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

2. d

3. c

4. a

5. d

6. b

7. c

8. d

9. a

10. c

11. b

12. c

13. a

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

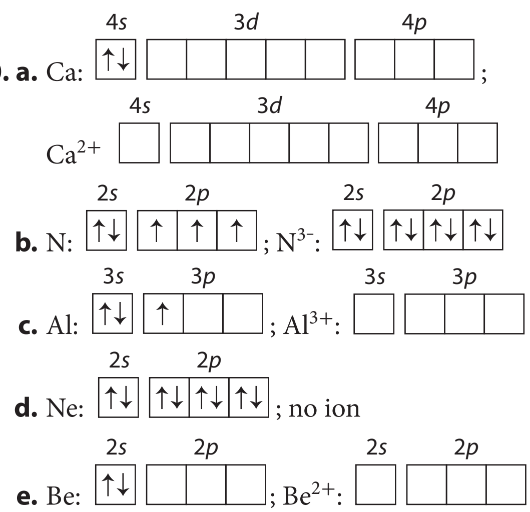
27. a. often soluble in water, brittle

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

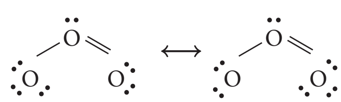
28. BF_3 is a trigonal planar molecule and the three polar bonds are positioned such that they cancel each other. Water is a V-shaped molecule and the two polar bonds are not along the same line and therefore cannot cancel each other.

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	90°/120°
b.	octahedral	octahedral	90°
c.	trigonal bipyramidal	T-shape	<90°
d.	trigonal planar	trigonal planar	120°
e.	tetrahedral	bent	<109.5°
f.	octahedral	square planar	90°

36. a. non-polar

b. polar

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

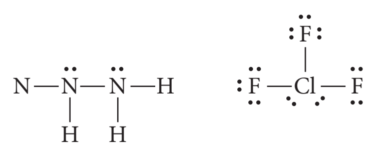
d. polar

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 180° angles

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

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

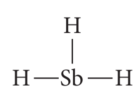
c. $SeCl_4$; $SeCl_4$ has a seesaw shape and is therefore polar allowing dipole-dipole forces between the molecules. $SiCl_4$ is tetrahedral and thus symmetrical, making it non-polar

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.

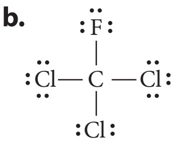
c. heptanes; Molecules of both compounds are identical in size and therefore have the same number of electrons. They are both non-polar so dispersion forces are the only intermolecular forces. Heptane is linear whereas 2,4-dimethyl pentane is spherical. Therefore, Heptane molecules can have larger surface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

45. a.



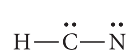
3 BP; 1 LP

b.



4 BP; 12 LP (total in molecule)

c.

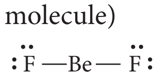


2 BP; 2 LP

d.

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

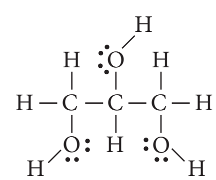
Metallic: conducts electric current in the solid state, does not crush in a mortar but rather deforms (malleable, not brittle), spreads to form a sheet when pounded with a hammer, lustrous, grey, may or may not melt in a deflagrating spoon over a Bunsen burner flame, but can deform more easily in a mortar

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

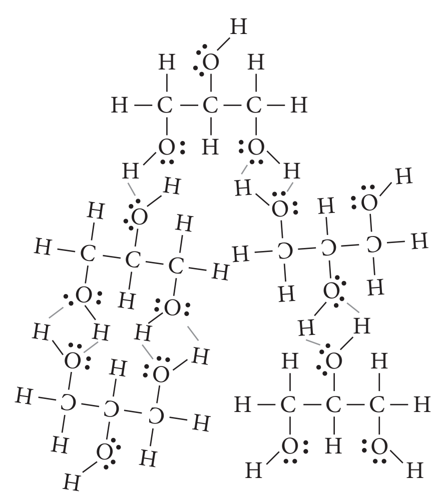
Molecular (Polar): crushes easily in a mortar with a pestle, dissolves in water, cannot conduct electric current in the dissolved state, can melt in a deflagrating spoon held in a Bunsen burner flame

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

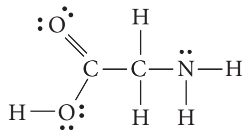
47. a.

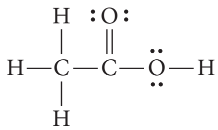


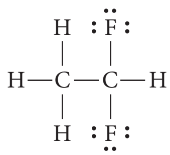
b. The lighter/grey lines represent the hydrogen bonds.

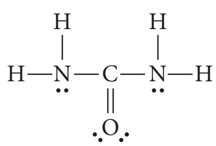


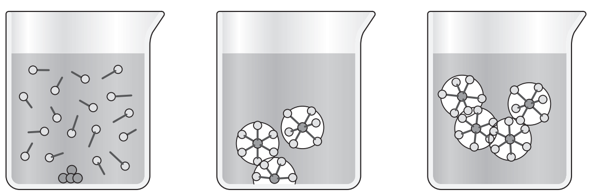
48. The report should state the general definition, list advantages and adverse effects, uses, toxicity, levels needed to have an effect, and examples, including the chemical structure.

49. a. 

b. 

c. 

d. 

50. 

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

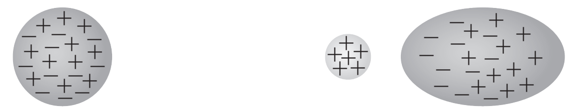
Step 1 Position the least electronegative atom in the centre of the molecule or polyatomic ion. Draw a skeleton structure for the molecule by placing the other atoms around the central atom. Draw a single bond between each pair of atoms. Always place a hydrogen atom or a fluorine atom at an end position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

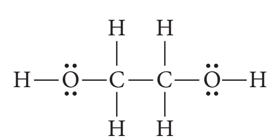
52. 

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.

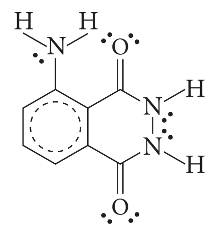
57. a. $[Kr]5s^2 4d^{10} 5p^3$

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

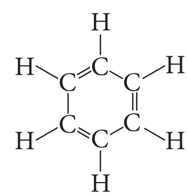


b. Luminol is mainly non-polar with some dipolar regions. It is insoluble in water but soluble in most polar organic solvents such as alcohols, aldehydes, and ketones.

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.



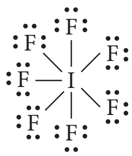
b. The electrons forming the second bond in the double bonds are really delocalized and shared among all six of the carbon atoms in the ring. All of the carbon-carbon bonds are identical in length and strength. The circle represents these delocalized electrons.

60.

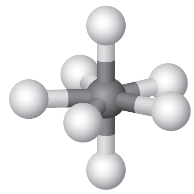
Compound	Hybridization Rough Work	Hybrid Orbital Required	Bonding Groups	Lone Pairs	Electron Group Arrangement	Molecular Shape	Bond Angle	Polarity
SnF_2	1 *s* orbital + 2 *p* orbitals = 3 sp^2 orbitals sp^2 sp^2 sp^2 ↑ ↑ ↑ — — —	sp^2	2	1	trigonal planar	bent	<120°	polar
NH_3	1 *s* orbital + 3 *p* orbitals = 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.

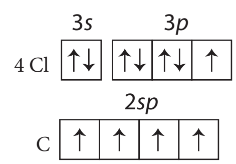
$BaCl_3$: Barium is a Group 2 metal which only forms ions with the charge of 2+. It would not combine with three chloride ions.

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

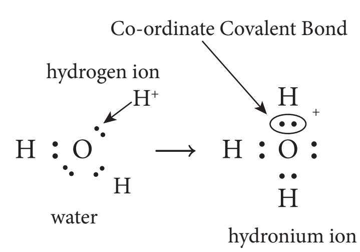
15. a. $\Delta EN = 1.8$, therefore ionic

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

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

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

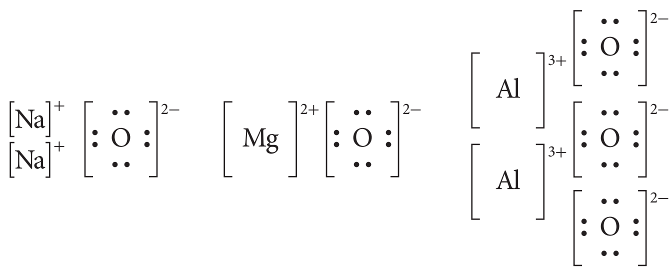
16. Both of the electrons in a co-ordinate bond are donated by only one of the atoms. For example, when a hydrogen ion combines with a water molecule to form a hydronium ion, the new bond is a co-ordinate bond because both electrons came from the oxygen atom.



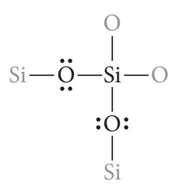
17. Compounds a, b, c, and e would be expected to dissolve in water because they are all polar molecular shapes (single polar covalent bond, trigonal pyramid, asymmetrical tetrahedron, bent, respectively); d is a trigonal plane and is non-polar.

18. Na_2O, MgO, Al_2O_3, $(SiO_2)_n$, P_4O_{10}, P_4O_6, SO_3, SO_2, Cl_2O_7, Cl_2O: Other compounds exist but these are the most important.

Na_2O, MgO, Al_2O_3 are ionic

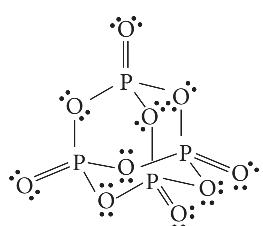


$(SiO_2)_n$ Silicon dioxide is a network solid. The shaded symbols in the Lewis diagram show the atoms in the next unit to which the atoms in the central unit are bonded.

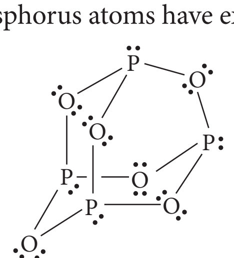


P_4O_{10}, P_4O_6, SO_3, SO_2, Cl_2O_7, Cl_2O are molecular compounds.

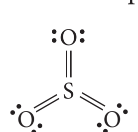
P_4O_{10}: The phosphorus atoms have expanded valences.



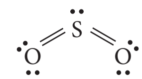
P_4O_6: The phosphorus atoms have expanded valences.



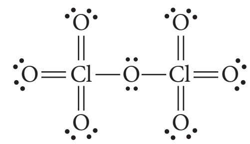
SO_3: The sulfur atom has an expanded valence.



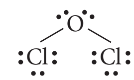
SO_2: The sulfur atom has an expanded valence.



Cl_2O_7: The chlorine atoms have expanded valences.



Cl_2O



19. Electronegativity is a measure of how strongly an atom attracts electrons in a chemical bond. Electron density describes the relative amount of electrons per unit of volume in an atom or in a bond. Bond dipole refers to the development of a partial positive end and a partial negative end in a covalent bond.

20. Because the electronegativity of chlorine is much higher than that of hydrogen, the electron density in the bond shifts toward the chlorine atom to form a bond dipole.

21. **a.** trigonal pyramidal

b. trigonal planar

c. linear

d. square bipyramidal

22. Compounds a, b, and d are polar.

23. The only intermolecular forces acting among the molecules of each of the two compounds are dispersion forces. C_2H_4 is a very small molecule with few electrons. Therefore, it would experience only very small dispersion forces. $C_{20}H_{40}$, on the other hand, is very large and has a large surface area. Dispersion forces could be acting at many points on the molecule at once so it experiences stronger dispersion forces.

24. **a.** Two different salts such as NaCl and $AgNO_3$ were probably combined at $t = 0$. With time, one new salt, such as AgCl, was insoluble and precipitated out of the solution. The other salt ($NaNO_3$) was still soluble and remained in the clear solution.

b. The two clear colourless liquids that were mixed at $t = 0$ were probably a polar liquid, such as water, and a non-polar liquid, such as hexane. The two liquids cannot mix and therefore remain separate.

25. Answers should include cutting with other diamonds and cutting with lasers.

Answers to Unit 2 Review Questions

(Student textbook pages 264-7)

1. b

2. a

3. d

4. e

5. e

6. a

7. a

8. c

9. c

10. b

11. e

12. e

13. a

14. b

15. d

16. *Similarities*: Both models have a positive charge with enough negatively charged particles (electrons) to balance the positive charge.

Differences: The Rutherford model has all of the positive charge and most of the mass concentrated in a very small sphere at the centre of the atom. The Thomson model has the positive charge spread out throughout the entire atom.

17. a. Bohr's model explained the observed line spectra of hydrogen.

b. Bohr's model says that the energy, and therefore the radius of the orbit, of electrons are quantized. Atoms do emit energy in the form of electromagnetic radiation but only in specific quanta which are emitted when an electron drops from one allowed orbit to another allowed orbit.

c. The Bohr model could not explain the more complex spectra of atoms larger than hydrogen.

18. As the atomic number increases, the positive effective charge in the nucleus increases, which increases the attractive force on the electrons, pulling them closer to the nucleus.

19. The aufbau principle states that each electron occupies the lowest energy orbital available. Thus, when filling orbitals in an orbital diagram, place electrons on the lowest energy orbital. When there are several equal-energy orbitals available, follow Hund's rule which states that each equal-energy orbital receives one electron before any pairing occurs. Also, when equal energy orbitals are being filled, the electrons must all have the same spin.

20. a. ultraviolet, X ray, gamma rays

b. violet

c. infrared, microwaves, radio waves

d. The speed of electromagnetic waves in a vacuum is the same (3×10^8 m/s).

e. red

21. Bond polarity forms a continuum, from non-polar to mostly ionic. If the electronegativity difference of the atoms forming the bond is between 0 and 0.4, the bond is mostly covalent. If it is between 0.4 and 1.7, the bond is polar covalent. If it is between 1.7 and 3.3, the bond is mostly ionic.

22. The quantum number l is the orbital shape quantum number. It determines the shape of the orbital such as spherical, dumbbell, four leaf clover, or other shape.

23. Answers should include any three of the following:

i. Electrons exist in circular orbits with the electrostatic force acting as the central force.

ii. Electrons can exist only in allowed orbits.

iii. While in an orbit, electrons do not radiate energy.

iv. Electrons can jump between orbits by absorbing or emitting energy in an amount equal to the energy difference between the orbits.

24. Tempering of steel reduces its brittleness and makes it softer and more malleable. Tempering is done by heating to a temperature well below its melting temperature and letting it cool very slowly.

25. a. solid but malleable and ductile

b. hard with a high melting point but very brittle, aqueous solutions are electrolytes

c. solid and extremely hard, difficult to ignite

26. For an ionic compound to be soluble in water, the attractive forces between the ions and water molecules must be stronger than the attractive forces among the ions themselves.

27. The shared electrons in a covalent bond are electrostatically attracted by the nuclei of both atoms.

28. (electron group arrangement; molecular shape)

a. octahedral; square pyramidal

b. trigonal bipyramidal; seesaw

c. trigonal bipyramidal; linear

d. tetrahedral; trigonal pyramidal

e. octahedral; square planar

f. tetrahedral; tetrahedral

29. If no attractive forces existed among non-polar molecules, they would be gases under any conditions of temperature and pressure. However, all compounds will eventually become liquid and even solid if the temperature is low enough. Therefore, some type of force must be holding the molecules together.

30. a. Bader: developed AIM theory (atoms in molecules); links quantum mechanics to the atoms and bonds in a molecule

b. LeRoy: developed the "LeRoy radius," which is related to the distance between bonded atoms that are about to split into individual atoms

c. Gillespie: developed the concepts of VSEPR

31. a. B: [He]$2s^2 2p^1$; B^{3+}: $1s^2$ does not form

b. Mg: [Ne]$3s^2$; Mg^{2+}: [He]$2s^2 2p^6$

c. S: [Ne]$3s^2 3p^4$; S^{2-}: [Ne]$3s^2 3p^6$

d. K: [Ar]$4s^1$; K^+: [Ne]$s^2 3p^6$

e. H: $1s^1$; H^+: no electrons

32. i. c

ii. a

iii. b

iv. c

v. b

33. a. boron

b. fluorine

c. manganese

d. calcium

e. hydrogen

f. argon

34. carbon

35. iodine

36. a. Group 1

b. i. [Uuo 118]$8s^1$

ii. $n = 8, l = 0, m_l = 0, m_s = +\frac{1}{2}$ or $-\frac{1}{2}$

c. very soft solid or a liquid

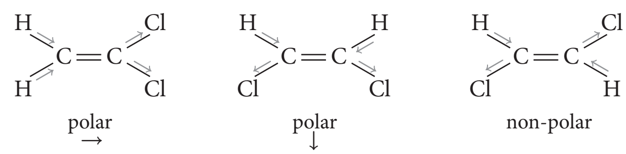
d. reacts explosively with water

37. The fluoride ion is much smaller than the iodide ion; thus the ions pack much closer together in NaF than in NaI. Therefore, the attractive forces among the sodium ions and fluoride ions in the crystals are greater than the attractive forces between the sodium ions and iodide ions. The weaker attractions in the NaI crystals require less energy to break; thus NaI melts at a lower temperature than NaF crystals, which have stronger attractions.

38. The attractions among the sodium ions and iodide ions are weaker than those among sodium and iodide ions for reasons given in the answer to question 37. Therefore, it is easier for the dipoles of the water molecules to pull apart the ions in the NaI crystals than in the NaF crystals.

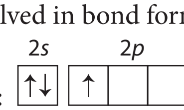
39. The cesium ion is much larger than the sodium ion and therefore the ions in NaF pack much closer together than do the ions in CsF. As a result, the attractive forces among the sodium and fluoride ions are greater than those between cesium and fluoride ions. It is the strength of the attractive forces that influences the melting points and solubilities of ionic compounds. If the attractive forces differ then you would expect that the melting points and solubilities would differ.

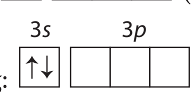
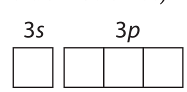
40. **A** **B** **C**

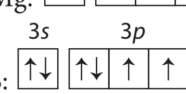
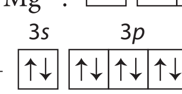


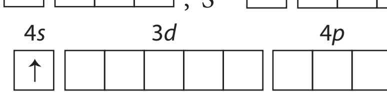
The three isomers of $C_2H_2Cl_2$ are shown above. The shape around all of the carbon atoms is trigonal planar, therefore, the molecules themselves are planar. The grey/light arrows show the direction of the polarity of each individual bond. Isomers A and B are polar with the net direction of the polarity shown by the arrows below the labels. Isomer C is non-polar because the directions of the individual polarities of the bonds cancel each other.

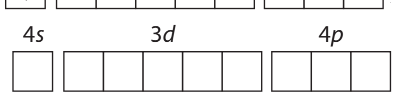
41. Defining the valence electrons as those that are involved in bond formation:

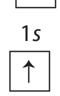
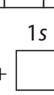
a. B:  (boron ions do not exist)

b. Mg:  ; Mg^{2+}: 

c. S:  ; S^{2-} 

d. K:  ;

K^+ 

e. H:  ; H^+ 

42. a. nitrogen (1), oxygen (2), fluorine (3), neon (4)

b. beryllium (1), boron (1), carbon (1), nitrogen (1), oxygen (2), fluorine (3), neon (4)

43. a. polar covalent; $\Delta EN = 1.3$

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

c. polar covalent; $\Delta EN = 1.4$

d. mostly ionic; $\Delta EN = 3.1$

44. a. non-polar molecular solid. Weak dispersion forces are the only attractive forces between molecules

b. *sample answer*: lard

45.

	bonding pairs	lone pairs	electron group	molecular shape
a. SnF_4	4	0	tetrahedral	tetrahedral
b. H_3O^+	3	1	tetrahedral	trigonal pyramidal
c. AsF_5	5	0	trigonal bipyramidal	trigonal bipyramidal
d. ICl_2^-	2	3	trigonal bipyramidal	linear

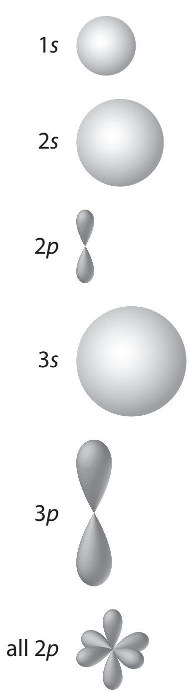
46. a. H_2O: H_2O forms hydrogen bonds whereas H_2S does not. Therefore, the intermolecular forces holding H_2O molecules are stronger.

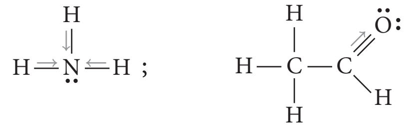
b. $SiCl_4$: Sulfur tetrachloride is unstable and decomposes before it can boil.

c. C_8H_{18}: Both molecules are non-polar but octane is larger and will thus experience greater dispersion forces.

47. Because propanol has an –OH group, it is polar and can form hydrogen bonds. The –C=O bond in propanal is polar but there are no hydrogen atoms bonded to an electronegative atom and thus it cannot form hydrogen bonds. As a result, although the two molecules are very similar in size, you would expect that propanol would have the higher melting and boiling points, which it does.

48.

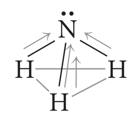


49. a. 

b. Count the number of bonding pairs and lone pairs around the central atom. Total number of groups (BP + LP) determines electron group arrangements. Number of BP and LP determines molecular shape.

NH_3: 3 BP, 1 LP	CH_3CHO: 3 BP, 0 LP
tetrahedral	trigonal planar
trigonal pyramidal	trigonal planar

c. The trigonal pyramidal shape of ammonia orients the polar N–H bonds as shown below. The components of the polarity vectors in the plane of the hydrogen atoms (shown by the gray lines) cancel each other. The components of the polarity vectors that point toward the nitrogen atom add up. The central, vertical vector shows the net polarity of the molecule.



The only bond that is very polar in the ethanal is shown with an arrow in part a. This, alone, causes the molecule to be polar.

50. *All types:* act as covalent bonds between (mostly) non-metal atoms to hold molecules together. All consist of a pair of shared electrons.

Polar covalent bonds and co-ordinate covalent bonds: can be polar

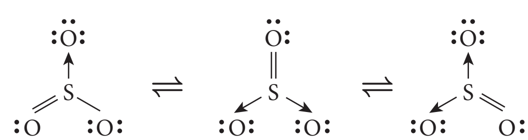
Mostly covalent bonds and co-ordinate covalent bonds: can be non-polar

Polar covalent bonds and mostly covalent bonds: One electron contributed by each atom.

Co-ordinate covalent bonds: Both electrons contributed by same atom.

Polar covalent bonds: Always polar

Mostly covalent bonds: Always non-polar

51. 

52. a. I: linear; II: trigonal planar; III: tetrahedral

b. I: bent; II: trigonal pyramidal; III: seesaw

c. I: bent; II: T shaped; III: square planar

53.

	Ionic	Metallic	Molecular (polar)	Molecular (non-polar)
Unit particle	ion	atom	molecule	molecule
Attraction between particles	electrostatic (positive ion to negative ion)	electrostatic (positive ion to "electron sea")	intermolecular (dipole-dipole and dispersion)	intermolecular (dispersion forces)
Relative Melting Point	very high	very high	intermediate	low
Electrical conductivity	none (only when in solution or molten)	yes	none	none
Hardness/ brittleness	hard, brittle	hard, not brittle	brittle	neither
Example	table salt	copper electrical cord	table sugar	wax

54. Discussions should include the following points.

The forces that hold gecko feet to a surface are all dispersion forces. The forces are very strong because of the large surface area created by many, very small, hair-like projections called setae. Each seta has about 1000 nano-scale hairs which create a huge surface area.

Several research teams have made materials based on the gecko foot that can hold large amounts of mass. For example, one team has made a pad about the size of an index card that can attach an 18 kg object to a smooth, vertical wall.

55. Answers should describe the basic principles of IR technology, and the training required for a job or profession that uses an infrared technology. The answer should describe how infrared technology is used, and the benefits, risks and costs of using the technology in a particular application. Poster presentations should include text with headings and visuals.

56. a. Use of spectroscopy allows scientists to see the spectral lines emitted by celestial bodies. Knowing the spectra of atoms, scientists can identify the chemical composition of these celestial bodies.

b. English physicist Sir William Ramsay (1852-1916) found helium on Earth in 1895. He found it in a mineral that contained uranium. He analyzed it spectroscopically and discovered that it was identical to the element that Lockyer had discovered on the Sun. Lockyer had named it helium because it was found on the Sun (*helios*).

57. a. $[Kr]5s^24d^{10}5p^4$

b. Atoms form ions by gaining or losing electrons in a way that leaves a filled outer shell that is stable. The electron configuration of tellurium is $[Kr]4d^{10}5s^25p^4$.

- Te can form the ion Te^{2-} by gaining 2 electrons, resulting in the configuration $[Kr]4d^{10}5s^25p^6$. The $5s^2$ and $5p^6$ orbitals are thus filled.
- Te can form Te^{4+} by losing 4 electrons, resulting in the configuration $[Kr]4d^{10}5s^2$. The 5*p* shell is empty and the 5*s* shell is filled.
- Te can form Te^{6+} by losing 6 electrons, resulting in the configuration $[Kr]4d^{10}$. The 5*p* and 5*s* shells are both empty leaving the 4*d* shell filled.

58. The electron in hydrogen experiences only the attractive electrostatic force of the nucleus. Electrons of all other elements experience the additional repulsive electrostatic forces of other electrons in the same atom.

59. i. d
ii. c
iii. b
iv. a
v. b
vi. c
vii. a
viii. d

SiF_4 forms a tetrahedron; XeF_4 forms a square plane

60. a. $[Xe]6s^2$

b. Group: 2; period: 6; block: *s*

61. a. iron

b. Group: 8; period: 4

62. a. oxygen

b. $[He]2s^22p^4$

c. Group: 16; period: 2; block: *p*

63. The phosphorus atom in PCl_5 has an expanded valence that includes *d* electrons. The hybrid orbital is sp^3d. The compound NCl_5 cannot exist because nitrogen, being in Period 2, has no *d* electrons. It cannot, therefore, have an expanded valence.

Answers to Unit 2 Self-Assessment Questions

(Student textbook pages 268-9)

1. b

2. d

3. e

4. b

5. c

6. a

7. e

8. d

9. c

10. d

11. a. i. *Rutherford's planetary model*

ii. It explains the small size of the nucleus. It explains the existence of electrons with negative charges that balance the positive charge in the nucleus.

iii. The model contradicts the classical electromagnetic wave theory that says that a charge moving in a circular path should radiate energy. It does not explain the observed line spectra of atoms.

b. i. *The quantum mechanical, or electron cloud model*

ii. It explains nearly all of the experimentally observed characteristics of atoms.

iii. It is difficult to see how electrons can have wave properties and particle properties at the same time.

c. i. *The Bohr model*

ii. It explains nearly all of the observed characteristics of the hydrogen atom. It explains the line spectra of hydrogen. It explains why orbiting electrons do not radiate energy.

iii. It does not explain the complex spectra of multi-electron atoms.

d. i. *Dalton's model, or the billiard ball model*

ii. It explained the particulate nature of matter. It explained that atoms of different elements had different chemical properties.

iii. It could not explain the existence of the electron as a sub-particle of atoms. It cannot explain the observed spectra of atoms and any of the modern observations about atoms.

e. i. *Thomson's plum pudding model*

ii. It explained the existence of negative charges (electrons) that can be emitted from atoms. It explained the electrically neutral nature of atoms.

iii. It could not explain Rutherford's discovery of the small size of the nucleus. It could not explain the line spectra of atoms.

12. a. A positive nuclear core was the only way to explain the path of the alpha particles that followed angles of more than 90° off of their original path.

b. Most of the positive alpha particles passed through undeflected.

c. Based on Thomson's model, they expected that the alpha particles would pass through the gold foils with only very slight deviations from a straight path.

13. An emission spectrum is a line spectrum with very thin lines of specific colours. An absorption spectrum is a nearly continuous spectrum with very small black lines that indicate that specific wavelengths of light were absorbed from the spectrum.

14. Co

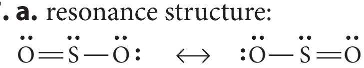
15. The electron configuration shows only the principle (n) and orbital shape (l) quantum numbers. It gives no information about the magnetic (m_l) or spin (m_s) quantum numbers.

16. a. $m_l = -1, 0, +1$

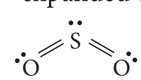
b. 5*p*

c. 3

17. a. resonance structure:



expanded octet:

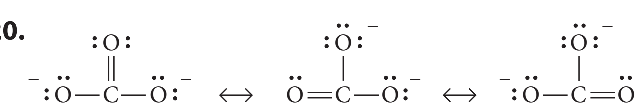


b. The experimentally measured bond lengths of the S-O bonds are shorter than would be predicted by the resonance structure (between single and double bond), but instead, are in agreement with the length of a double bond. Therefore, the expanded octet, or expanded valence, is the most likely structure.

18. Group 13; Period 5; *p* block

19. H—C≡N:

20.



21. Noble gases are inert, indicating that they are very stable. Thus, it would seem that a filled outer shell, or octet, of electrons makes atoms stable. When atoms form bonds, either ionic or covalent, they become more stable than they were as individual atoms. Chemists noted that, when combined, atoms generally have a filled outer shell. These observations led to the octet rule.

22. a. $^{\delta+}C—O^{\delta-}$ not polar $C—S$

b. $^{\delta+}C—F^{\delta-}$ $^{\delta+}C—N^{\delta-}$

c. not polar $P—H$ $^{\delta+}P—Cl^{\delta-}$

23. The students' diagrams should clearly give all of the information that is found in an orbital diagram. The discussion should explain how the diagram and rules fulfil all of the steps in the aufbau principle on page 184 of the student textbook.

24. The students' procedures should include only those tests that have been studied in class or in the textbook. They should be as efficient as possible. The flow diagram should be clear enough that a fellow classmate could follow them successfully. Conductivity of the solid, solubility in water and hardness would be easy steps to use.

25. The first carbon atom (in the methyl group) is *sp*^{3}. The second carbon atom is *sp*. The orbitals in the nitrogen atom are not hybridized.