Chapter 4 Chemical Bonding and Properties of Matter

Answers to Learning Check Questions

(Student textbook page 211)

1. Three types of bonding are ionic, covalent, and metallic. All chemical bonds involve the electrons of atoms interacting with the other atoms of the same or different elements.
2. The general trend is for electronegativity to decrease down a group and increase across a period from left to right. Electronegativity values are a measure of how strongly an atom of a given element attracts shared electrons in a bond. A comparison of the electronegativity of two elements that are bonded gives information about the characteristics of the bond.
3. Bonds with a ΔEN between 1.7 and 3.3 are classified as mostly ionic. Bonds with a ΔEN between 0.4 and 1.7 are classified as polar covalent. Bonds with a ΔEN between 0.0 and 0.4 are classified as mostly (non-polar) covalent.
4. Since metal atoms cannot attract and hold electrons of other atoms well enough to form filled valence shells, the valence electrons of metals in the solid or liquid state have the ability to move freely from one atom to the next. The electrons are said to be *delocalized*, because they do not remain in one location. This is often referred to as the electron-sea model of metals, in which a metal is described as being a relatively ordered array of cations in a "sea" of freely moving electrons, with the positively charged ions all being attracted to many of the electrons in the "sea" simultaneously.
5. False. Change "disordered" to "relatively ordered" array of cations.
6. Metals are made up of aggregates of millions of tiny crystals, called grains, which range in size from a few nanometres to several millimetres, depending on the metal and the conditions under which it has formed. The atoms of the crystals (grains) form precise, regularly repeating patterns, whereas the atoms at the boundaries between grains are arranged randomly.

(Student textbook page 214)

7. High melting and boiling points indicate strong attractions between atoms. A large amount of kinetic energy is required to pull the particles apart. The kinetic energy of the particles is directly related to the temperature of the substance. A high temperature will be required to provide sufficient kinetic energy to cause melting and boiling.
8.

Periodic Table	Trend	Explanation
Down a group	Decrease	As you go down a group, the atoms have one more electron shell than atoms in the previous group. Therefore, the free valence electrons are progressively farther from the nucleus and the strength of the attractive forces decreases, resulting in lower melting and boiling points.
Across a period (except for metals with electrons in *d* orbitals)	Increase	The number of valence electrons increases across a period and the ions have a larger positive charge. The larger number of electrons in the "electron sea" and the larger charge on the ions result in a stronger attractive force. As the ions are held in place more tightly, the melting and boiling points increase.
Group 12	Decrease	Since *d* shells in Group 12 are filled, electrons cannot freely move away from the atom. There are fewer electrons in the "electron sea." There will be less attraction between the positively charged ions and the electrons in the "electron sea." The melting and boiling points decrease.

9. The valence electrons in metals are free to move from one atom to another. When a potential difference is applied to a piece of metal, the electrons are drawn toward the positive end and repelled from the negative end. Similarly, metals conduct heat when freely moving electrons receive kinetic energy from the source of heat and pass the energy along to other electrons.
10. Metals can be hammered or stretched without breaking, because the ions in the metal crystals can slide past one another without a change in the sea of electrons that surrounds them. Refer to Figure 4.6 on page 213 of the student textbook.
11. The hardness of a metal depends on the size of its crystalline grains. Metals are malleable because the atoms in a crystal can slide over one another. The boundaries of grains resist this sliding, because the atoms are not aligned in layers that slide smoothly. A metal that has a very large number of small crystals has a greater boundary surface and is harder than a metal that has a larger grain size with fewer boundaries.

12. *Substitutional alloy only*: the atoms of the different metals are similar in size and one kind of atom readily takes the place of a different kind.

Substitutional alloy and interstitial alloy: In both types of alloys, electrons attract any kind of positive ion so that all the ions are held together in the same "sea."

Interstitial alloy only: atoms of one or more of the alloy metals are smaller than those of another metal and will fit into spaces between the larger atoms, making this type of alloy stronger.

(Student textbook page 218)

13. A crystal lattice is a three-dimensional pattern of alternating positive and negative ions in an ionic solid. The ratio of positive to negative ions in a crystal lattice is the smallest whole-number ratio of ions in the crystal and is called a formula unit.

14. The relative size and the charge on ions affect how they pack together so that oppositely charged ions are as close together as possible.

15. Similarities: In one formula unit of each compound, the ratio of positive ion to negative ion is 1:1 and both form cubic-shaped crystals. Differences: In NaCl, each Na^+ is surrounded by 6 Cl^-, and each Cl^- is surrounded by 6 Na^+; the Na^+ and Cl^- alternate in a row; Na^+ and Cl^- alternate above and below one another; the Na^+ is about half as large as Cl^-. In CsCl, each Cs^+ is surrounded by 8 Cl^-, and each Cl^- is surrounded by 8 Cs^+; layers of Cs^+ alternate with layers of Cl^-; layers of oppositely charged ions are offset so that Cs^+ is not directly above a Cl^-; Cs^+ is more than twice as large as Cl^-.

16. **a.** High melting point in 900°C range and boiling point in 1500°C range

b. Hard and brittle

c. Not malleable

d. Good electrolyte, conducts an electric current

e. Non-conductor of electric current

17. This is a cut-away three-dimensional image in which the ions on the sides of each unit cell are shared by the next unit cell. If you add up the "halves" and "eights" of the ions, you will find a total of four sodium ions and four chloride ions.

18. For an ionic compound to be soluble in water, the attractive forces between the ions and the water molecules must be stronger than the attractive forces among the ions themselves. For ionic solids of low solubility, the attractive forces between the ions are stronger than the attractive forces between the ions and the water molecules.

(Student textbook page 223)

19. A covalent bond exists because of electrostatic forces. The nuclei of both atoms exert attractive forces on both of the shared electrons. Repulsive forces exist between positive nucleus and positive nucleus, and between negative electron and negative electron. The bond between two atoms occurs at a distance where the repulsive forces balance the attractive forces.

20. A σ bond is a bond that is formed by the overlap of half-filled atomic orbitals that are symmetrical around the bond axis of the two nuclei; for example, overlap of two half-filled *s* orbitals, two half-filled *p* orbitals, or half-filled *s* and *p* orbitals. The σ bond formed has an energy level that is lower than the energy of the half-filled atomic orbitals, making the molecule more stable than the separate atoms.

21. $[Ne]3s^23p^2$; Mix one 3*s* orbital with three 3*p* orbitals to produce four $3sp^3$ orbitals.

22. Ethene is a planar molecule, meaning that all of its atoms lie in the same plane. All of the bond angles in the molecule are approximately 120° and there are three orbitals of the carbon atoms that form a plane. This planar structure is explained by three sp^2 hybrid orbitals formed by mixing the 2*s* orbital and two of the 2*p* orbitals of carbon. One σ bond forms by the overlap of a half-filled sp^2 hybrid orbital along the axis between the two carbon atoms, and the other sp^2 hybrid orbitals from each carbon overlap with the half-filled *s* orbital from a hydrogen atom. These are also σ bonds. The lobes of the half-filled pure *p* atomic orbital not hybridized from each carbon atom overlap above and below the axis of the σ bond to form a π bond between the two carbon atoms.

23. Like a hybrid offspring resulting from crossing two species, a hybrid orbital results from the combination of atomic orbitals of the valence electrons. For example, a sp^2 hybrid orbital results from combining one *s* atomic orbital and two *p* atomic orbitals.

24. $H-C\equiv N:$

The triple bond between the carbon atom and the nitrogen atom consists of one sigma bond and two pi bonds. One σ bond forms by the overlap of a half-filled *sp* orbital of C with a half-filled 2*p* orbital of N. Two π bonds form from the overlap of the half-filled *p* orbital from C with two of the half-filled 2*p* orbitals of N. The second half-filled *sp* orbital of C overlaps with the half-filled 1*s* orbital of H to form a σ bond.

(Student textbook page 230)

25. Drawing a Lewis structure for a molecule shows how many electrons are involved in each bond, as well as how many, if any, lone pairs of electrons are present.

26. A line represents a shared pair of electrons or a covalent (or polar covalent) bond.

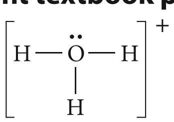
27. A co-ordinate covalent bond is a covalent bond in which one of the atoms contributes both of the shared electrons. Refer to Figure 4.27 on page 229 of the student textbook.

28. An expanded valence refers to more than eight electrons in the valence shell of a bonded atom. This is possible for a central atom that forms hybrid orbitals that include *d* orbitals, such as sp^3d and sp^3d^2 hybrids. Refer to SF_6(g) in Figure 4.28 on page 229 of the student textbook.

29. Carbon cannot have an expanded valence as a central atom because it does not have *d* orbitals in the second energy level, and the *d* orbitals in its third energy level are too high in energy to combine with *s* or *p* orbitals in the second energy level.

30. This Lewis structure is not accurate because experimental evidence has indicated that all of the oxygen–oxygen bonds are the same in length and energy. In this depiction of the molecule, there are single and double bonds that would be of different bond lengths. The structure is explained by using two Lewis diagrams, which if "averaged," would give the accurate picture. These representations are called hybrid resonance structures.

(Student textbook page 237)

31. 

Oxygen is sp^3 hybridized from mixing one 2*s* orbital with three 2*p* orbitals producing four sp^3 orbitals. Three of the sp^3 orbitals form a bond with hydrogen atoms and the fourth sp^3 orbital becomes the lone pair.

32. sp^3d

33. sp^2

34. **a.** sp^3, the N–H bond angle is a little less than the tetrahedral angle, about 107°.

b. sp^3, the N–H bond angle is the tetrahedral angle of 109.5°.

35. The condensed configuration for Br is [Ar]$4s^23d^{10}4p^5$. To form hybrid orbitals, the bromine atom uses the 4*s* orbital, three 4*p* orbitals, and two 4*d* orbitals to give six sp^3d^2 hybrid orbitals. Five of these hybrid orbitals are half-filled and bond with the five F atoms. The other contains a lone pair of electrons. The overall shape is square pyramidal.

36. The condensed configuration for Cl is [Ne]$3s^23p^5$. To form hybrid orbitals, the chlorine atom uses the 3*s* orbital, three 3*p* orbitals, and one 3*d* orbital to give five sp^2d hybrid orbitals. Three of these hybrid orbitals are half-filled and bond with the three F atoms. The other two contain two non-bonding electrons. The overall shape is T shaped.

Answers to Caption Questions

Figure 4.1 (Student textbook page 209): For the transition elements, electrons are added to *d* orbitals, which are filled after the $(n + 1)s$ orbital in any period. The *d* orbitals have a variety of shapes and half-filled *d* orbitals are energetically more favourable. This results in the nuclear charge having a varying effect on this shell and the trend in electronegativity will be somewhat irregular.

Figure 4.15 (Student textbook page 218): Current will flow through a solution if charged particles are free to move independently. When an ionic compound is dissolved in water, the ions are free to move past one another and will migrate to the electrodes. This will be registered as current flow on the conductivity meter.

Figure 4.35 (Student textbook page 239): The molecule CH_3OH is polar. The molecular shape is tetrahedral but the C–O and O–H bonds are polar. This results in a permanent dipole in the molecule.

Figure 4.36 (Student textbook page 242): The strength of a dipole-dipole force depends on the size of the charge difference between the two poles of the molecule. This depends on the difference in the electronegativity of the two atoms in the molecule that results in it being polar. For example, the polarity of the F–Cl molecule is different from the polarity of the F–Br molecule because Cl and Br have different electronegativity values.

Answers to Section 4.1 Review Questions

(Student textbook page 227)

1. Formula unit: a formula that describes the type and ratio of ions in an ionic compound

Unit cell: The smallest repeating pattern of particles in a crystal

Molecule: A neutrally charged combination of atoms held together by covalent bonds

2. *Substitutional alloy*: a mixture of atoms of two or more different metal elements that consists of many tiny crystals of atoms in which atoms of one or more metals take the place of atoms of the most abundant metal

   Interstitial alloy: a mixture of two or more metals in which the most abundant metal forms a crystalline structure with atoms of one or more metals that have small atoms that fit into the spaces between the larger atoms

3. **a.** The rapid cooling prevents large crystalline grains from forming. The small grains make the metal very hard.

   b. The sword will probably be tempered by heating to a temperature below the melting point and cooling it slowly to allow the crystalline grains to become larger. The process makes the metal less brittle.

4.

	metals	ionic crystals
melting and boiling points	very high	high
electrical conductivity	good	none in its solid form
malleability	very	none
hardness	most are very hard	hard

5. Sodium fluoride and all other ionic compounds have no sets of ions that can be identified as a unit or molecule because each ion is equally attracted to the adjacent oppositely charged ions. Thus, ionic compounds do not form molecules.

6. Yes. Atoms of elements on the left side of the periodic table form positive ions and atoms of elements on the far right side of the periodic table tend to form negative ions. These oppositely charged ions are attracted to one another and form ionic compounds.

7. **a.** H–N is the most covalent; ΔEN for H–N = 0.9; ΔEN for H–O = 1.4; ΔEN for H–F = 1.9

   b. O–O is the most covalent; ΔEN for O–O = 0.00; ΔEN for O–N = 0.5; ΔEN for O–C = 1.0

8. **a.** mostly ionic; $\Delta EN = 3.2$

   b. mostly covalent; $\Delta EN = 0.0$

   c. polar covalent; $\Delta EN = 0.5$

   d. mostly ionic; $\Delta EN = 2.5$

9. **a.** One *s* orbital, three *p* orbitals, and one *d* orbital

   b. Five hybrid orbitals

   c. trigonal bipyramidal

10. *Both*: very hard, extended structures

    Ionic crystals: very brittle and can easily be crushed by pressure, consist of oppositely charged ions and form crystals

    Network solids: consist of atoms of one or two elements connected by covalent bonds. Extremely hard and not brittle

11. The pattern in which the carbon atoms are bonded together is very different in diamond and graphite. In diamond, each carbon atom in the crystal is at the centre of a tetrahedron with another carbon atom at each corner, giving a three-dimensional network solid. Diamond is very hard with several planes of atoms within the crystal that can reflect light. In graphite, each carbon atom is bonded to 3 other carbon atoms in a trigonal planar orientation. This pattern repeats in layers that are weakly attracted so that these layers can slide over one another. The fourth pair of electrons is delocalized, which allows this form of carbon to conduct an electric current.

    Applications:
    Diamond: cutting of very hard substances, jewellery
    Graphite: lubrication, writing

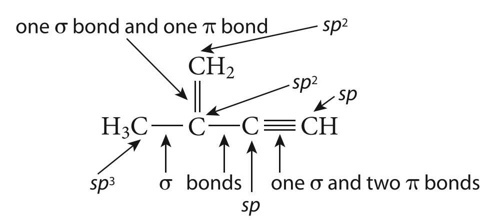
12. **a.** sp^3

    b. sp^2 with one unhybridized *p* orbital

    c. sp^2 with one unhybridized *p* orbital

    d. sp^2 with one unhybridized *p* orbital

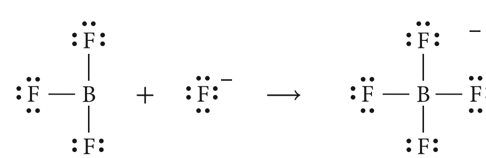
13.



Answers to Section 4.2 Review Questions

(Student textbook page 247)

1. Co-ordinate covalent bond: a bond in which both of the electrons were contributed by one of the atoms



BF_3 F^- BF_4^-

2. *Sample answer*: Two or more possible Lewis structures of the same molecule in which the structures differ only in the positioning of their bonding and lone pairs. The actual structure of the molecule is most likely a combination, or hybrid, of all possible Lewis structures. They are used when no single Lewis structure properly describes the observed characteristics of the molecule. For example, when measurements of bond lengths are between the lengths of a single and a double bond but a Lewis structure cannot be drawn to represent it. (See examples on page 230 of the student text.)

3. To determine the shape of a molecule, you need to know the total number of electron groupings (bonding pairs and lone pairs) around the central atom. A Lewis structure is needed in order to determine the number of electron groupings.

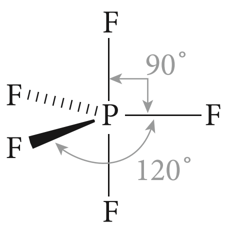
4. a. bent

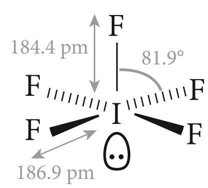
b. trigonal pyramidal

c. trigonal pyramidal

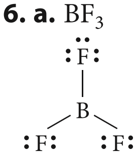
d. linear

5. PF_5 has five electron groupings, all of which are bonding pairs. IF_5 has six electron groupings and one of them is a lone pair.



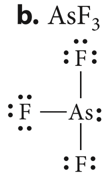


6. a. BF_3



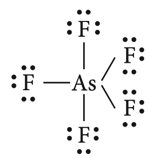
Boron trifluoride has three electron groupings around the boron atom so it is trigonal planar.

b. AsF_3



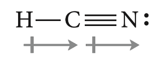
Arsenic trifluoride has four electron groupings (three bonding pairs and one lone pair) so the electron pairs have a tetrahedral arrangement, but the shape is trigonal pyramidal.

c. AsF_5



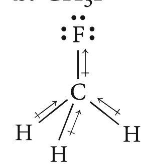
Arsenic pentafluoride has five electron groupings so it is trigonal bipyramidal.

7. a. HCN



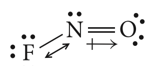
The shape is linear and the molecule is polar.

b. CH_3F



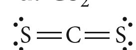
The shape is tetrahedral and the molecule is polar.

c. NOF



The shape is bent, and the molecule is polar.

d. CS_2



The molecule is linear and the bonds are not polar, so the molecule is non-polar.

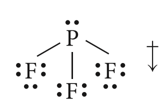
8. I. Check for solubility in water. The polar compound should dissolve but the non-polar compound should not dissolve.

II. Check for the melting point. The polar compound will have a higher melting point than will the non-polar compound.

III. Check for the boiling point. The polar compound will have a higher boiling point than will the non-polar compound.

9. In *sp* hybridization, one *s* orbital and one *p* orbital combine to form two identical *sp* orbitals. In sp^2 hybridization, one *s* orbital and two *p* orbitals combine to form three identical sp^2 orbitals. Hydrogen cyanide has *sp* orbitals. Carbon dioxide has sp^2 orbitals.

10. a.



b. There are four electron groupings with one being a lone pair. Therefore, the electron group arrangement is tetrahedral.

c. The hybridization of the orbitals in the phosphorus atom is probably sp^3.

d. With three bonding pairs, the molecular shape is trigonal pyramidal.

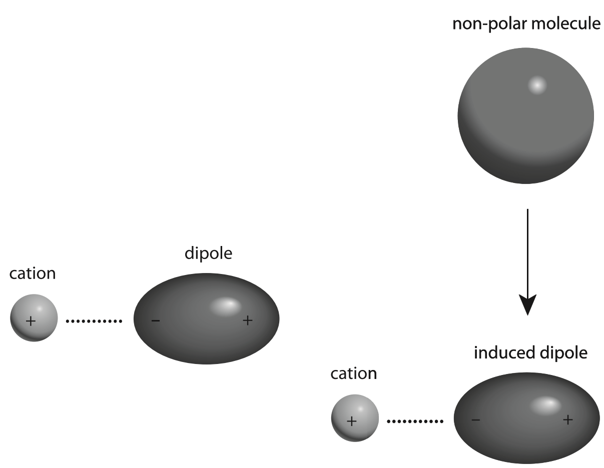
11. a. *dispersion forces*: Neon consists of individual atoms that do not form ions, so polarity is not a factor. Therefore, the only possible type of bonding would be dispersion forces.

b. *hydrogen bonding (a special case of dipole-dipole bonding)*: Each water molecule has two –OH bonds so a single molecule can form hydrogen bonds with several other water molecules.

c. *dipole-dipole bonding*: $CHCl_3$ is slightly polar so each molecule is a dipole.

d. *dispersion forces*: BF_3 is trigonal planar in shape so it is completely symmetrical. The bond polarities cancel each other out. Therefore, it is a non-polar molecule.

12. In an ion-dipole force, an ion and the oppositely charged end of a dipole attract each other electrostatically. An ion-induced dipole force acts between an ion and a non-polar molecule. When the ion approaches the non-polar molecule, it draws negative charges within the molecule toward itself and repels the positive charges in the molecule. These forces cause the once non-polar molecule to become a dipole. Once formed, the ion attracts the oppositely charged end of the induced dipole.



13. a. dipole-dipole forces

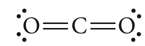
b. ion-dipole forces

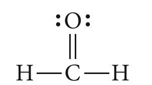
14. The observation that non-polar gases such as the noble gases and the halogens will form liquids, and in some cases solids, when the temperature is lowered sufficiently, demonstrates that some type of force must be acting between the molecules.

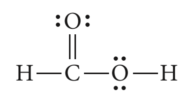
Answers to Practice Problems

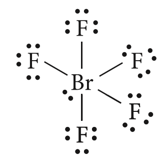
For full solutions to Practice Problems, see Part B of this Solutions Manual.

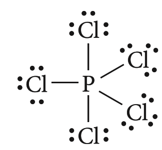
(Student textbook page 232)

1. 

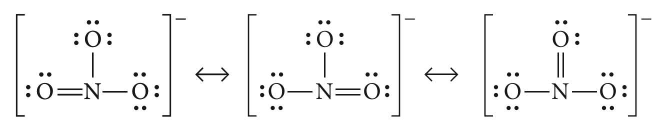
2. 

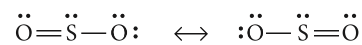
3. 

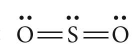
4. 

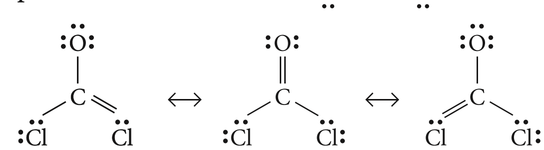
5. 

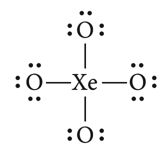
6.

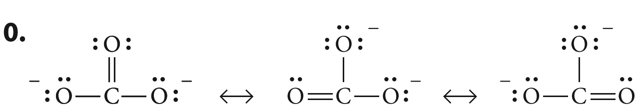


7. resonance structure: 

expanded valence structure: 

8. 

9. 

10. 

(Student textbook page 236)

11. a. trigonal planar

b. square pyramidal

12. a. 32

b. 24

13. **a.** 4 BP, 0 LP
b. 2 BP, 0 LP
c. 4 BP, 2 LP
d. 2 BP, 1 LP

14. bent

15. trigonal pyramidal

16. tetrahedral

17. tetrahedral, 109.5°

18. trigonal planar

19. OF_2, NF_3, CF_4, BF_3, BeF_2

20. trigonal bipyramidal; tetrahedral; octahedral

Answers to Chapter 4 Review Questions

(Student textbook pages 253-7)

1. e	**8.** d
2. d	**9.** a
3. c	**10.** c
4. a	**11.** b
5. d	**12.** c
6. b	**13.** a
7. c	**14.** e

15. Metals and non-metals: Metals form positive ions and non-metals form negative ions. When metals donate electrons and non-metals accept electrons, both types can form a complete octet of electrons and, therefore, become stable.

16. Non-metals: The valence shells of non-metals are more than half full but they do not lose electrons so they cannot accept electrons from each other to form a complete octet of electrons. They can, however, share electrons and form a complete octet and become stable.

17. The metal is heated to a temperature at which it begins to become soft. This allows the atoms to move somewhat randomly. The metal is then plunged into cold water, which prevents crystalline grains from growing. When the metal is cool, it consists of very tiny grains with atoms situated randomly in the boundaries around the grains. This makes the metal hard because the boundaries resist sliding.

18. **a.** Delocalized means that the electrons that originate from one atom can move away from that atom and become a part of a larger group of shared electrons (the "sea" or "pool" of electrons surrounding a cluster of metal cations). In ionic bonds, the electrons are transferred specifically from one to another atom and held there to form discrete ions. In covalent bonds, an electron from one atom pairs up with an electron from another atom and is held in a location near those two nuclei.
b. Metallic bonds can be viewed as non-directional. The positive ions, layered in a very organized pattern, will slide over one another when stress (such as a pounding hammer) is applied, and not break the pattern. Throughout the impact, the pool of delocalized electrons helps to keep the metal together by continuing to exert a uniform attraction on the positive ions.

19. Ionic solids are brittle because the oppositely charged ions are aligned in a regular pattern. When the crystals are stressed, the ions along a plane move in a way that causes like charges to be aligned beside each other. The repulsive electrostatic force then causes the crystal to break apart.

20. Single bonds form between two atoms when unfilled orbitals overlap along a line between the two nuclei. The electrons are then shared between the two atoms. According to molecular orbital theory, when two atomic orbitals overlap, a new type of orbital, a molecular orbital, is formed. The sharing of the electrons in the molecular orbital constitutes the bond.

21. All four of the C–H bonds in methane are identical in length and strength. Thus the four orbitals of the carbon that have overlapped with an orbital in the hydrogen atom have to be identical. Thus they cannot consist of one *s* and three *p* orbitals.

22. linear, trigonal planar, tetrahedral, trigonal bipyramidal, and octahedral

23. A non-polar molecule can move close to (1) an ion, or (2) a dipole. The charge on the ion or the end of the dipole will distort the electron distribution of the once non-polar molecule, making it a temporary dipole.

24. In symmetrical molecules, any polar bonds are aligned so that the polarities of the bonds cancel, leaving the molecule non-polar. If an asymmetrical molecule has any polar bonds, the asymmetry will prevent them from cancelling each other.

25. Solubility in polar and non-polar substances; miscibility with other polar and non-polar substances; relative melting and boiling points; physical state at room temperature; hardness; surface tension

26. **a.** have a crystalline structure, generally have high melting and boiling points
b. metallic substances are not brittle but ionic substances are, metallic substances are malleable and ductile but ionic substances are not, metallic solids

conduct electric current but solid ionic substances do not

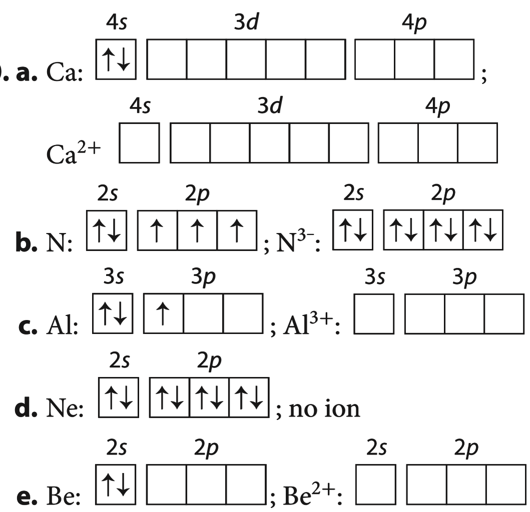
27. a. often soluble in water, brittle

b. ionic solids have a high melting point and molecular solids do not, ionic solids are much harder than polar molecular solids

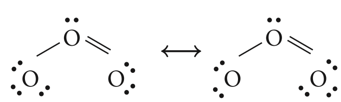
28. BF_3 is a trigonal planar molecule and the three polar bonds are positioned such that they cancel each other. Water is a V-shaped molecule and the two polar bonds are not along the same line and therefore cannot cancel each other.

29. The solubility will probably increase as the melting point decreases. The lower melting point of rubidium iodide indicates that the attractive forces holding the ions together are not as strong as in rubidium bromide, which has a higher melting point. If the force is not as strong, the polarity of the water molecules is more likely to be able to pull the ions apart.

30.



31. a. There are three electron groupings around the central oxygen so the electron group arrangement would be trigonal planar. However, there is one lone pair so the molecular shape would be bent.



b. Quite close.

32. a. H–F is the most ionic; ΔEN for H–Cl = 0.9; ΔEN for H–Br = 0.7; ΔEN for H–F = 1.9

b. K–O is the most ionic; ΔEN for Na–O = 2.6; ΔEN for Li–O = 2.5; ΔEN for K–O = 2.7

33. a. ionic; $\Delta EN = 2.0$

b. covalent; $\Delta EN = 0.0$

c. polar covalent; $\Delta EN = 0.5$

d. ionic; $\Delta EN = 2.6$

34. a. lithium: [He]$2s^1$

b. argon: [Ne]$3s^2 3p^6$

c. chlorine: [Ne]$3s^2 3p^5$

d. phosphorus: [Ne]$3s^2 3p^3$

35.

	Electron Grouping	Molecular Shape	Bond Angles
a.	trigonal bipyramidal	seesaw	90º/120º
b.	octahedral	octahedral	90º
c.	trigonal bipyramidal	T-shape	<90º
d.	trigonal planar	trigonal planar	120º
e.	tetrahedral	bent	<109.5º
f.	octahedral	square planar	90º

36. a. non-polar

b. polar

c. polar (but it does not exist as a single unit)

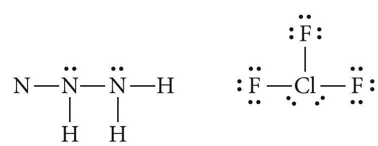
d. polar

37. The molecules of the two different liquids are attracted to one another more strongly than to other like molecules and therefore, they pack together more closely than the molecules in the separate liquids. The tight packing causes the volume of the combined liquids to be smaller than the sum of the liquids separately.

38. *Sample answer*: First, try to mix the unknown liquid with water which is polar. If the unknown liquid does not mix with water, it is non-polar. Next, try to mix the unknown liquid with a non-polar liquid such as a light oil. If it mixes, it is probably non-polar. If it does not mix, the unknown liquid is probably polar.

39. In co-ordinate covalent bonds, one of the atoms contributes both of the electrons for the bond.

40. a.



b. hydrazine: sp^3 ; chlorine trifluoride: sp^3d

c. T-shaped; polar 41. Moh's hardness scale was developed in 1832 to compare the hardness of minerals. The scale is from 1–10 with talc being the softest mineral having a Moh's value of 1 and diamond being the hardest having a Moh's value of 10. To determine hardness, the metal or mineral is scratched. Scratching breaks metallic bonds as atoms are displaced. On Moh's scale, the hardness of some metals is: Pb 1.5; Au, Ag 2.5-3; Cu 3; Ni 4, W 7.5; hardened steel 7-8. Alloying a metal with another generally increases its hardness. For example, pure gold is malleable and easily worked for jewellery but is too soft to support the holding of a stone or gem. It is alloyed with other metals like Cu and Zn to increase its strength.

42. a. 2 BP, 3 LP
linear shape with 180º angles

b. 4 BP, 0 LP
tetrahedral shape with 109.5º angles

c. 5 BP, 0 LP
trigonal bipyramidal with 90º (axial) 120º (equatorial)

43. a. NH_3; N–H bonds are polar and P–H bonds are not, NH_3 has hydrogen bonding and is a dipole having dipole-dipole attractions.

b. C_4H_{10}; C_4H_{10} is larger, with more electrons, so a greater number of dispersion forces can form between molecules

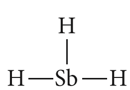
c. $SeCl_4$; $SeCl_4$ has a seesaw shape and is therefore polar allowing dipole-dipole forces between the molecules. $SiCl_4$ is tetrahedral and thus symmetrical, making it non-polar

44. a. O_2; oxygen has more electrons causing the dispersion forces to be greater than for N_2

b. ethanol; Ethanol has an –OH group and can form hydrogen bonds whereas methoxymethane does not. Therefore, ethanol molecules are more strongly attracted to one another and thus would have a higher boiling point.

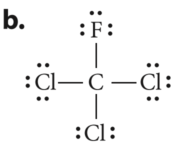
c. heptanes; Molecules of both compounds are identical in size and therefore have the same number of electrons. They are both non-polar so dispersion forces are the only intermolecular forces. Heptane is linear whereas 2,4-dimethyl pentane is spherical. Therefore, Heptane molecules can have larger surface areas in contact with one another and thus experience stronger dispersion forces.

45. a.



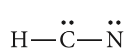
3 BP; 1 LP

b.



4 BP; 12 LP (total in molecule)

c.

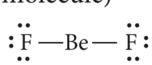


2 BP; 2 LP

d. H : C ::: C : H

5 BP, 0 LP

e.



2 BP; 6 LP

46. *Ionic*: crushes easily in a mortar with a pestle (brittle), dissolves in water, conducts electric current in the dissolved state but not the solid state, does not melt in a deflagrating spoon over a Bunsen burner flame

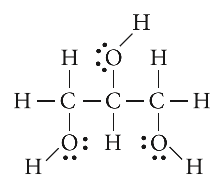
Metallic: conducts electric current in the solid state, does not crush in a mortar but rather deforms (malleable, not brittle), spreads to form a sheet when pounded with a hammer, lustrous, grey, may or may not melt in a deflagrating spoon over a Bunsen burner flame, but can deform more easily in a mortar

Covalent Network: does not crush or deform in a mortar with a pestle, does not conduct electric current in the solid state, cannot dissolve in water or in acetone

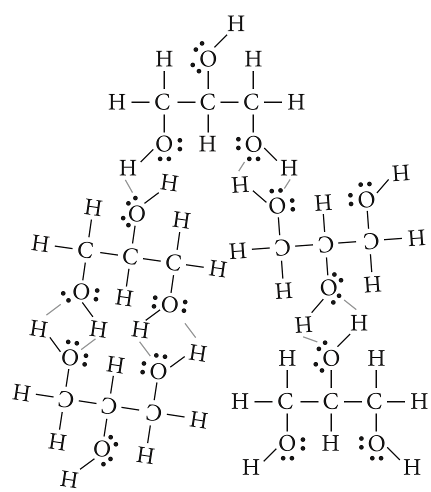
Molecular (Polar): crushes easily in a mortar with a pestle, dissolves in water, cannot conduct electric current in the dissolved state, can melt in a deflagrating spoon held in a Bunsen burner flame

Molecular (Non-polar): soft (can be scratched easily), can be deformed with relatively little pressure in a mortar with a pestle but does not shatter (not brittle), cannot dissolve in water, can dissolve in acetone, cannot conduct electric current in any state, melts very readily in a deflagrating spoon held in a Bunsen burner flame

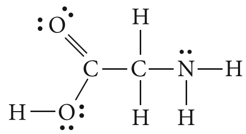
47. a.

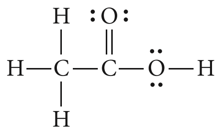


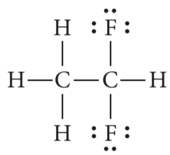
b. The lighter/grey lines represent the hydrogen bonds.

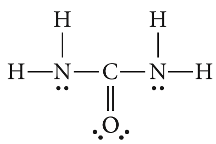


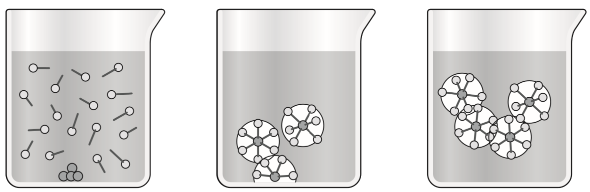
48. The report should state the general definition, list advantages and adverse effects, uses, toxicity, levels needed to have an effect, and examples, including the chemical structure.

49. a. 

b. 

c. 

d. 

50. 

51. Student answers should show the following steps (from page 228 of the student textbook):

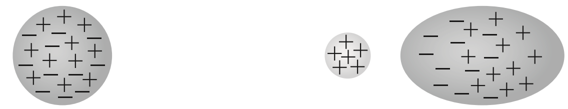
Step 1 Position the least electronegative atom in the centre of the molecule or polyatomic ion. Draw a skeleton structure for the molecule by placing the other atoms around the central atom. Draw a single bond between each pair of atoms. Always place a hydrogen atom or a fluorine atom at an end position in the structure.

Step 2 Determine the total number of valence electrons (V) in all the atoms in the molecule. For polyatomic ions, add or subtract electrons to account for the charge. For example, for a cation with a charge of 2+, subtract two electrons from the total number of valence electrons that are calculated for the ion.

Step 3 Determine the total number of electrons (T) needed for each atom to achieve a noble gas electron configuration. For all atoms besides hydrogen this corresponds to the octet rule. For hydrogen, a complete valence shell is two electrons, not eight.

Step 4 Subtract the number of valence electrons (V) from the number of electrons needed to satisfy the octet rule (T). This represents the number of shared electrons (S) involved in bonding, $S = T - V$. Divide this number by 2 to give the number of bonds, bonds $= S/2$. Double bonds count as two bonds. Triple bonds count as three bonds.

Step 5 Subtract the number of shared electrons from the number of valence electrons to get the number of non-bonding electrons, $NB = V - S$. Add these electrons as lone pairs to the atoms to achieve a noble gas electron configuration for each atom.

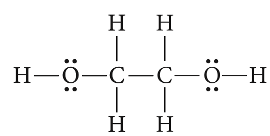
52. 

53. First, use a magnet to remove the nickel balls. Second, add water to dissolve the table salt. Filter to remove the solids—the glass balls and wax. Allow the water to evaporate to obtain the dry salt. Next, heat the solids to melt the wax. Pour off the wax and let it cool. The only remaining solids are the glass balls.

54. A graphic organizer such as a spider map, fishbone diagram, or a concept map should identify key concepts and relationships among the concepts shown in the Chapter 4 Summary on page 252 of the student textbook. Concepts relate to models of chemical bonding, shapes, intermolecular forces, and properties of molecules.

55. a. Group 1 metals have very low first ionization energy values. The first ionization energy decreases as you go down the group so cesium has the lowest ionization energy of all of the stable elements. Thus, electrons are easily removed from cesium. Therefore, it would not require much light energy to activate the photovoltaic cell.

b. Rubidium, because its first ionization energy is also very low.

56. a. 

b. It will have a higher boiling point than water. Because it can form hydrogen bonds at both ends of the molecule, it will require even more energy to change state. (B.P. = 197.3°C)

c. Yes, through the formation of hydrogen bonds between the two compounds, ethylene glycol and water can mix very thoroughly.

When ethylene glycol and water mix, they form hydrogen bonds with each other.

d. In order for water to freeze, it needs to form hydrogen bonds with other water molecules. Ethylene glycol prevents water from freezing because it interferes with the ability for water molecules to form hydrogen bonds with each other.

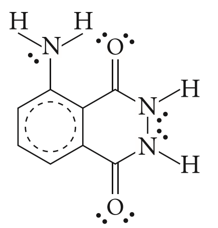
57. a. $[Kr]5s^2 4d^{10} 5p^3$

b. Sb^{-3}: $[Kr]5s^2 4d^{10} 5p^6$, the antimony atom gains 3 electrons to have a completely filled set of 5*p* orbitals and be more stable; Sb^{+3}: $[Kr]4d^{10} 5s^2$, the antimony atom loses its 3 highest energy electrons from the 5*p* orbitals to be more stable;

Sb^{+5}: $[Kr]4d^{10}$, the antimony atom loses all of the 5 electrons from its highest energy level to be more stable.

c. A metalloid is an element that can behave as a metal as well as a non-metal. The electron configurations in b. show that antimony can become an anion, typical of non-metals, as well as a cation, typical of metals.

58. a.

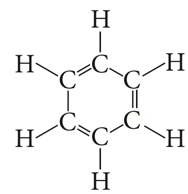


b. Luminol is mainly non-polar with some dipolar regions. It is insoluble in water but soluble in most polar organic solvents such as alcohols, aldehydes, and ketones.

c. Luminol has some drawbacks that may limit its use in a crime scene investigation:

- Luminol chemiluminescence can also be triggered by a number of substances such as copper or copper-containing alloys, and certain bleaches. As a result, if a crime scene is thoroughly cleaned with a bleach solution or horseradish, residual cleaner will cause the entire crime scene to produce the typical blue glow, effectively camouflaging any organic evidence, such as blood.
- Luminol will also detect the small amounts of blood present in urine, and it can be distorted if animal blood is present in the room that is being tested.
- Luminol reacts with fecal matter, causing the same glow as if it were blood.

59. a.

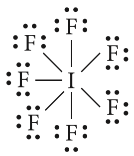


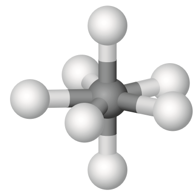
b. The electrons forming the second bond in the double bonds are really delocalized and shared among all six of the carbon atoms in the ring. All of the carbon-carbon bonds are identical in length and strength. The circle represents these delocalized electrons.

60.

Compound	Hybridization Rough Work	Hybrid Orbital Required	Bonding Groups	Lone Pairs	Electron Group Arrangement	Molecular Shape	Bond Angle	Polarity
SnF_2	1 *s* orbital + 2 *p* orbitals = 3 sp^2 orbitals sp^2 sp^2 sp^2 ↑ ↑ ↑	sp^2	2	1	trigonal planar	bent	<120°	polar
NH_3	1 *s* orbital + 3 *p* orbitals = 4 sp^3 orbitals sp^3 sp^3 sp^3 sp^3 ↑ ↑ ↑ ↑	sp^3	3	1	tetrahedral	trigonal pyramidal	<109.5°	polar
PF_5	1 *s* orbital + 3 *p* orbitals + 1*d* orbital = 5 sp^3d orbitals sp^3d sp^3d sp^3d sp^3d sp^3d ↑ ↑ ↑ ↑ ↑	sp^3d	5	0	trigonal bipyramidal	trigonal bipyramidal	90°/ 120°	non-polar
SF_6	1 *s* orbital + 3 *p* orbitals + 2 orbitals = 6 sp^3d^2 orbitals sp^3d^2 sp^3d^2 sp^3d^2 sp^3d^2 sp^3d^2 sp^3d^2 ↑ ↑ ↑ ↑ ↑ ↑	sp^3d^2	6	0	octahedral	octahedral	90°	non-polar

61. The sand fits in between the marbles similar to the way that the smaller metal atom fits in the spaces between the larger metal atoms in an interstitial alloy.

62. a. 

b. 

The bond angle in the equatorial plane is
$\frac{1}{5} \times 360° = 72°$

The bond angle between the axial and equatorial plane is 90° and the bond angle along the axial plane is 180°.

63. The two nail lacquers are both insoluble in water (they do not wash off with water); therefore, they are molecular compounds made up of mostly non-polar molecules. As the top coat dries (the solvent evaporates), cracks start to form in the top layer; therefore, the two nail lacquers are not attracted to each other. The bottom coat shows through the cracks completely unchanged; therefore, the two nail lacquers do not dissolve in one another or react with each other to cause a change in the dried undercoat. Cracks form, not beads like water forming on a surface of oil or plastic; therefore, the rate at which the top coat dries is faster than the rate at which the top coat is repelled away from the undercoat.

64. a. The students' discussions should include the following information. By adding 0.5 to 1.0 percent nickel to iron, engineers can build rotor hubs that hold the giant rotor blades and other components of wind mills. The nickel strengthens the iron without making it brittle. The alloy has good impact strength, even at temperatures as low as −20°C.

b. The students' discussions should include the following information. Cast iron, with the 3 to 4 percent carbon, is a runny liquid when molten and does not shrink when it is cooled in a mould. It is very hard but also quite brittle. Pure iron is not as hard but is more malleable and ductile.

Answers to Chapter 4 Self-Assessment Questions

(Student textbook pages 258-9)

1. d

2. b

3. e

4. e

5. b

6. d

7. b

8. a

9. b

10. e

11. CaKr: Krypton is a noble gas and does not form ions.

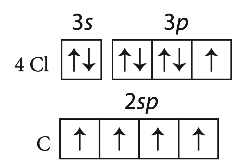
$BaCl_3$: Barium is a Group 2 metal which only forms ions with the charge of 2+. It would not combine with three chloride ions.

NCl_5 will not form since *d* orbitals must be available to accommodate 5 F atoms. There are no *d* orbitals at the 2nd energy level. The closest *d* orbitals are in the 3rd energy level.

12. a. Ionic bonds: Compounds with ionic bonds form ions when they dissolve in water and therefore conduct electric current.

b. Table salt is a white solid at room temperature. It dissolves in water and can conduct electricity.

13. CCl_4: The unpaired 2sp^3 electron in each of four chlorine atoms would be shared with each of the unpaired 2sp^3 electrons in the carbon atom.



14. a. CaO

b. KCl

c. RbCl

15. a. $\Delta EN = 1.8$, therefore ionic

b. $\Delta EN = 3.1$, therefore ionic

c. $\Delta EN = 1.0$, therefore polar covalent

d. $\Delta EN = 0.0$, therefore covalent

16. Both of the electrons in a co-ordinate bond are donated by only one of the atoms. For example, when a hydrogen ion combines with a water molecule to form a hydronium ion, the new bond is a co-ordinate bond because both electrons came from the oxygen atom.