around C, and bent around S.
(c) CH_3CN is tetrahedral around C bonded to H, linear around C bonded to N.
(d) C_2F_4 is trigonal planar.
52. (a) Lewis structure for PH_4^+:



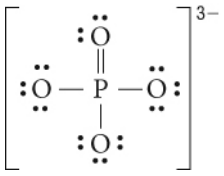
(b) Lewis structure for OH^-:



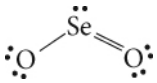
(c) Lewis structure for PO_4^{3-}:
Using the procedure in this text, the structure is



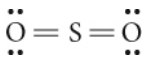
However, the actual structure is:



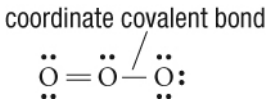
(d) Lewis structure for SO_2:



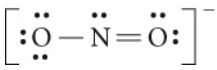
However, the actual structure is



(e) Lewis structure for O_3:



(f) Lewis structure for NO_2^-:



53. (a) $\Delta EN_{Li-H} = 1.1$, $\Delta EN_{Mg-H} = 0.9$, $\Delta EN_{B-H} = 0.1$, $\Delta EN_{N-H} = 0.9$, $\Delta EN_{O-H} = 1.4$
In order of increasing polarity: B–H < N–H = Mg–H < Li–H < O–H
(b) $\Delta EN_{Al-F} = 2.5$, $\Delta EN_{Ga-Br} = 1.2$, $\Delta EN_{Ge-Cl} = 1.2$, $\Delta EN_{Sn-F} = 2.2$
In order of increasing polarity: Ga–Br = Ge–Cl < Sn–F < Al–F
(c) $\Delta EN_{P-S} = 0.4$, $\Delta EN_{P-Cl} = 0.9$, $\Delta EN_{P-O} = 1.4$, $\Delta EN_{P-F} = 1.9$, $\Delta EN_{P-H} = 0.0$
In order of increasing polarity: P–H < P–S < P–Cl < P–O < P–F
54. (a) $\Delta EN_{P-F} = 1.9$, so P–F is polar covalent.
(b) $\Delta EN_{Li-H} = 1.1$, so Li–H is polar covalent.
(c) $\Delta EN_{Mg-Cl} = 1.8$, so Mg–Cl is ionic.
(d) $\Delta EN_{P-H} = 0.0$, so P–H is covalent.
(e) $\Delta EN_{Si-Br} = 1.0$, so Si–Br is polar covalent.
(f) $\Delta EN_{Ge-H} = 0.3$, so Ge–H is covalent.

55. (a) $EN_{Cl} = 3.0,\ EN_{Si} = 1.8$

Partial charges on atoms forming Cl–Si bond: Si δ^+; Cl δ^-

(b) $EN_{Li} = 1.0,\ EN_{H} = 2.1$

Partial charges on atoms forming Li–H bond: Li δ^+; H δ^-

(c) $EN_{Al} = 1.5,\ EN_{I} = 2.5$

Partial charges on atoms forming Al–I bond: Al δ^+; I δ^-

(d) $EN_{Ge} = 1.8,\ EN_{H} = 2.1$

Partial charges on atoms forming Ge–H bond: Ge δ^+; H δ^-

56. Given: $EN_B = 2.8$

Required: all atoms that can form a polar covalent bond with bromine such that the bromine atom has a partial positive charge

Solution: Bromine's electronegativity is 2.8.

If an atom X forms a polar covalent bond with bromine, $0.5 \leq |\Delta EN_{B-X}| \leq 1.7$.

If bromine has a partial positive charge, $EN_B < EN_X$.

If $EN_B < EN_X,\ EN_X - EN_B \geq 0.5$

$$EN_X \geq 0.5 + EN_B$$

$$EN_X \geq 0.5 + 2.8$$

$$EN_X \geq 3.3$$

Atoms meeting this criterion are O and F.

Statement: The atoms that can form a polar covalent bond with bromine such that the bromine atom has a partial positive charge are O and F.

57. (a) CS_2 would not orient in an electric field because it is non-polar. It is non-polar because it is linear with two identical C–S bonds, so the molecule is symmetrical and bond dipoles will cancel out.

(b) CH_3OH would orient in an electric field because it is polar. It is polar because the geometry of C–O–H is bent and thus asymmetrical, so there is a net dipole.

(c) C_5H_{12} would not orient in an electric field because it is non-polar. The C–H bonds are non-polar ($\Delta EN_{C-H} = 0.4$), as are the C–C bonds, so there is no net dipole.

(d) SCl_2 would orient in an electric field because it is polar. It is polar because the S atom has 4 electron pairs, 2 of which are lone pairs, so its structure is bent (asymmetrical), so there is a net dipole.

(e) OF_2 would orient in an electric field because it is polar. It is polar because the O atom has 4 electron pairs, 2 of which are lone pairs, so its structure is bent (asymmetrical), so there is a net dipole.

58. (a) There is no dipole in the B–B bond in B_2H_4, so no partial charges on the B atoms.

(b) $EN_S = 2.5,\ EN_C = 2.5$

There is no dipole in the S–C bond in CH_3SH, so no partial charges on the atoms.

(c) $EN_H = 2.1,\ EN_C = 2.5$

The partial charges on the atoms forming the H–C bond in C_5H_{12} are: H δ^+; C δ^-.

(d) $EN_N = 3.0,\ EN_C = 2.5$

The partial charges on the atoms forming the C–N bond in HOCN are: C δ^+; N δ^-.

59. (a) $\Delta EN_{Br-Si} = EN_{Br} - EN_{Si}$

$= 2.8 - 1.8$

$\Delta EN_{Br-Si} = 1.0$

$SiBr_4$ has polar Si–Br bonds ($0.5 < \Delta EN_{Br-Si} < 1.7$) but no net dipole because the molecule has a symmetrical shape (tetrahedral with 4 identical bonds), so the bond dipoles cancel one another.

(b) $\Delta EN_{Br-Si} = EN_{Br} - EN_{Si}$ $\quad \Delta EN_{H-Si} = EN_{H} - EN_{Si}$

$= 2.8 - 1.8$ $\quad = 2.1 - 1.8$

$\Delta EN_{Br-Si} = 1.0$ $\quad \Delta EN_{H-Si} = 0.3$

$SiHBr_3$ has polar Si–Br bonds and has a net dipole because it is asymmetrical. It has a symmetrical tetrahedral shape but the bonds are not identical: there are 3 Si–Br bonds and 1 Si–H bond.

(c) $\Delta EN_{F-B} = EN_{F} - EN_{B}$ $\quad \Delta EN_{H-B} = EN_{H} - EN_{B}$

$= 4.0 - 2.0$ $\quad = 2.1 - 2.0$

$\Delta EN_{F-B} = 2.0$ $\quad \Delta EN_{H-B} = 0.1$

BH_2F has a polar bond and has a net dipole because it is asymmetrical. It has a symmetrical trigonal planar shape but the bonds are not identical: there are 2 B–H bonds and 1 B–F bond.

(d) $\Delta EN_{H-P} = EN_{H} - EN_{P}$

$= 2.1 - 2.1$

$\Delta EN_{H-P} = 0.0$

PH_3 has non-polar bonds.

60. (a) SeO_2 is polar. Since $EN_{Se} = 2.4$ and $EN_{O} = 3.5$, it has polar bonds. There are 3 electron groups around the central atom (a single bond, a double bond, and a lone pair), so the molecular structure is bent, which is an asymmetrical structure.

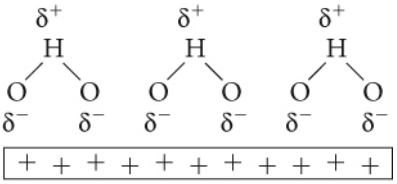
(b) SBr_2 is polar. Since $EN_{S} = 2.5$ and $EN_{Br} = 2.8$, it does not have polar bonds. But of the 4 pairs of electrons around the central atom, 2 are lone pairs, so the molecular structure is bent, which is an asymmetrical structure.

(c) SiO_2 is non-polar. $EN_{S} = 1.8$ and $EN_{O} = 3.5$, it has polar bonds. However, the molecule is linear, with two identical Si=O double bonds, so it is symmetrical, and the bond dipoles cancel each other.

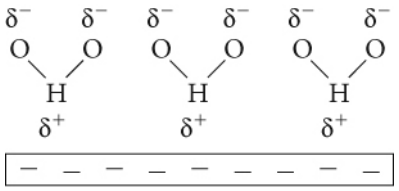
(d) SeO_3 is non-polar. Since $EN_{Se} = 2.4$ and $EN_{O} = 3.5$, it has polar bonds. However, the molecule is trigonal planar, with three identical Se=O double bonds, so it is symmetrical, and the bond dipoles cancel each other.

(e) N_2O is polar. Since $EN_{N} = 3.0$ and $EN_{O} = 3.5$, it has a polar bond. The structure is linear but the bonds are not identical, so the molecule is not symmetrical.

61. (a) The reason why water is attracted to both a negatively charged ebonite rod and a positively charged glass rod is that water has a positive and negative dipole, each of which are attracted to the oppositely charged rod.
(b) Water attracted to a positively charged rod:

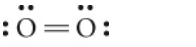


Water attracted to a negatively charged rod:

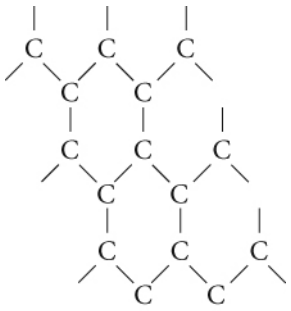


62. (a) The geometry around the C in CH_3OH is tetrahedral. C's 4 valence electrons are in orbitals 2*s* and 2*p,* so the hybridization of C in CH_3OH is sp^3.
(b) NCl_3 is trigonal pyramidal, with 3 N–Cl bonds. N has 3 unpaired electrons in its 2*p* orbital, and each of the 3 Cl atoms has 1 unpaired electron in its 3*p* orbital. The 3*p* orbital from the chlorine atom can overlap with the 2*p* orbital in the nitrogen atom to form a covalent bond, and no hybridization is needed.
(c) BBr_3 is trigonal planar, and B's 3 valence electrons are in its 2*s* and 2*p* orbitals, so the hybridization of B in BBr_3 is sp^2.
(d) The hybridization of C in HCOH is sp^2. HCOH has 2 H–C bonds and 1 C=O bond, and the O has 2 lone pairs of electrons. The C atom's 4 valence electrons are in the 2*s* and 2*p* orbitals. The C atom undergoes partial hybridization to form three sp^2 orbitals. The remaining *p* orbital is left with 1 electron in it, which will form a pi bond with O.
63. (a) SiF_4 is tetrahedral, with 4 Si–F bonds. The Si atom's 4 valence electrons are in orbitals 3*s* and 3*p*, so the hybridization of the central atom in SiF_4 is sp^3.
(b) The hybridization of the central B atom in $BeBr_2$ is *sp*. $BeBr_2$ is a linear molecule; the Be atom forms a bond with each Br atom. An atom with two groups of electrons is *sp* hybridized.
(c) $AlCl_3$ has a trigonal planar structure; the 3 valence electrons of the central Al atom are in the 3*s* and 3*p* orbitals, so the hybridization of the central atom in $AlCl_3$ is sp^2.
(d) The hybridization of each carbon atom in C_2Cl_4 is sp^2. The geometry of C_2Cl_4 is trigonal planar around each C. Each C forms 2 C–Cl bonds and one C=C bond. The C atom's 4 valence electrons are in the 2*s* and 2*p* orbitals. The C atom undergoes partial hybridization to form three sp^2 orbitals. The remaining *p* orbital is left with 1 electron in it, which will form a pi bond with the other carbon atom.
(e) The hybridization of the central C atom in CO is *sp*. CO is a linear molecule; each atom has a lone pair and the two atoms are joined by a triple bond. An atom with two groups of electrons is *sp* hybridized.

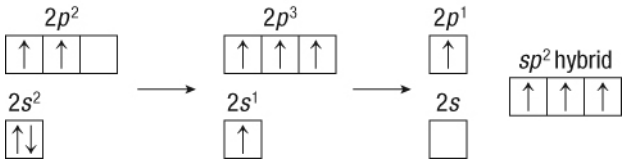
64. (a) There are 4 single Si–F bonds in SiF_4, so sigma bonds are formed in SiF_4.
(b) There are 2 single Be–Br bonds in $BeBr_2$, so sigma bonds are formed in $BeBr_2$.
(c) There are 3 single Al–Cl bonds in $AlCl_3$, so sigma bonds are formed in $AlCl_3$.
(d) In C_2Cl_4, each C forms 2 C–Cl bonds and one C=C bond. The C–Cl bonds are sigma bonds; there is 1 sigma bond and 1 pi bond between the two C atoms.
(e) The bond between C and O in the CO molecule is a triple bond, so 1 sigma bond and 2 pi bonds are formed in CO.
65. (a) Lewis structure of an oxygen molecule, O_2:



Each oxygen atom has 3 electron groups (1 double bond and 2 lone pairs), so the hybridization of each oxygen atom will be sp^2.
(b) There is 1 sigma bond and 1 pi bond in the oxygen molecule.
66. (a) Diagram of carbon atoms in a layer of graphite, including at least 12 C atoms:



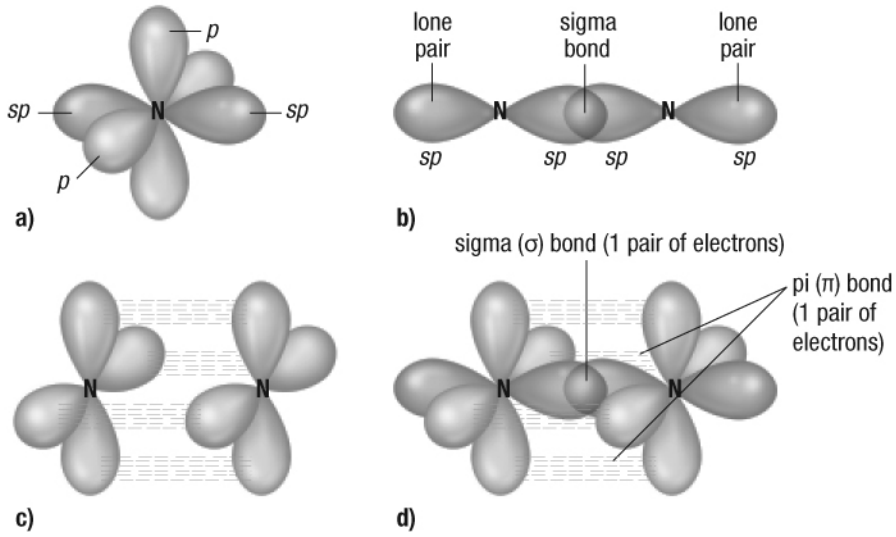
(b) Since each sp^2 orbital of the carbon atoms in graphite contains 1 electron, and each carbon has 4 valence electrons, the sp^2 hybridization is only partial, and the fourth electron will be in the *p* orbital. Students may choose to include a diagram:



(c) London dispersion forces are what cause each layer in graphite to be attracted to the layer above it and the layer below it.
67. Triple bonds are shorter and stronger than double or single bonds because there are additional electrons between the atoms; additional electrons attract the atoms together and strengthen the bond between the atoms. For example:

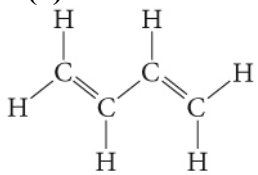
Double and triple bonds are more reactive than single bonds due to their electron richness compared to that of single bonds.

68. Illustration of the bonding in the nitrogen molecule, N_2:



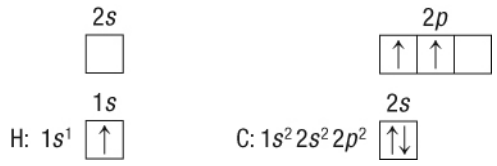
An *sp* hybridized nitrogen atom. There are two *sp* hybrid orbitals and two unhybridized *p* orbitals. (b) The sigma bond in the N_2 molecule. (c) The two pi bonds in N_2 are formed when electron pairs are shared between two sets of parallel *p* orbitals. (d) The total bonding picture for N_2.

69. (a) Lewis structure of 1,3-butadiene, $CH_2CHCHCH_2$:

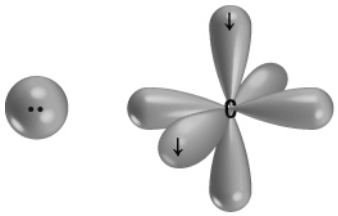


(b) Each carbon atom in 1,3-butadiene has 3 electron groups (1 double and 2 single bonds), so the hybrid orbitals used by each carbon atom are sp^2 orbitals.

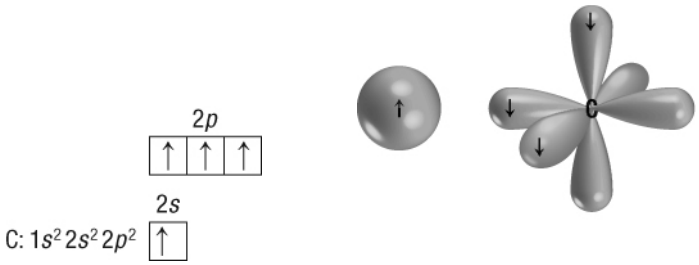
(c) & (d) Orbitals of the 1,3-butadiene molecule before bonding:



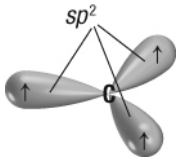
In each C, there are 2 unpaired electrons in 2*p*.



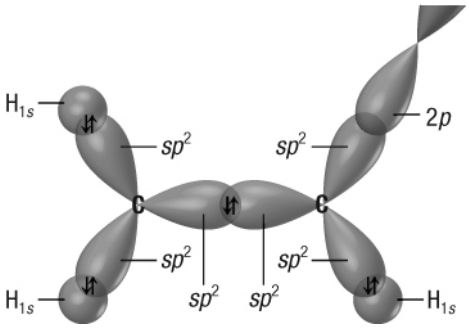
In each H, there is 1 unpaired electron in 1*s*.
The C atoms each have 3 electron groups, so carbon requires 3 hybrid orbitals. Therefore, the hybridization must be sp^2. An electron is promoted from carbon's 2*s* to its 2*p* orbital just prior to bonding:



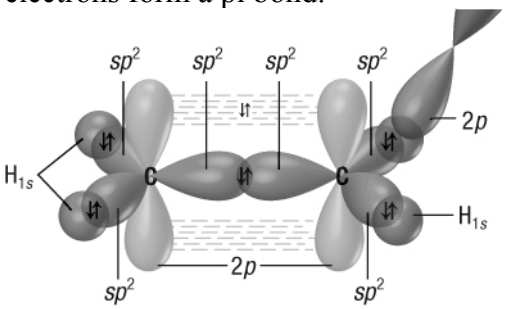
The carbon atoms are partially hybridized. The *s* orbital and 2 of the *p* orbitals become 3 sp^2 hybrids.



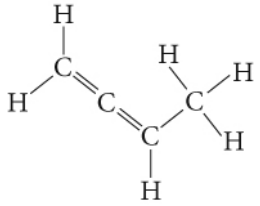
The carbon atoms can now each form 3 sigma bonds using the sp^2 orbitals: either 2 H–C and 1 C–C bond, or 1 H–C and 2 C–C bonds. For simplicity, only 1 of the p orbitals in the carbon at the top right of the diagram is shown.



Each carbon atom has 1 additional unpaired electron in a *p* orbital. These unpaired electrons form a pi bond.



(e) The same orbitals would not be used in 1,2-butadiene, CH_2CCHCH_3, as in 1,3-butadiene, $CH_2CHCHCH_2$. Consider the Lewis structure of 1,2-butadiene, CH_2CCHCH_3:



In 1,2-butadiene, the first and third carbon have 3 electron groups, so they use sp^2 orbitals; the second carbon has 2 electron groups, so it uses an *sp* orbital; and the fourth carbon has 4 electron groups, so it uses sp^3 hybridized orbitals. Therefore, different orbitals would be used.

70. Ethanol, CH_3CH_2OH, has a higher boiling point than methoxymethane, CH_3OCH_3, because the shape of the ethanol molecule gives it increased polarity. The stronger dipole–dipole forces make the molecules of ethanol more attracted to one another, giving ethanol a higher boiling point.

71. (a) *n*-pentane, $CH_3CH_2CH_2CH_2CH_3$, has a higher boiling point than neopentane, $CH_3CH_3CCH_3CH_3$, because its long chain arrangements allows its molecules to pack closer together than neopentane's molecules can. This produces more London dispersion forces between the molecules, which have to be broken in order for *n*-pentane to boil. Hence, the boiling point of *n*-pentane is higher than the boiling point of neopentane.

(b) HF has a higher boiling point than HCl because it has a stronger dipole.

(c) LiCl has a higher boiling point than HCl because it has a stronger dipole.

(d) Hexane, $CH_3CH_2CH_2CH_2CH_2CH_3$, has a higher boiling point than pentane, $CH_3CH_2CH_2CH_2CH_3$, because it has a greater molar mass which results in more London dispersion forces between its molecules.

72. (a) Ethanol, C_2H_5OH, will have greater intermolecular forces than propane, C_3H_8, because it has a hydrogen bond.

(b) SiH_4 will have greater intermolecular forces than CH_4 because it is a larger molecule with larger London dispersion forces.

(c) H_2S will have greater intermolecular forces than H_2O because it is a larger molecule with larger London dispersion forces.

(d) C_2H_5OH will have greater intermolecular forces than CH_3COH because it is a larger molecule with larger London dispersion forces.

(e) C_4H_9OH will have greater intermolecular forces than C_2H_5OH because it is a larger molecule with larger London dispersion forces.

73. (a) Acetone has a lower boiling point than carbon tetrachloride because carbon tetrachloride is a larger molecule. As a result, the London dispersion forces in carbon tetrachloride are stronger than the dipole–dipole attractions and London dispersion forces in acetone.

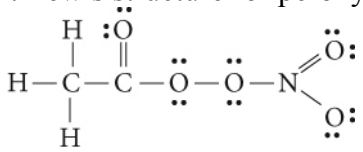
(b) Naphthalene has a higher ΔH_{vap} than benzene because naphthalene is a larger molecule with larger London dispersion forces than benzene has.

(c) Benzoic acid has a higher boiling point than benzene because benzoic acid is a polar molecule and has dipole–dipole forces, including hydrogen bonding, that benzene does not have.

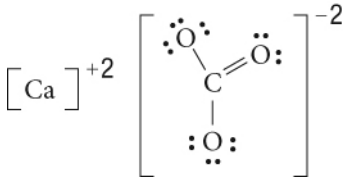
(d) Naphthalene has a lower ΔH_{vap} than benzoic acid because benzoic acid has dipole–dipole forces, which are stronger than the London dispersion forces in naphthalene.
74. (a) $CaSO_4$ forms an ionic solid.
(b) NaOH forms an ionic solid.
(c) U forms a metallic solid.
(d) PH_3 forms a molecular solid.
75. (a) I would expect K_2S to form ionic bonds.
(b) I would expect SiH_4 to form covalent bonds within its molecules.
(c) I would expect Zn to form metallic bonds.
(d) I would expect SiO_2 to form covalent bonds within its covalent network.
76. (a) A molecular element that is a gas at room temperature would form molecular bonds and would have a low melting point.
(b) A hard, silver-grey solid that conducts electricity as a solid would form metallic bonds and would be insoluble.
(c) A very hard, insoluble crystal that does not conduct electricity as a solid or a liquid would form covalent network bonds and would have a very high melting point.
(d) A hard, brittle solid that dissolves in water would form ionic bonds and would conduct electricity in solution.
77. Solid aluminum is a metallic crystal. It is expected to be hard, malleable, conduct electricity, and have a high melting point. Solid aluminum oxide is an ionic crystal. It is expected to be hard, brittle, not conduct electricity, and have a high melting point.
78. (a) SiC has a higher melting point than Pb. It is a covalent network solid and has stronger bonding than a metallic solid.
(b) Na_3PO_4 has a higher melting point than I_2. It is an ionic solid and has stronger bonding than a molecular solid.
(c) H_2O has a higher melting point than H_2Se. It has the stronger dipole–dipole forces, including hydrogen bonding, of the two molecular solids.

Evaluation

79. Lewis structure for peroxyacetylnitrate (PAN):



80. (a) Lewis structure for calcium carbonate, $CaCO_3$:

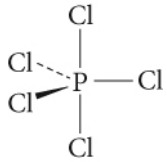


(b) The carbonate ion is trigonal planar with 120º bond angles because it has 3 groups of electrons on the central carbon atom.
(c) A trigonal planar structure uses sigma bonds. A double bond always consists of 1 sigma bond and 1 pi bond. The sp^2 orbital is used by the carbon atom to form the sigma bonds, and the *p* orbital is used to form the pi bond.

81. Answers may vary. Sample answer: I disagree with the statement, "A bond is either purely covalent, polar, or purely ionic. It is all very black and white." The polarity of intramolecular bonds depends on the electronegativity difference between the bonded atoms, and this difference affects whether a bond will be non-polar covalent, polar covalent, or ionic. However, the possible differences between electronegativities between various elements form a spectrum. Therefore, all bonds will have mixtures of covalent and ionic character.

82. (a) The orientation that would allow the greatest separation between bonding electrons in sulfur hexafluoride, SF_6, is four S–F bonds in the same plane separated by 90° angles, and the other two S–F bonds perpendicular to this plane, one above it and one below it. This gives the molecule an octahedral shape.

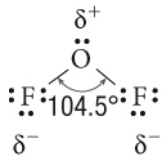
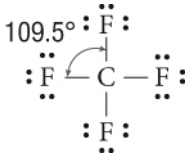
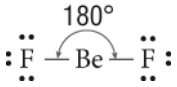
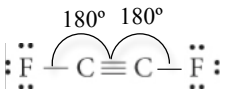
(b) Three-dimensional structure of phosphorus pentachloride, PCl_5:



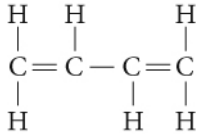
83. Of carbon dioxide, CO_2; carbonate ion, CO_3^{2-}; methanol, CH_3OH; and methyl mercaptan, CH_3SH, methanol would be the most polar, due to its C–O bond. Carbon dioxide and carbonate also have C–O bonds, but their structures are linear and trigonal planar, respectively. These symmetrical shapes cause opposing dipoles to cancel out. Methyl mercaptan will be non-polar because the C–S bond is non-polar.

84. Water is a better choice than hexane, $C_6H_{14}(l)$, for use in a capacitor because, unlike hexane, it is a polar molecule. This allows it to polarize (orient) in an electric field, allowing more electrical energy to be stored.

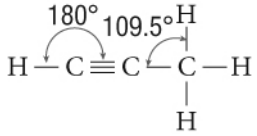
85. (a)

	Molecule	Lewis structure	Geometry	Bond angle	Polarity	Hybridization
(a)	OF_2		bent or V-shaped	104.5°	polar	sp^3
(b)	CF_4		tetrahedral	109.5°	non-polar	sp^3
(c)	BeF_2		linear	180°	non-polar	sp
(d)	C_2F_2		linear	180°	non-polar	sp

86. (a) Lewis structure for $H_2C=CHCH=CH_2$:



(b) The structure of $H_2C=CHCH=CH_2$ is trigonal planar around the C atoms, with bond angles of 120º. Each carbon is sp^2 hybridized.
(c) The molecule is non-polar.
87. The PO_4^{3-} ion is tetrahedral; the bonds are sp^3 hybridized with bond angles of 109.5º.
88. Proposed structure for propyne, $C_3H_4(g)$:



The first two carbons are *sp* hybridized with bond angles of 180º. There is one sigma bond and two pi bonds between the first two carbons. The third carbon is sp^3 hybridized with bond angles of 109.5º. This carbon has 4 sigma bonds.
89. (a) An I atom has 7 valence electrons, and an F atom has 7 valence electrons, so IF_5 will have 5 I–F bonds and 1 lone pair, for a total of 12 electrons around I. Thus, IF_5 does not follow the octet rule.
(b) The valence shell of iodine has 12 electrons.
90. Hydrogen peroxide is more viscous and has a higher boiling point than water because the additional oxygen in H_2O_2 strengthens the dipole–dipole interactions, as well as the amount of hydrogen bonding between molecules, in comparison to water. These increase the boiling point and the viscosity.
91. Each compound in Table 2 contains a different atom from the same period bonded to chlorine. Since electronegativity increases going from left to right across a period, the leftmost elements have a greater electronegativity difference with Cl. This results in greater bond polarities, which lead to stronger intermolecular attractions and higher melting points.

Reflect on Your Learning

92. Answers may vary. Answers should include a description of ionic bonding and covalent bonding and the significance of these types of bonding for the properties of the molecule. Research to answer unanswered questions could be done using peer-reviewed chemistry literature and reputable Internet sources.
93. Answers may vary. Sample answer: Learning about types of bonds before learning about VSEPR theory helped me understand which atoms will form covalent bonds, how electrons from each atom form covalent bonds, and how there are non-bonding pairs of electrons around atoms. An understanding of these concepts was helpful to me as I learned about VSEPR theory.

94. Answers may vary. Answers should include information about the structure as well as the physical and chemical properties of four types of solids: ionic, molecular, metallic, and covalent network solids.

- Covalent network solids are very hard, have the highest melting points of the four, can act as semiconductors, and are held together by covalent bonding.
- Metallic solids are hard and malleable, have very high melting points, conduct electricity in their solid and liquid state, and are held together in a "sea" of electrons and positively charged nuclei.
- Ionic solids are hard and brittle, have high melting points, conduct electricity in solution and are held together with ionic bonding.
- Molecular solids are weak and brittle, have low melting points, and are held together by dipole–dipole forces, London dispersion forces, and hydrogen bonding.

95. Answers may vary. Sample explanations of concepts to a fellow student who has not taken chemistry:

Molecular polarity Think of a molecule as a seesaw with electrons on it. The heavier side of the seesaw has a negative charge and the lighter side has a positive charge. Some atoms pull electrons closer to their side and offset the balance of charge on the seesaw, making the molecule polar.

Intermolecular forces

Consider two pieces of metal.

London dispersion forces: Imagine rubbing the two pieces of metal together. The frictional force between the two pieces of metal has some affect on the motion of the pieces of metal, but only a relatively small effect. The force of friction between the pieces of metal is analogous to London dispersion forces between molecules.

Dipole–dipole forces: When you magnetize two pieces of metal, the pieces of metal will be attracted to each other. Similarly, polar molecules are attracted to each other.

Hydrogen bonding: If you increased the magnetization on one part of each of the two pieces of metal, you would create a stronger attraction of a particular part of the metal pieces to each other, strengthening the overall attractive force between them. This is analogous to hydrogen bonds, which are strong dipole–dipole forces.

Structure of solids

Imagine building a bridge of the same design out of various substances.

Molecular solid: This is like a bridge built out of folded paper. The bridge will be a stable structure, but it will be very weak.

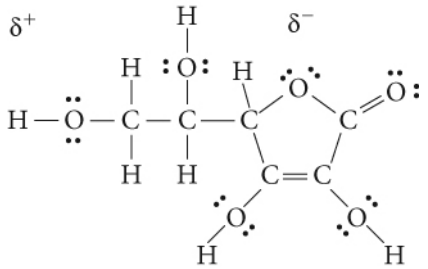
Ionic solid: This is like a bridge built out of toothpicks and glue. The bridge will be stronger than the paper bridge, but it will be brittle.

Metallic solid: This is like a bridge built out of plastic and putty. The bridge will be stronger than the ionic one, and it will be malleable.

Covalent network: This is analogous to building the bridge out of metal and contact cement. This bridge will be the strongest and will only break if a lot of force is applied.

Research

96. Lewis structure of vitamin C, including polarity:



97. DNA has a backbone composed of alternating, covalently linked deoxyribose sugars and phosphate groups. The nitrogenous bases adenine, guanine, cytosine, or thymine are attached to each sugar. Two strands of DNA are held together by the pairing of complementary bases. Bases pair using hydrogen bonding; there are 2 hydrogen bonds between adenine and thymine, and 3 hydrogen bonds between cytosine and guanine. Adjacent bases also undergo stacking interactions due to London dispersion forces between them. The stacking interactions of base pairs along with the hydrogen bonds between base pairs allow DNA to form a stable double helix.

98. Answers may vary. Sample answer: Many stain removers utilize the polarity of molecules to do their work. Water's polarity makes it an effective stain remover for other polar molecules and ionic compounds. Detergents are polar molecules used to emulsify compounds that are not usually soluble in water. Non-polar molecules tend not to dissolve in polar solvents such as water, but different non-polar molecules tend to mix well with one another, since the forces holding non-polar molecules together are relatively weak. For example, oil and grease molecules contain long carbon–carbon chains with hydrogen atoms bonded to the carbons. A C–H bond is non-polar. When detergent is added to water, the oil, water, and detergent form a mixture that does not separate. The detergent molecule has a non-polar hydrocarbon end and a polar end. Non-polar molecules in dirt, oils, and grease adhere to the non-polar end of a detergent molecule while the polar end increases the solubility of the detergent-dirt particle complex and allows for the removal of dirt, oil, and grease.

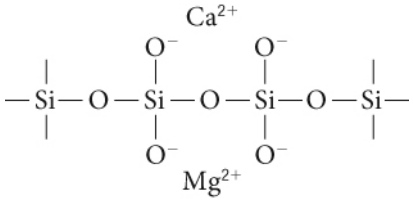
99. Answers may vary. Sample answer:

Bonding in carbon nanotubes: The bonding in carbon nanotubes is sp^2 covalent bonding, as in graphite; in contrast, diamond is sp^3 hybridized. The sheet of bonded carbons rolls into a "tube" using varying arrangements of the sheet.

Chemical and physical properties: Carbon nanotubes are very light and yet extremely stiff and strong, due to the strength of the sp^2 carbon–carbon bonds. Carbon nanotubes have the highest strength-to weight ratio of any known material. They are 5 times as stiff as steel and their tensile strength is around 50 times that of steel. Their thermal and electrical conductivity are also very high. Carbon nanotubes can be metallic or semiconducting depending on their structure, so some have electrical conductivities higher than that of copper, while others behave more like silicon. Because of their long, narrow shape, carbon nanotubes can easily penetrate membranes such as cell walls.

Another property of carbon nanotubes is that their electrical resistance changes significantly when other molecules attach themselves to the carbon atoms. The properties of nanotubes change depending on the diameter, length, and twist of the nanotube.
Applications of carbon nanotubes: Some applications of nanotubes are high-strength materials, circuit transistors, circuit cooling, molecular hydrogen storage, and solar cells. Nanotubes are used in flat-panel displays, scanning probe microscopes, and various sensing devices. Medical researchers are attaching molecules that are attracted to cancer cells to nanotubes to deliver drugs directly to diseased cells and using nanotubes as scaffolds for bone formation. Due to the unique properties of carbon nanotubes, it is expected that many more applications will be developed.

100. (a) Asbestos is a silicate, derived from the silicate ion, SiO_4^{4-}. The Si atom is covalently bonded to 4 oxygen atoms, each of which has a negative charge. Thus, each tetrahedral unit has a charge of –4. When linked together, the extended units are also negatively charged. The presence of other metallic ions, such as Ca^{2+}or Mg^{2+}, is necessary for electrical neutrality. The covalent Si–O bond is particularly strong and the linkage –Si–O–Si–O– is very stable. As a result of its SiO_4 covalent network combined with ionic interactions with metals, asbestos has a fibrous structure and forms stringy one-dimensional chains. The term asbestos actually refers to two kinds of minerals: serpentines and amphiboles. Serpentines have a sheet structure while amphiboles have a double chain structure. Both types of structure are held together by the electrostatic attraction between the cations and the negative silicate structures.



(b) The covalent network of SiO_4 in its structure gives asbestos its fire-resistant properties.

(c) Asbestos has been important to the Canadian economy in the past because Canada produced the largest amount of the world's asbestos supply in the early 1900s, and Canada has continued to be a significant producer of asbestos.

(d) Asbestos causes respiratory and cancer-related diseases in humans. Its fibrous structure allows the asbestos to fracture easily and eventually create fibers that are small enough to become airborne, which can be inhaled and cause the development of cancer in the lungs.

(e) Answers may vary. Sample answer: Asbestos mining in Canada should not continue, because the adverse health effects of using the material are known. Plus, selling a cheaper harmful product to other countries, mainly developing ones where it is legal for use, is not an ethical business practice.