Chemistry 12

Solutions Manual Part A

Unit 2 Structures and Properties of Matter

Answers to Unit 2 Preparation Questions
(Student textbook pages 158-61)

1. a. Poisonous and infectious material causing immediate and serious toxic effects

b. Use care when handling hot objects

2. Tie back long hair. Do not have any flammable material near the burner. Secure the burner. For example, attach it to a retort stand. Handle hot glassware with thermal gloves or beaker tongs.

3. The sources must be reliable. Web sites such as university and government sites are most reliable. If the source is posted by an individual, the credentials of the author must be shown to be reliable. Corporate sites are often biased.

4. a

5. Note to the teacher: The numbers in the first printing of the text were accidentally combined. The values in the shaded squares are the data that should be provided and the open squares are the answers. BLM 3-1 has been provided with the correct working chart.

Name of Isotope	Notation for Isotope	Atomic Number	Mass number	Number of Protons	Number of Electrons
bromine-81	$^{81}_{35}$Br	35	81	35	35
neon-22	$^{22}_{10}$Ne	10	22	10	10
calcium-44	$^{44}_{20}$Ca	20	44	20	20
silver-107	$^{107}_{47}$Ag	47	107	47	47

6. Valence electrons are those electrons that exist in the outer energy level. They are the only electrons that are involved in chemical bonds.

7. b

8. c

9. The first four electrons, in a Lewis diagram, are drawn as far apart as possible because their negative charges repel each other.
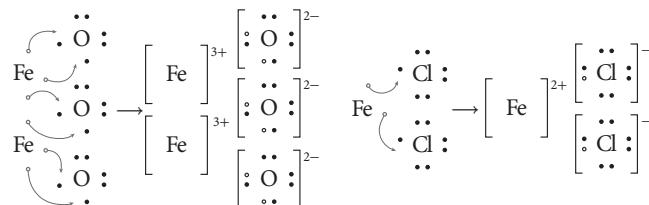
10. a. The distance from the centre of the atom to the boundary within which the electrons spend 90 percent of their time.

b. The energy absorbed or released when an electron is added to a neutral atom.

c. An indicator of the relative ability of an atom to attract shared electrons.

d. The amount of energy required to remove the outermost electron from an atom or ion in the gaseous state.

11.

a. down a group:	b. across a period:
a. increases	a. decreases
b. decreases	b. increases
c. decreases	c. increases
d. decreases	d. increases

12. a. yellow: halogens; very reactive non-metals

b. purple: transition metals; very hard metals with high boiling points

c. brown; other non-metals of the main group elements; most of these atoms are essential building blocks of molecules found in living organisms. Their chemistry typifies how non-metals combine with other atoms.

13. In both cases, the metal is donating, in one case 1 e⁻ and in the other 2 e⁻, to a non-metal. This leaves a positive metal ion and a negative non-metal ion that are attracted to each other by electrostatic forces and they form an ionic compound.

14. The octet rule states that when bonds form between atoms, the atoms gain, lose, or share electrons in a way that will create an outer shell containing eight (an octet) of electrons. However, the electron configurations of the transition elements are such that they can have more than eight electrons in their outer shell.

15.

16. d

17. Both a lone pair and a bonding pair consist of two paired electrons in the valence energy level of an atom. The lone pair is **not** involved in a covalent bond in the molecule but may interact with a nearby molecule forming an intermolecular bond. A bonding pair consists of 1 electron from the valence energy level of two separate atoms forming a covalent bond between two atoms within the molecule.

18. Ionic: a., d., and e.

Molecular: b. and c.

19. a. ammonium ion

b. cyanide ion

c. chromate ion

d. carbonate ion

e. nitrite ion

20. Titanium can form ions having more than one charge. The Roman numeral IV indicates that the ion has lost four electrons and thus carries a charge of 4^+.

21. a. ionic: iron(III) oxide

b. molecular: phosphorus pentaiodide

c. molecular: dinitrogen monoxide

d. ionic: calcium iodide

e. ionic: sodium cyanate

22. a. ionic: $ZnBr_2$

b. molecular: SF_6

c. ionic: $CuCrO_4$

d. molecular: Cl_2O

e. ionic: CsCl

23.

Number	Prefix
1	mono-
2	di-
3	tri-
4	tetra-
5	penta-
6	hexa-
7	hepta-
8	octa-
9	nona-
10	deca-

24. chlorine, methanol, potassium oxide:
Ionic compounds such as potassium oxide form crystals with tightly bound positive and negative ions and thus have very high boiling points. Alcohols such as methanol have hydroxyl groups that can form hydrogen bonds among the molecules. Although hydrogen bonds are not nearly as strong as the bonds between ions, they nevertheless increase the boiling point above those of non-polar compounds. Chlorine forms a non-polar molecule. The only attractive forces among the molecules are very weak and therefore non-polar compounds have very low boiling points.

25. c

26. When ionic compounds are dissolved in water, the ions are separated from each other and are free to move independently. Positive ions are attracted to the negative electrode and negative ions are attracted to the positive electrode. The moving ions constitute an electric current. When the ions reach their respective electrode, the negative ions donate electrons to the positive electrode and the positive ions accept electrons from the negative electrode. The electrons move through the external circuit and form a complete loop of electric charge flow.

Each molecule in a molecular compound has an overall neutral charge and thus they do not move towards any electrode. There is no movement of charges.

27. a. Because the chlorine atoms are more electronegative than carbon atoms, the electrons in the carbon-chlorine bond spend more time near the chlorine atoms thus making the bonds polar.

b. Because the CCl_4 molecule is perfectly symmetrical, the polarity of the bonds cancel each other and the overall molecule is not polar.

c. Because the CCl_4 molecules are non-polar, they would likely not dissolve in water.

28. In a sodium chloride crystal, each sodium ion is attracted to six adjacent chloride ions. The distances between the ions and therefore the strength of the attraction, is the same in every case. The same can be said for the chloride ions, each of which is attracted equally to six adjacent sodium ions. Since all of these attractive forces are the same, it is not possible to identify any specific pair of sodium and chloride ions that form a "molecule."

29. Intermolecular forces are those forces that attract molecules to one another. Covalent bonds are the interactions between atoms in a molecule, holding the individual atoms together to form a molecule.

30. Each elliptical shape represents a polar molecule or a dipole. The oppositely charged ends of the dipoles attract each other with electrostatic forces. This type of attraction is called a dipole-dipole force.

Chapter 3 Atomic Models and Properties of Atoms

Answers to Learning Check Questions
(Student textbook page 168)

1. Thomson's discovery of the electron in 1897 invalidated Dalton's atomic theory.

2. Alpha particles would have passed straight through the foil with minimal or no deflection from encounters or collisions with nearby electrons. There would be no deflection caused by the positive charge because Thomson's model postulates a uniform, positive charge spread throughout the atom.

3. Some radioactive elements emit positively charged alpha particles. Rutherford studied them and then used the alpha particles to bombard thin foils including gold foils. This led to the model in which all of the positive charge and most of the atom's mass were confined to a very small region at the centre of the atom, which Rutherford called the nucleus.

4. Diagrams should be based on Figures 3.3 and 3.6. Both models are spherical and include electrons and the positive charge. In Thomson's model, the positive charge is spread throughout the sphere and electrons are embedded in the sphere like raisins in a muffin. In Rutherford's model, the positive charge is found in a tiny, extremely dense nucleus and the electrons orbit the nucleus like planets.

5. Scientists tend to name their models, or other discoveries, after something that is common to their own everyday lives. Rutherford's model is sometimes called the planetary model.

6. In Thomson's model, negative charges were scattered evenly throughout a large positively charged mass. The alpha particles were highly energetic and would not be expected to be deflected very much by such atoms.

(Student textbook page 183)

7. In a hydrogen atom, orbital energy depends only on n. For example, electrons in $2s$ and $2p$ have the same energy. In multi-electron atoms, orbitals in different sublevels have different energies associated with them, even if they have the same value of n. For example, $2s$ and $2p$ are associated with different energies.

8. a. $2s$, $2p$, $3s$, $3p$
   b. $3p$, $4s$, $3d$, $4p$
   c. $5s$, $4d$, $5p$, $6s$, $4f$, $5d$, $6p$, $5f$

9. An orbital is "full" when it contains two electrons.

10. Method one: There are five possible orbitals for $n = 1$ and $n = 2$: one $1s$ orbital, one $2s$ orbital, and three $2p$ orbitals. Each of these can contain a maximum of two electrons. Therefore, 10 electrons can occupy all possible orbitals with $n = 1$ and $n = 2$. Method two: Using the formula $2n^2$, $n = 1$ can contain two electrons and $n = 2$ can contain eight electrons, for a total of 10.

11. No. Two arrows pointing in the same direction would indicate that two electrons in the same orbital have the same spin quantum number. This violates the statement made in the Pauli exclusion principle that no two electrons can have the same four quantum numbers.

12. a. ☐   **b.** ↑   **c.** ↑↓

(Student textbook page 187)

13. Orbitals fill in order of increasing energy. At energy levels above $n = 3$, the different sublevels overlap. As a result, the $5s$ orbital has a lower energy than the $4d$ orbitals.

14. boron: $1s^2 2s^2 2p^1$; $[\text{He}]2s^2 2p^1$
   neon: $1s^2 2s^2 2p^6$; $[\text{He}]2s^2 2p^6$

15. $1s^2 2s^2 2p^6 3s^2 3p^4$

16.

↑↓	↑	↑	
3s		3p	

17. a. sodium: $1s^2 2s^2 2p^6 3s^1$
   b. vanadium: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

18. titanium: $[\text{Ar}]4s^2 3d^2$

Answers to Caption Questions

Figure 3.3 (Student textbook page 165): These descriptions act as models for real structures that are complex or incapable of being seen. Each part of the model is related to a part of the real structure (for example, raisins represent electrons), so that it is easier to visualize the whole structure and its component parts.

Figure 3.16 (Student textbook page 182): $4s$ is higher since it has the higher principal quantum number (even though it is filled before the $3d$).

Answers to Section 3.1 Review Questions
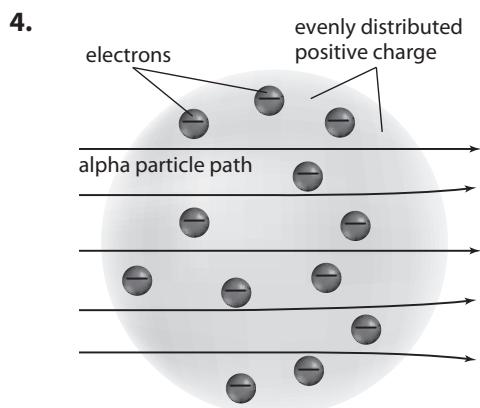
(Student textbook page 173)

1. *Sample answer:* People call Thomson's model the "plum pudding" model because the electrons were embedded in a positively charged sphere like plums in a pudding. Plum pudding was common in Thomson's time but now, many people haven't even heard of it. So we could maybe think of the electrons as blueberries in blueberry muffins.

2. **a.** In a chemical reaction, the mass of products equals the mass of reactants. According to Dalton's theory, atoms are not destroyed or created in chemical reactions—instead they are rearranged. Therefore the mass doesn't change.

   b. Dalton's theory could not explain why atoms combined in certain ratios.

3. The discovery of the neutron solved the problem. The neutron adds mass to the nucleus without adding any charge.

4.

5. In order for the foil to be penetrated by alpha particles, it had to be extremely thin. Because gold is a highly malleable metal, it could be used to make an extremely thin foil.

6. Thomson and Rutherford: atoms contain positive and negative charges, contains electrons; Thomson: electrons distributed throughout sphere, positive charge spread throughout sphere; Rutherford: positive charge in tiny, dense nucleus; electrons orbit the nucleus.

7. Rutherford's electrons are free to orbit anywhere in the volume of space around the nucleus; Bohr's have only specific allowable energy levels. This change provides a model that can explain emission and absorption spectra.

8. These are the only transitions that produce photons with wavelengths in the visible region of the spectrum.

9. **a.** microwave, visible, gamma ray
   Gamma rays have the highest energy and frequency and the shortest wavelength; microwaves have the lowest energy and frequency and the longest wavelength.

   b. infrared, visible, X ray
   X rays have the highest energy and frequency and the shortest wavelength; infrared light has the lowest energy and frequency and the longest wavelength.

   c. radio wave, ultraviolet, X ray
   X rays have the highest energy and frequency and the shortest wavelength; radio-frequency waves have the lowest energy and frequency and the longest wavelength.

10. According to 19th century physics, electrons in Rutherford's model of the atom should emit a complete spectrum of radiation while spiraling into the nucleus. Experimentally, excited atoms emit a line spectrum or a set of very specific wavelengths of light.

11. **a.** $n = 2$

    b. $n = 5$

12. In Dalton's symbols for water and methane, he had the proportions wrong. Water has two hydrogen atoms with one oxygen atom and methane has four hydrogen atoms with one carbon atom.

13. The amount of energy that electrons could have was quantized or has an exact singular value. Electrons in Bohr's model could exist only in allowed energy levels, emitting or absorbing only energy in specific quantities when moving from one energy level to another.
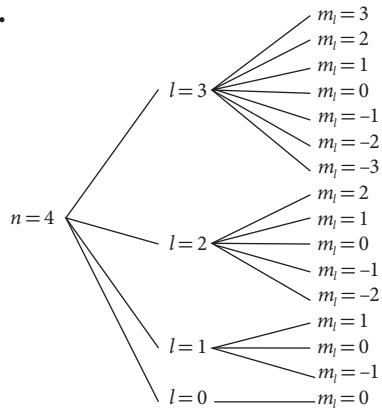
14. One example might be adding sugar to tea using sugar cubes vs. adding sugar from a bag.

Answers to Section 3.2 Review Questions

(Student textbook page 180)

1. In Bohr's atomic model, the *n* represents the allowed orbit which has a specific allowed radius and energy level. In the quantum mechanical model, *n* is the principle quantum number which describes a given energy level or shell.

2.

$n = 4$

$l = 3$
- $m_l = 3$
- $m_l = 2$
- $m_l = 1$
- $m_l = 0$
- $m_l = -1$
- $m_l = -2$
- $m_l = -3$

$l = 2$
- $m_l = 2$
- $m_l = 1$
- $m_l = 0$
- $m_l = -1$
- $m_l = -2$

$l = 1$
- $m_l = 1$
- $m_l = 0$
- $m_l = -1$

$l = 0$ —— $m_l = 0$

3. The fourth quantum number is the spin quantum number symbolized, m_s (m sub s). The spin quantum number has two possible values, $\frac{1}{2}$ and $-\frac{1}{2}$. It represents the orientation of the axis of the electron. The orientations are called spin up and spin down.

4.

Summary of Quantum Number of Electrons in Atoms

Name	Symbol	Allowed Values	Property
Principal (shell)	n	Positive integers (1, 2, 3, etc.)	Orbital size and energy level
Orbital-shape (subshell)	l	Integers from 0 to $(n-1)$	Orbital shape (*l* values 0, 1, 2, and 3 correspond to *s, p, d,* and *f* orbitals)
Magnetic	m_l	Integers from $-l$ to $+l$	Orbital orientation
Spin	m_s	$+\frac{1}{2}$ or $-\frac{1}{2}$	Spin orientation

5. *Sample answer:* The *s* orbital is like a tennis ball. The fuzzy surface represents the concept that the surface is not precisely defined. The *p* orbitals are shaped like putty that is pulled until the two pieces are almost pulled apart. The first four *d* orbitals are shaped like a four-leafed clover and the fifth is like a *p* orbital with a ring around the centre.

6. Disagree: An orbital simply defines a region in the space around a nucleus. There are no physical barriers. Electrons having a specific amount of energy can exist in that region of space or orbital.

7. a. not allowed: When $n = 2$, l can have values of only 0 or 1.

b. not allowed: When $l = 0$, m_l can have a value of only 0. It cannot be -1.

c. allowed

d. not allowed: When $n = 2$, the values of l are restricted to 0 or 1.

8. a. 1

b. 7

c. 3

d. 5

e. 1

f. infinite

9. Six sets:

$n = 3, l = 1, m_l = 0, m_s = \frac{1}{2}$

$n = 3, l = 1, m_l = 0, m_s = -\frac{1}{2}$

$n = 3, l = 1, m_l = 1, m_s = \frac{1}{2}$

$n = 3, l = 1, m_l = 1, m_s = -\frac{1}{2}$

$n = 3, l = 1, m_l = -1, m_s = \frac{1}{2}$

$n = 3, l = 1, m_l = -1, m_s = -\frac{1}{2}$

10. a. The value of n cannot be 0. It has to be at least one more than l. $n = 2, l = 1, m_l = 0$

b. The magnitude of m_l cannot be larger than l. $n = 3, l = 2, m_l = -2$

11. Hi, So you missed class today. Well, orbitals are really weird. They are just a region in space around a nucleus in which an electron having a specific energy can exist. There is really nothing holding the electron there except its energy.

12. For any given energy level, or shell, the *d* orbitals have five different regions in space that electrons can occupy whereas *p* orbitals have only three and *s* orbitals have only one. The value of *l* for *d* orbitals is always 2 whereas for *p* orbitals, $l = 1$, and for *s* orbitals, $l = 0$.

13. a. The statement assumes that all of the electrons in an atom are always in their ground state. However, electrons can absorb energy and move to higher energy levels. The largest *known* atom has 118 electrons. However, new elements might be found or created experimentally in the future.

b. Any number of energy levels is possible. If the atom absorbs the right amount of energy, an electron could be in the $n = 1000$ level.

14. The sphere represents the region in space around the nucleus within which two electrons (spin up and spin down) can exist.

15. The Pauli exclusion principle states that no two electrons in the same atom can have the same four quantum numbers. As a consequence, the number of electrons in the atom determines the number of filled orbitals.

16. Yes: The atom would be in an excited state. That means that the electron would have absorbed energy.

17. The size of the region in space that the electrons in the $n = 2$ orbital can occupy is larger than the region that electrons in the $n = 1$ orbital can occupy. The energy of electrons in the $n = 2$ orbital is greater than the energy of the electrons in the $n = 1$ orbital.

18. a. 2
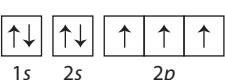
 b. 8

 c. 18

 d. 32

19. For any value of l, m_l can be anything from $-l$ to l. You could think of the 1 in $(2l + 1)$ as being $l = 0$. Then there are l positive values of m_l and l negative values of m_l, making the $2l$.

20. No. Energy levels of all values of n have an $l = 0$ orbital.

Answers to Section 3.3 Review Questions

(Student textbook page 193)

1.

N $1s^2 2s^2 2p^3$ ↑↓ ↑↓ ↑ ↑ ↑
　　　　　　　　　 1s 2s 　2p

Quantum Numbers for Electrons in Ground State

electron	n	l	m_l	m_s
1	1	0	0	$+\frac{1}{2}$
2	1	0	0	$-\frac{1}{2}$
3	2	0	0	$+\frac{1}{2}$
4	2	0	0	$-\frac{1}{2}$
5	2	1	-1	$+\frac{1}{2}$
6	2	1	0	$+\frac{1}{2}$
7	2	1	$+1$	$+\frac{1}{2}$

2. spin $+\frac{1}{2}$

 When filling p orbitals (or any equal energy orbitals) electrons with the same spin occupy each of the orbitals before electrons of opposite spin fill any orbital.

3. a. Oxygen: If the atom has eight electrons and it is neutral, it must have eight protons, which makes it oxygen.

 b. Each of the $2p$ orbitals has one electron with spin $+\frac{1}{2}$. The fourth p electron has opposite spin, thus filling one of the p orbitals.

 c. The Pauli exclusion principle states that only two electrons of opposite spin can occupy any given orbital. In this orbital diagram, no orbital has more than two electrons and the electrons that do occupy the same orbital are of opposite spin.

4. a. In neon, the first and second energy levels are filled. Thus the electron configuration is $1s^2 2s^2 2p^6$. The next energy level to be filled would be $n = 3$. Thus, the electrons following the [Ne] would have to be $3s$ electrons, not $4s$.

 b. Krypton ends the fourth period so electron configurations that begin with [Kr] are in the fifth period. In the fifth period, the $4d$ orbitals begin to fill before the $5p$ orbitals. Therefore $5p^6$ cannot follow $5s^2$.

 c. Bromine does not have all filled orbitals in the outer energy level so it cannot be used as a base for any electron configuration.

5. According to the electron configuration, there are 17 electrons in the outer shell so the element (iodine) is in group 17.

6. Because there are 14 electrons in the outer shell, the element is in group 14. Because the $n = 4$ energy level is being filled, the element is in period 4. Because the p orbitals are the last ones being filled, the element is in the p block.

7. a. The s^1 puts the elements in Group 1, so the elements would be H, Li, Na, K, Rb, Cs, and Fr.

 b. The p^6 puts the elements in Group 18 and the absence of any d electrons in groups above Group 2 eliminates elements in Period 4 and above, so the elements would be Ne and Ar.

 c. The p^4 puts the elements in Group 16 and the presence of 10 d electrons in the $n - 1$ period places the elements in Period 4 and above, so the elements would be Se, Te, and Po.

8. a. potassium: $[Ar]4s^1$

 b. selenium: $[Ar]4s^2 3d^{10} 4p^4$

 c. xenon: $[Kr]5s^2 4d^{10} 5p^6$

9. Cr and Cu are exceptions. When the d orbitals are either exactly half-filled (as in the case of Cr) or completely filled (as in the case of Cu) the configurations are more stable.

10. a. Be, Ca, Ba: Atomic radii increase going down a group.

 b. Cl, Al, Na: Atomic radii increase going left to right across a period.

 c. Sn, Li, Rb: Atomic radii increase going left to right across a period and Sn is the farthest to the right so it is smallest. Atomic radii increase going down a group so Rb is larger than Li.

11. Z represents the actual charge of the nucleus. In atoms of elements with two or more filled inner shells, the negative charges of the electrons slightly cancel the effect on the positive charges of the nucleus on the outer electrons. Z_{eff} is the effective charge of the nucleus on the outer electrons or the charge that the nucleus appears to have with respect to the outer electrons.

12. a. The first ionization energy is the amount of energy needed to remove an electron from a neutral atom. If the first ionization energy is low, that means that it is easy to remove the electron and thus the atom's chemical reactivity is high.

b. The first ionization energy tends to increase across a period from left to right. As you go across a period, the charge on the nucleus becomes greater and the outer electrons are closer to the nucleus, thus it takes more energy to remove an electron. The first ionization energy tends to decrease down a group.

The nuclear charge is increasing down a group and the outer energy level is getting farther away from the nucleus. The force of attraction varies inversely as the square of the distance between the charges. The distance factor has a greater effect and as the radius gets larger down the group ionization energy decreases.

c. As the radius of an atom becomes larger, the ionization energy decreases.

13. Recall that atoms are most stable when orbitals are half-filled or completely filled. For B ($[He]2s^22p^1$) and Al ($[Ne]3s^23p^1$), removal of the only p orbital electron in each will produce more stable filled-orbital electron configurations for these atoms. For O ($[He]2s^22p^4$) and S ($[Ne]3s^23p^4$), removal of one p orbital electron in each will produce more stable half-filled orbital electron configurations for these atoms. Therefore, less energy is needed to remove these electrons.

14. a. To shield the metals from oxygen and moisture in the air that they would react with.

b. *Sample answer*: Work with tiny amounts of Group 1 metals. When it is necessary to work with larger amounts, create a system in which the reaction can be carried out remotely and stay a large distance away.

c. Group 1 metals have only one electron in the outer shell. They have the lowest ionization energy in their respective periods and they become very stable after that electron has been removed. Much more energy is given off when this stable electron configuration is attained than is needed to remove the electron so the reactions are highly exothermic and violent.

Answers to Practice Problems

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

(Student textbook page 179)

1. l can be 0, 1, or 2; for $l = 0$, $m_l = 0$; for $l = 1$, m_l can be -1, 0, or $+1$; for $l = 2$, m_l can be -2, -1, 0, 1, or 2; There are 9 orbitals.

2. $5d$; m_l can be -2, -1, 0, $+1$, or $+2$; There are five orbitals.

3. a. $n = 2$, $l = 0$, $m_l = 0$

b. $n = 3$, $l = 1$, $m_l = -1$, 0, or $+1$

c. $n = 5$, $l = 2$, $m_l = -2$, -1, 0, $+1$, or $+2$

d. $n = 4$, $l = 3$, $m_l = -3$, -2, -1, 0, $+1$, $+2$, or $+3$

4. a. $2s$

b. $5f$

5. a. 1

b. 7

c. 7

d. 3

6. $n = 4$, $l = 2$, $m_l = -2$; $n = 4$, $l = 2$, $m_l = -1$; $n = 4$, $l = 2$, $m_l = 0$; $n = 4$, $l = 2$, $m_l = +1$; $n = 4$, $l = 2$, $m_l = +2$

7. a. $m_l = -1$

b. $l = 3$ Note: The question should read $n = 4$, $l = ?$, $m_l = -3$.

8. $n = 6$, $l = 1$, $m_l = -1$; $n = 6$, $l = 1$, $m_l = 0$ (many more)

9. *Sample answers*:
   a. $l = 2$ does not exist with $n = 1$; change $n = 1$ to $n = 3$

   b. $m_l = -2$ does not exist with $l = 1$; change $l = 1$ to $l = 2$

10. a. Allowed

b. Not allowed; $l = 1$ does not exist with $n = 1$

c. Not allowed; $n = 0$ does not exist

d. Not allowed; m_l can be only -1, 0, or $+1$ when $l = 1$

(Student textbook pages 188-9)

11. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^1$; $[Kr]5s^24d^1$

12. $1s^22s^22p^63s^23p^64s^23d^{10}4p^65s^24d^{10}5p^66s^24f^{14}5d^{10}6p^2$; $[Xe]6s^24f^{14}5d^{10}6p^2$

13. Group 2 elements

14. Group 5; V, Nb, Ta, Db

15. The two exceptions are copper and chromium. The expected diagram for chromium would have a filled $4s$ orbital and four $3d$ orbitals containing one electron each (the fifth $3d$ orbital would be empty). The expected diagram for copper would have a filled $4s$ orbital, four filled $3d$ orbitals, and one $3d$ orbital with one electron.

The actual diagram for chromium has one electron in its $4s$ orbital and one electron in each of its five $3d$ orbitals. The actual diagram for copper has one electron in its $4s$ orbital and five filled $3d$ orbitals. The discrepancy arises because the predicted electron configurations have a slightly higher energy than the actual electron configurations. When the $3d$ orbitals are either all half-filled or all completely filled, the configuration becomes stable. The predicted electron configurations are just 1 electron off this energetically more favourable situation.

16. The configuration for the valence electrons, s^2, indicates that strontium is in Group 2.

17. a. nickel: $[Ar]4s^23d^8$

   b. lanthanum, La: $[Xe]6s^25d^1$

18. The value of 4 in $4s^2$ indicates that titanium is in Period 4.

19. The configuration of the valence electrons shows full s and d orbitals, but a half-filled p orbital, so arsenic must be in the p block.

20. a. The electron configuration of the valence electron is s^1, so francium must be in Group 1 and belong to the s block. Because of the 7 in $7s^1$, francium must be in Period 7.

   b. The electron configuration of the valence electrons is $s^2f^{14}d^4$, so tungsten must be in Group 6 (2 + 4) and belong to the d block. Because of the 6 in $6s^2$, tungsten must be in Period 6.

   c. The electron configuration of the valence electrons is $s^2d^{10}p^3$, so antimony must be in Group 15 (2 + 10 + 3) and belong to the p block. Because of the 5 in $5s^2$, antimony must be in Period 5.

Answers to Chapter 3 Review Questions

(Student textbook pages 199-203)

1. b

2. a

3. c

4. a

5. e

6. d

7. a

8. a

9. b

10. d

11. e

12. a

13. c

14. d

15. In both models atoms were spherical and solid (no empty space) and were neutral-particles. As well, in both models, atoms were considered the building blocks of all matter.

Thomson's model contained internal structures and discrete structures that were smaller than the atom itself. Specifically, Thomson's model had negatively charged subatomic particles, now known as electrons, embedded in a continuous mass of positively charged matter.

Dalton's atom was the smallest possible particle of matter. It was indivisible, meaning that it could not be broken down into smaller particles.

16. Thomson's and Rutherford's models of the atom were spherical and contained a sub-atomic particle believed to be the same particle in all matter (the electron). Both models included equal amounts of positively and negatively charged matter to create an overall neutral atom.

In Rutherford's model, the positively charged matter was contained in a very small region at the centre of the atom which he called the nucleus. Most of the matter of the atom was contained within this nucleus. Negatively charged electrons orbited the nucleus. Most of the atoms consisted of empty space.

Thomson's model suggested that the positive and negatively charged matter was evenly distributed and no empty space existed within the atom.

17. In all atomic models, the atom is the building block of all matter, atoms are neutrally charged, and atoms are spherical.

18. The speed of all electromagnetic radiation, in a vacuum, is the same, 3×10^8 m/s.

19. Bohr's electron moves in a circular pathway a fixed distance from the nucleus—its position is known. The electron in the quantum mechanical model moves constantly within a region of space that is described by using wave equations. Its exact position cannot be known for certain.

20. By using the basic laws of physics and Planck's concepts, Bohr was able to calculate the energy and radii of the allowed orbits for the hydrogen atom.

21. Groups 1 and 2 metals have low electron affinity and low ionization energies and therefore lose electrons easily, forming positive ions. These positive ions are then attracted to negative ions to form ionic compounds.

22. There are three equal-energy *p* orbitals. According to Hund's rule, single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. When the three electrons in the *p* orbitals of nitrogen have occupied the *p* orbitals, there are no additional electrons to pair with the single electrons in the *p* orbitals.

23. The exact position of an electron in an atom cannot be known with certainty. The energy of electrons in any electron shell and the relative size of the orbital are known with certainty as defined by the principal quantum number *n*.

24. Quantum numbers provide information about the energy, size, shape, spatial orientation, and the relative direction of the axis of the electron of the orbital.

25. The principle quantum number provides information about the energy and size of the orbital.

26. The electron in a higher energy level would have more energy to overcome the attractive pull of the positive charge of the nucleus and therefore be able to venture farther away from it.

27. The Pauli exclusion principle states that only two electrons of opposite spin could occupy an orbital. Each electron in an electron pair within the same orbital are identical but have opposite spins—the spin quantum number is the same number for any electron, but the sign is either + (spin up) or − (spin down).

28. Hund's rule states that the lowest energy state for an atom has the maximum number of unpaired electrons allowed by the Pauli exclusion principle in a given energy sublevel. This rule can also be stated as single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. This rule determines the order of the filling of orbitals.

29. An atom's chemical properties are mainly associated with its ground state electron configuration. As well, there are a tremendous number of possible excited states.

30. The atomic radius decreases across a period (from left to right) and increases down a group. The reason that it decreases across a period is because the amount of charge in the nucleus becomes greater across a period. Thus the attractive force on the electrons becomes greater, pulling the electrons closer to the nucleus. The radius increases down a group because each successive group has one more energy level, or shell, than the previous one and thus electrons are farther from the nucleus.

31. None of the experiments described in the book would have remained the same because the tests depend on the charge carried by the particles.

32. The element described, would be in Period 6, *d* block, and Group 12.

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

34. 3*s*, 3*d*, 4*p*, 5*s*, 5*d*

35. The highest possible value of *l* is 3 (because there are electrons in a *d* orbital). Therefore, m_l can be anything from −3 to +3.

36. nickel, Ni

37. aluminum, Al

38. tellurium, Te

39. chlorine, Cl

40. ruthenium, Ru

41. oxygen, O

42. barium, Ba

43. In these cases, the *d* electrons do not typically participate in chemical reactions. As well, all *p*-block elements in Period 4 and above have 10 *d* electrons.

44. Boron will be larger than fluorine.

45. Magnesium will be larger than silicon.

46. Calcium will be larger than selenium.

47. a. Electron configuration: $[Rn]7s^2 5f^{14} 6d^{10} 7p^6$

   b. Quantum number for last electron: $n = 7$, $l = 1$, $m_l = 1$, $m_s = -\frac{1}{2}$

   c. Physical state at room temperature: gas

   d. Reactivity: inert, just like all other noble gases

48. a. Na, Si, Ar: First ionization energy tends to increase across a period. Sodium, silicon, and argon are all in Period 3.

b. Ba, Ca, Mg: First ionization energy tends to decrease down a group. Barium, calcium, and magnesium are all in Group 2.

c. Li, Be, He: Helium would have the highest first ionization energy because it is the farthest right in Period 1. Lithium would have the lowest first ionization energy because it is to the left of beryllium in Period 2.

49. a. Oxygen: There are eight electrons in the neutral atom, thus there must be eight protons in the nucleus. Oxygen has the atomic number eight.

b. The electron configuration indicates the values of the quantum numbers n and l.

c. The electron configuration does not provide information about the magnetic and spin quantum numbers, m_s and m_l.

50. Distribute BLM G-23 Risk/Benefit Analysis to assist students in organizing their thinking. Answers may be in the form of a chart, a Venn diagram, or risk/benefit analysis and should include a clear and succinct description of the technology and its use, plus a brief summary of each cost/risk and benefit. Sources of information should also be included.

51. Answers should follow the outline of the question and include descriptions of:

- how X rays can be released in cathode ray tubes
- the precautions that are taken to prevent the unguarded release of X ray radiation from the CRT computer monitor
- any other precautions or warnings that are necessary that extend from the use of a cathode ray tube in a computer monitor.

Ideas that could be further developed include:

- discovery of X rays and early use as an experimental tool
- construction of first CRT, how X rays are produced
- work of J.J. Thomson, Henry Moseley and others
- parts of CRT used in television picture tubes and oscilloscopes
- dangers for those working with CRT, precautions to take
- disposal problems, environmental concerns

Sources of research information should be cited.

52. Venn Diagram:

- *Orbit only:* 2-D pathway; circular shape; know exactly the location of the electron (orbit has a fixed radius from the nucleus); $2n^2$ electrons can fit in an orbit; classical mechanics; also known as energy level shell.
- *Both orbit and orbital:* electron realm
- *Orbital only:* 3-D space; many shapes; do not know the exact location of the electron—a high probability that it is anywhere in a certain space; only 2 electrons can fit in an orbital; quantum mechanics; also known as: sublevel of energy; subshell; wave function.

53. Students might suggest such examples as twisting balloons together, balloons of one colour within balloons of a different colour, etc.

54. The condensed electron configuration for bromine is $[Ar]4s^23d^{10}4p^5$.
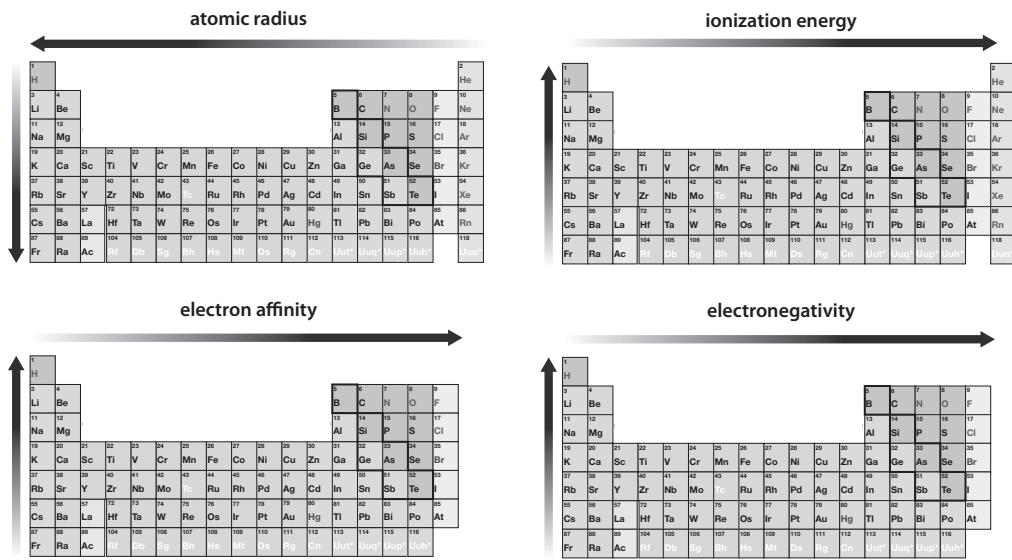
The [Ar] represents the electron configuration for argon, which would be $1s^22s^22p^63s^23p^6$. If you put this electron configuration in place of the [Ar], you would have the complete electron configuration for bromine. Because all of the occupied energy levels in argon and all other noble gases are completely filled, these electron configurations are present in the electron configuration of the elements in the next period. Therefore, for any given period, you can indicate the electron configuration of all of the filled energy levels by writing the symbol for the noble gas in square brackets.

55. The effective nuclear charge is the net positive charge that an electron experiences from the nucleus. It is equal to the nuclear charge less the shielding or screening from the inner core of electrons.

Comparing a Mg atom (2,8,2) with a S atom, (2,8,6), the outer valence electron of Mg experiences a nuclear charge of 12+ and a shielding from 8 inner electrons. On the other hand, the valence electron in an atom of S experiences a nuclear charge of 16+ and a shielding of 10 inner electrons. The effective nuclear charge is greater in S and it is a smaller atom than Mg.

Estimates have been made of the shielding effect contributed by electrons in different orbitals that allow for a comparison of effective nuclear charge to be made between different types of atoms. These estimates reasonably agree with the differences in atomic radius between different types of atoms.

56.

atomic radius

ionization energy

electron affinity

electronegativity

57.

Scientist	Definition/Description (Formal)	Key Concept for the quantum mechanical model
Planck	suggested that matter, at the atomic level, can absorb or emit only discrete quantities of energy. Each of these specific quantities is called a quantum of energy.	the energy of an atom was quantized, meaning that it can exist only in certain discrete amounts.
Einstein	asserted that light was *also* quantized. This meant that light occurs as quanta of electromagnetic energy having particle-like properties. These particle-like "packets" of energy were later called photons.	photons of light are emitted in gas discharge tubes to create discrete lines in atomic spectra
Heisenberg	Using mathematics, Heisenberg showed that it is impossible to know both the position and the momentum of an object beyond a certain measure of precision. According to this principle, if you can know an electron's precise position and path around the nucleus, as you would by defining its orbit, you cannot know with certainty its velocity. Similarly, if you know its precise velocity, you cannot know with certainty its position.	Based on the uncertainty principle, Bohr's atomic model is flawed because you cannot assign fixed paths (orbits) to the motion of electrons.
de Broglie	as energy has matter-like properties (a photon of light), matter can have wave-like properties	developed an equation that enabled him to calculate the wavelength associated with any object — large, small, or microscopic.
Schrödinger	used mathematics and statistics to combine de Broglie's idea of matter waves and Einstein's idea of quantized energy particles (photons). Schrödinger's mathematical equations and their interpretations, together with another idea called Heisenberg's uncertainty principle, resulted in the birth of the field of *quantum mechanics*. This is a branch of physics that uses mathematical equations to describe the wave properties of sub-microscopic particles such as electrons, atoms, and molecules. Schrödinger used concepts from quantum mechanics to propose a new atomic model: the quantum mechanical model of the atom. This model describes atoms as having certain allowed quantities of energy because of the wave-like properties of their electrons.	Schrödinger used a type of equation called a wave equation to define the probability of finding an atom's electrons at a particular point within the atom. There are many solutions to this wave equation, and each solution represents a particular wave function. Each wave function gives information about an electron's energy and location within an atom.

58. 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 3 Summary on page 252 of the student textbook. Concepts relate to types of bonds.

59. a. The chemist would approach atomic theory from an experimental point of view, building a model to explain what is observed in an experiment. Mathematicians and physicists would look at the theoretical possibilities for an explanation; the mathematician by trying to quantify results in the form of mathematical equations and the physicist by looking at how the results can be explained by the fundamental laws of the universe.

  b. The physicist would specifically look at the laws that describe the interactive forces between particles.

60. The Thomson model of the atom suggests that the atom is made up of matter that has an overall positive charge in which the negatively charged electrons are embedded for an overall neutral-charge atom—there is no discrete particle similar to the electron that contains a certain positive charge. Therefore, the Thomson model has a continuum of positive matter inside it. The Rutherford model claims that all of the positive charge of a nucleus is located in a discrete nuclear core and the negative charges of the atom are found in the discrete particles called electrons moving outside the nucleus. Because the different types of matter are found in discrete structures, the matter can be referred to as "quantized."

61. a. • all the positive charge was concentrated in a central nucleus

  • the positive charge was electrically balanced by negative electrons

  • most of the atom was empty space

  b. These features of the atom challenged known properties of matter. Putting all the positive charge into a small volume did not make sense because positive charges repel each other more strongly as they get closer and closer together. It was difficult to imagine that the matter that can be seen and felt is mainly empty space. Rutherford proposed that the negatively charged electrons are orbiting the positively charged nucleus. Accepted concepts of physics indicated that moving charged particles (in orbits or otherwise) should emit energy and therefore lose energy—in the case of electrons, this loss of energy would cause them to move closer and closer to the nucleus until they crashed into it.

62. Bohr's model was very successful at predicting, matching, and explaining the spectral lines for hydrogen. However, it was very unsuccessful at predicting, matching, and explaining the spectral lines (especially clustering) in all multi-electron elements.

63. Answers should identify the health-care professional(s) interviewed and include answers to all parts of the question. The answer to question 3 should include some detail to explain it.

64. The electromagnetic spectrum of radiation is a continuum of energy that can be thought of in two ways: a wave of energy in which the frequency and wavelength are inversely related or as packets or photons of energy that are directed related to the frequency of the radiation. Ultraviolet light is one part of this continuum that lies just beyond the violet portion of the visible spectrum. This UV energy can excite electrons in some atoms, which, when they fall back to a lower energy level, emit visible violet light.

Canada's new currency has molecules embedded in it that will react in this way under UV radiation. The receptors in the newer bills specifically respond to a narrow band of UV radiation (wavelength of 370 nm) and emit a very faint white-purple light. This specificity of receptors to respond to a single wavelength makes it difficult to duplicate and therefore the bills are less likely to be counterfeited.

Older Canadian banknotes included dots that glowed blue under ultraviolet light. (These notes are being phased out as polymer notes are being introduced.)

65.

Firework Colours

Colour	Compound
Red	strontium salts, lithium salts
Orange	calcium salts
Green	barium salts, $BaCl_2$
Blue	copper compounds, $CuCl_2$
Purple	mixture of strontium (red) and copper (blue) compounds

Source:
http://chemistry.about.com/od/fireworkspyrotechnics/a/fireworkcolors.htm

66.

| Scenario	Example	
	Continuous Version	**Quantized Version**
Climbing	Using a ramp	Using a ladder
Paying an amount of money	(sample) using debit	(sample) using cash (bills and coins)
Getting down a snowy hill	(sample) skiing	(sample) using snowshoes

67.

H	He	Li	Be	B																									C
N	O	F	Ne	Na	Mg															Al	Si	P	S	Cl	Ar	K	Ca	Sc	Ti
V	Cr	Mn	Fe	Co	Ni															Cu	Zn	Ga	Ge	As	Se	Br	Kr	Rb	Sr
Y	Zr	Nb	Mo	Tc	Ru	Rh	Pd	Ag	Cd	In	Sn	Sb	Te	I	Xe	Cs	Ba	La	Ce	Pr	Nd	Pm	Sm	Eu	Gd	Tb	Dy	Ho	Er
Tm	Yb	Lu	Hf	Ta	W	Re	Os	Ir	Pt	Au	Hg	Tl	Pb	Bi	Po	At	Rn	Fr	Ra	Ac	Th	Pa	U						

68. carbon, titanium, strontium, and erbium
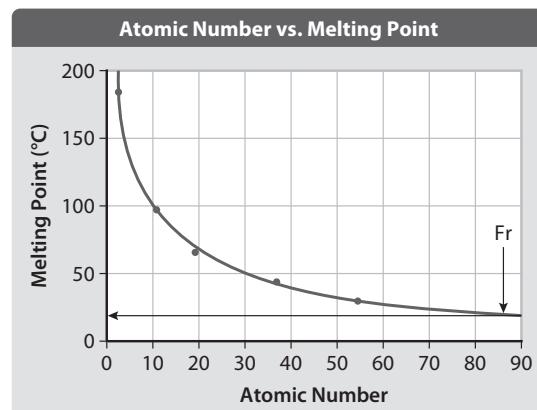
69. [Uuo]$8s^2 5g^7$

70. 1. For main group elements, the last numeral of the group number is the same as the number of valence electrons.

   2. The n value of the highest occupied energy level is the period number in which the element is found.

   3. The number of orbitals in an energy level is equal to n^2, and the maximum number of electrons allowed in an energy level is equal to $2n^2$.

   4. Each block of the periodic table contains twice as many groups as there are orientations of that type of orbital—this corresponds to the maximum number of electrons that can fill all the same type of orbitals within an energy level (2×1 groups in the s block, 2×3 groups in the p block, 2×5 groups in the d block, and 2×7 groups in the f block).

71. Carbon contains no d electrons and lead does. The presence of d electrons makes the electron configuration very complex. In the case of lead, the valence electrons form a "sea of electrons" that causes metallic bonding.

72. a.

Atomic Number vs. Melting Point

b. Li to Na: $\Delta T = 83°C$

  Na to K: $\Delta T = 34°C$

  K to Rb: $\Delta T = 25°C$

  Rb to Cs: $\Delta T = 11°C$

c. The arrow labelled Fr shows where francium would be on the graph. The horizontal arrow shows what the melting temperature probably is. The melting point of francium should be approximately 20°C.

d. Francium will probably be liquid at 25°C.

73. Answers should include a brief biography of the chosen subject, including birth, education, and career summary of work and/or achievements, as well as their work in quantum mechanics and quantum chemistry.

Answers to Chapter 3 Self-Assessment Questions

(Student textbook pages 204-5)

1. c

2. a

3. d

4. e

5. d

6. a

7. e

8. d

9. c

10. a

11. a. When using a cathode ray tube and generating a potential difference across the electrodes that were in a partial vacuum, a particle was emitted from the negative electrode that was drawn to the positive electrode.

   b. When the experimenter placed a magnet near the tube, the 'ray' was bent in the direction in which negative particles were known to move.

   c. The same results were obtained regardless of which metal was used for the electrodes.

12. *Nuclear model*: When Rutherford discovered that all of the positive charge and most of the mass was located in a very small region at the centre of the atom, he called it the nucleus.
Planetary model: According to Rutherford's model, the electrons orbited the nucleus much like planets orbit the Sun.
Beehive model: Some people thought that the electrons moving around the nucleus were like bees flying around a hive.

13. A continuous spectrum includes all wavelengths whereas a line spectrum contains several specific wavelengths of electromagnetic waves.

14. The symbol n generally represents an energy level in both the Bohr and quantum mechanical models. In the Bohr model, n refers to a specific allowed energy and orbit with a specific radius. In the quantum mechanical model, n refers to an energy level or shell which has sub-levels or orbitals.

15. $n = 7; n = 5; n = 4; n = 2; n = 1$: Higher numbers of n are farther from the nucleus and therefore more energy is required to pull the electron away from the nucleus.

16.

Energy level	Subshell	m_ℓ	State	n	ℓ	m_ℓ	m_s	Total number of states
$n=1$	$\ell=0$	$m_\ell=0$	$m_s=+\frac{1}{2}$	1	0	0	$+\frac{1}{2}$	2
			$m_s=-\frac{1}{2}$	1	0	0	$-\frac{1}{2}$	
$n=2$	$\ell=1$	$m_\ell=1$	$m_s=+\frac{1}{2}$	2	1	1	$+\frac{1}{2}$	8
			$m_s=-\frac{1}{2}$	2	1	1	$-\frac{1}{2}$	
		$m_\ell=0$	$m_s=+\frac{1}{2}$	2	1	0	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	2	1	0	$-\frac{1}{2}$	
		$m_\ell=-1$	$m_s=+\frac{1}{2}$	2	1	-1	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	2	1	-1	$-\frac{1}{2}$	
	$\ell=0$	$m_\ell=0$	$m_s=+\frac{1}{2}$	2	0	0	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	2	0	0	$-\frac{1}{2}$	
$n=3$	$\ell=2$	$m_\ell=2$	$m_s=+\frac{1}{2}$	3	2	2	$+\frac{1}{2}$	18
			$m_s=-\frac{1}{2}$	3	2	2	$-\frac{1}{2}$	
		$m_\ell=1$	$m_s=+\frac{1}{2}$	3	2	1	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	2	1	$-\frac{1}{2}$	
		$m_\ell=0$	$m_s=+\frac{1}{2}$	3	2	0	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	2	0	$-\frac{1}{2}$	
		$m_\ell=-1$	$m_s=+\frac{1}{2}$	3	2	-1	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	2	-1	$-\frac{1}{2}$	
		$m_\ell=-2$	$m_s=+\frac{1}{2}$	3	2	-2	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	2	-2	$-\frac{1}{2}$	
	$\ell=1$	$m_\ell=1$	$m_s=+\frac{1}{2}$	3	1	1	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	1	1	$-\frac{1}{2}$	
		$m_\ell=0$	$m_s=+\frac{1}{2}$	3	1	0	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	1	0	$-\frac{1}{2}$	
		$m_\ell=-1$	$m_s=+\frac{1}{2}$	3	1	-1	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	1	-1	$-\frac{1}{2}$	
	$\ell=0$	$m_\ell=0$	$m_s=+\frac{1}{2}$	3	0	0	$+\frac{1}{2}$	
			$m_s=-\frac{1}{2}$	3	0	0	$-\frac{1}{2}$	

State / Total number of states table (quantum states for $n=4$):

m_s	n	ℓ	m_ℓ	m_s
$m_s = +\frac{1}{2}$	4	3	3	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	3	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	3	2	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	2	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	3	1	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	1	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	3	0	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	0	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	3	–1	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	–1	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	3	–2	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	–2	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	3	–3	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	3	–3	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	2	2	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	2	2	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	2	1	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	2	1	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	2	0	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	2	0	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	2	–1	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	2	–1	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	2	–2	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	2	–2	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	1	1	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	1	1	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	1	0	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	1	0	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	1	–1	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	1	–1	$-\frac{1}{2}$
$m_s = +\frac{1}{2}$	4	0	0	$+\frac{1}{2}$
$m_s = -\frac{1}{2}$	4	0	0	$-\frac{1}{2}$

$n = 4$; branches: $\ell = 3$, $m_\ell = 3, 2, 1, 0, -1, -2, -3$; $\ell = 2$, $m_\ell = 2, 1, 0, -1, -2$; $\ell = 1$, $m_\ell = 1, 0, -1$; $\ell = 0$, $m_\ell = 0$. Total number of states: **32**

17. When light interacts with matter, it does so in discrete quantities of energy, or photons, in an all-or-nothing manner. However, when it is travelling, it exhibits wavelike properties.

Electrons have a specific amount of mass but when they travel, they exhibit wavelike properties.

18. a. m_l can be any integer from –2 to 2; five d orbitals

b. m_l can be –1, 0, or +1; three p orbitals

c. m_l must be 0; one s orbital

19. 1. Always fill the lowest energy level and sublevel of energy first before filling ones with higher energies.

2. Place only two electrons maximum in each orbital.

3. Fill each equal-energy orbital (orbitals with the same value of l) with one electron only, to half fill all of them, before adding a second electron in each one.

20. rhodium, Rh (Note: configuration should read $[Kr]5s^1 4d^8$)

21. Group 14; Period 3; p block

22. a. main group elements

b. transition metals or transition elements

c. inner transition metals or inner transition elements

23. • *Bohr model only*: quantized radii for electrons; electrons move in circular orbits; no energy is emitted during orbiting; electrons can absorb energy difference between two energy levels and move to a higher energy level; could explain line spectra of hydrogen atom

• *Rutherford model only*: electrons orbit randomly in circular orbits; rotating electron is predicted to emit energy and electron should lose energy and be drawn to the nucleus; could not explain line spectra of any atom

• *Quantum mechanical model only*: electrons have wave-like and matter-like properties; quantum numbers define the characteristics of electrons in region of space called an orbital; exact position and energy of an electron cannot be known simultaneously; explains line spectra of some multi-electron atoms.

• *Bohr and Quantum model*: energy of electron is quantized; most probable position of electron is the same as Bohr's exact prediction of position; explain atomic spectra of hydrogen atom.

• *Bohr and Rutherford models*: electrons orbit the nucleus

• *Rutherford and Quantum:* nil

• *Rutherford and Bohr and Quantum*: Negative charge surrounds a central positive core; atom is a neutral species

24. The first ionization energy tends to increase across a period and decrease down a group.

25. electronegativity and electron affinity

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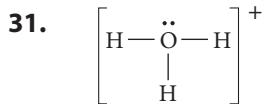
23. Like a hybrid offspring resulting from crossing two species, a hybrid orbital results from the combination of atomic orbitals of the valence electrons. For example, a sp^2 hybrid orbital results from combining one *s* atomic orbital and two *p* atomic orbitals.

24. H—C≡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.

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.

31.

$$\left[H - \overset{\cdot\cdot}{\underset{|}{O}} - H \right]^{+}$$
$$\qquad\quad H$$

Oxygen is sp^3 hybridized from mixing one $2s$ orbital with three $2p$ 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 $4s$ orbital, three $4p$ orbitals, and two $4d$ 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 $3s$ orbital, three $3p$ orbitals, and one $3d$ 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
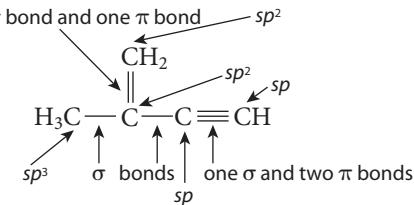
11. The pattern in which the carbon atoms are bonded together is very different in diamond and graphite. In diamond, each carbon atom in the crystal is at the centre of a tetrahedron with another carbon atom at each corner, giving a three-dimensional network solid. Diamond is very hard with several planes 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
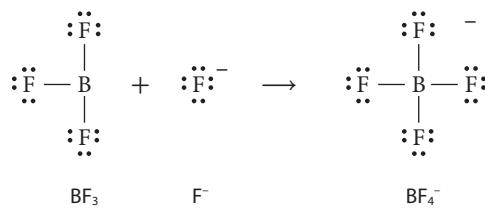
13.

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
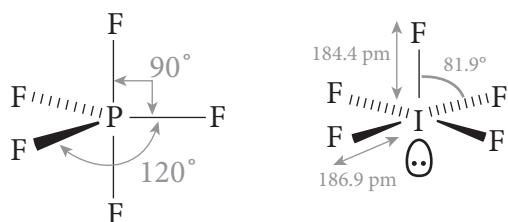
2. *Sample answer*: Two or more possible Lewis structures of the same molecule in which the structures differ only in the positioning of their bonding and lone pairs. The actual structure of the molecule is most likely a combination, or hybrid, of all possible Lewis structures. They are used when no 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.)
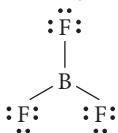
3. To determine the shape of a molecule, you need to know the total number of electron groupings (bonding pairs and lone pairs) around the central atom. A Lewis structure is needed in order to determine the number of electron groupings.

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.
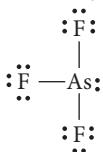
6. a. BF_3
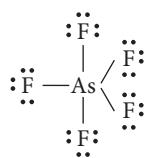
Boron trifluoride has three electron groupings around the boron atom so it is trigonal planar.

b. AsF_3
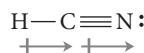
Arsenic trifluoride has four electron groupings (three bonding pairs and one lone pair) so the electron pairs have a tetrahedral arrangement, but the shape is trigonal pyramidal.

c. AsF_5
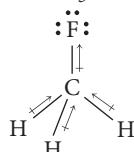
Arsenic pentafluoride has five electron groupings so it is trigonal bipyramidal.

7. a. HCN
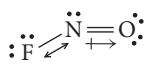
The shape is linear and the molecule is polar.

b. CH_3F
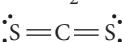
The shape is tetrahedral and the molecule is polar.

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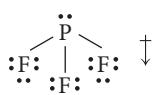
8. I. Check for solubility in water. The polar compound should dissolve but the non-polar compound should not dissolve.

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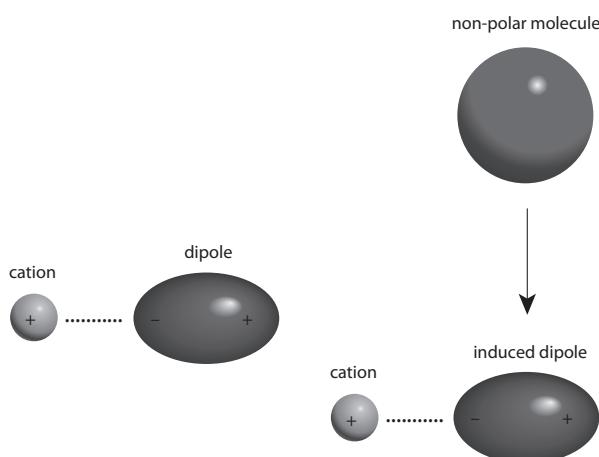
b. There are four electron groupings with one being a lone pair. Therefore, the electron group arrangement is tetrahedral.

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.

non-polar molecule

dipole

cation

cation

induced dipole
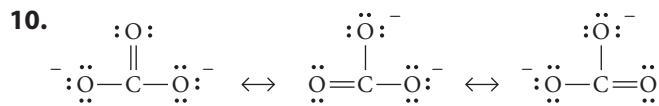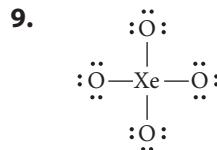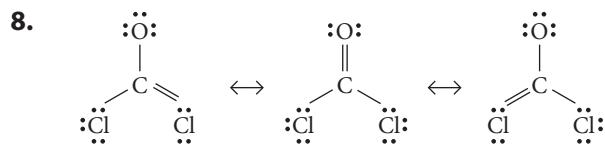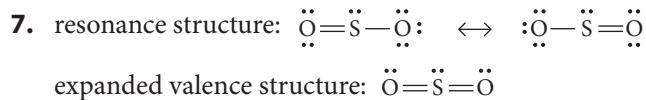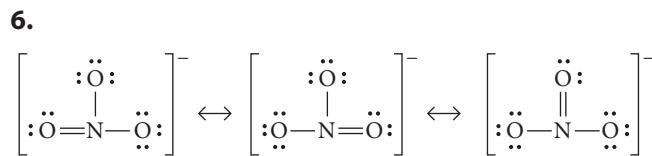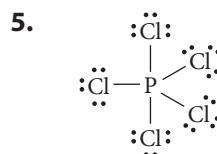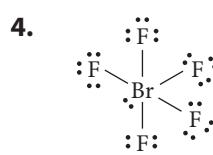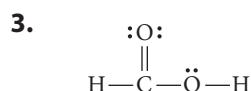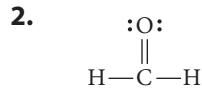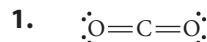
13. a. dipole-dipole forces

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. $\ddot{O}=C=\ddot{O}$

2.
$$H-\overset{\overset{\ddot{O}}{\|}}{C}-H$$

3.
$$H-\overset{\overset{\ddot{O}}{\|}}{C}-\ddot{O}-H$$

4.

5.

6.
$$\left[\right]^- \leftrightarrow \left[\right]^- \leftrightarrow \left[\right]^-$$

7. resonance structure: $\ddot{O}=\ddot{S}-\ddot{O} \leftrightarrow \ddot{O}-\ddot{S}=\ddot{O}$

expanded valence structure: $\ddot{O}=\ddot{S}=\ddot{O}$

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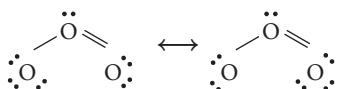
 b. ionic solids have a high melting point and molecular solids do not, ionic solids are much harder than polar molecular solids

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. a. Ca:

 Ca^{2+}

 b. N: ; N^{3-}:

 c. Al: ; Al^{3+}:

 d. Ne: ; no ion

 e. Be: ; Be^{2+}:

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: $[Ne]3s^23p^6$

 c. chlorine: $[Ne]3s^23p^5$

 d. phosphorus: $[Ne]3s^23p^3$

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)
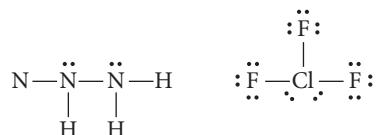
 d. polar

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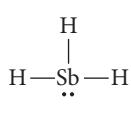
c. $SeCl_4$; $SeCl_4$ has a seesaw shape and is therefore polar allowing dipole-dipole forces between the molecules. $SiCl_4$ is tetrahedral and thus symmetrical, making it non-polar

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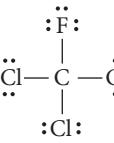
c. heptanes; Molecules of both compounds are identical in size and therefore have the same number of electrons. They are both non-polar so dispersion forces are the only intermolecular forces. Heptane is linear whereas 2,4-dimethyl pentane is spherical. Therefore, Heptane molecules can have larger surface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

45. a.

H—Sb—H

3 BP; 1 LP

b.

:Cl—C—Cl:
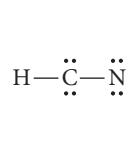
4 BP; 12 LP
(total in molecule)

c.

H—C—N

2 BP; 2 LP

d. H : C ::: C : H

5 BP, 0 LP

e. :F̈—Be—F̈:

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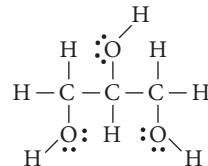
Metallic: conducts electric current in the solid state, does not crush in a mortar but rather deforms (malleable, not brittle), spreads to form a sheet when pounded with a hammer, lustrous, grey, may or may not melt in a deflagrating spoon over a Bunsen burner flame, but can deform more easily in a mortar

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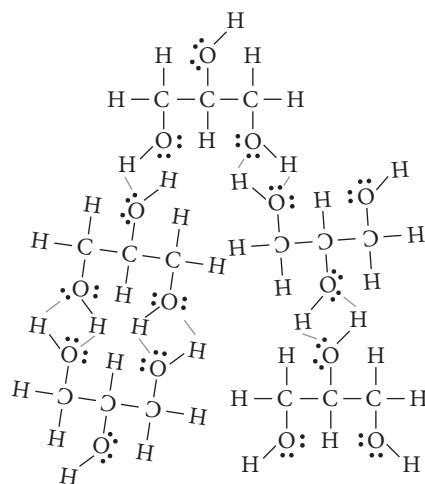
Molecular (Polar): crushes easily in a mortar with a pestle, dissolves in water, cannot conduct electric current in the dissolved state, can melt in a deflagrating spoon held in a Bunsen burner flame

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
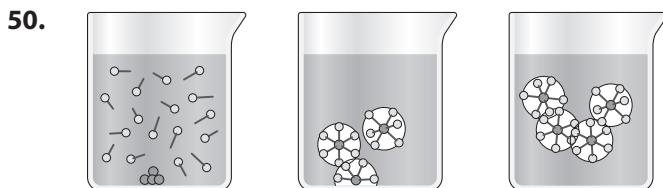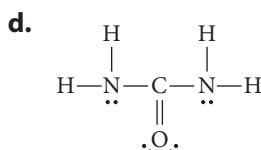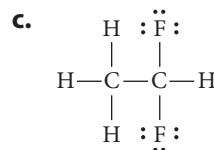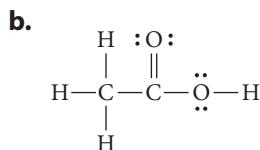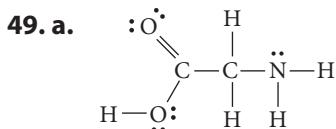
47. a.

**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.
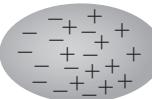
Step 2 Determine the total number of valence electrons (V) in all the atoms in the molecule. For polyatomic ions, add or subtract electrons to account for the charge. For example, for a cation with a charge of 2+, subtract two electrons from the total number of valence electrons that are calculated for 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.
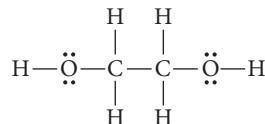
53. First, use a magnet to remove the nickel balls. Second, add water to dissolve the table salt. Filter to remove the solids—the glass balls and wax. Allow the water to evaporate to obtain the dry salt. Next, heat the solids to melt the wax. Pour off the wax and let it cool. The only remaining solids 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. $= 197.3^{\circ}$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.
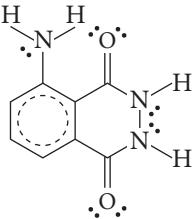
d. In order for water to freeze, it needs to form hydrogen bonds with other water molecules. Ethylene glycol prevents water from freezing because it interferes with the ability for water molecules to form hydrogen bonds with each other.

57. a. $[Kr]5s^24d^{10}5p^3$

b. Sb^{-3}: $[Kr]5s^24d^{10}5p^6$, 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.
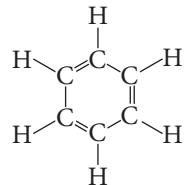
b. Luminol is mainly non-polar with some dipolar regions. It is insoluble in water but soluble in most polar organic solvents such as alcohols, aldehydes, and ketones.

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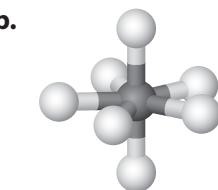
b. The electrons forming the second bond in the double bonds are really delocalized and shared among all six of the carbon atoms in the ring. All of the carbon-carbon bonds are identical in length and strength. The circle represents these delocalized electrons.

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 + 1d 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.
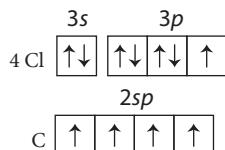
BaCl$_3$: Barium is a Group 2 metal which only forms ions with the charge of 2+. It would not combine with three chloride ions.

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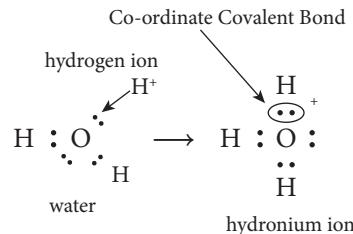
15. a. $\Delta EN = 1.8$, therefore ionic

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

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

d. $\Delta EN = 0.0$, therefore covalent
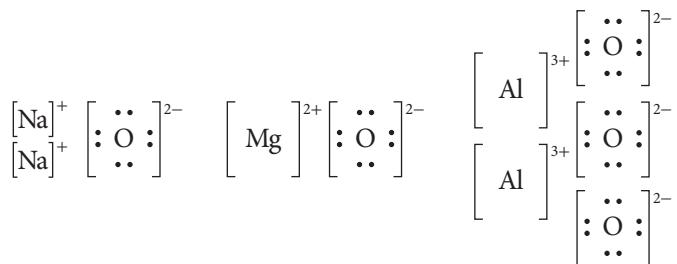
16. Both of the electrons in a co-ordinate bond are donated by only one of the atoms. For example, when a hydrogen ion combines with a water molecule to form a hydronium ion, the new bond is a co-ordinate bond because both electrons came from the oxygen atom.
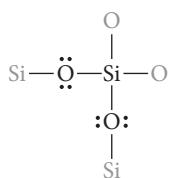
17. Compounds a, b, c, and e would be expected to dissolve in water because they are all polar molecular shapes (single polar covalent bond, trigonal pyramid, asymmetrical tetrahedron, bent, respectively); d is a trigonal plane and is non-polar.

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
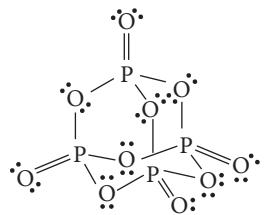
$(SiO_2)_n$ Silicon dioxide is a network solid. The shaded symbols in the Lewis diagram show the atoms in the next unit to which the atoms in the central unit are bonded.
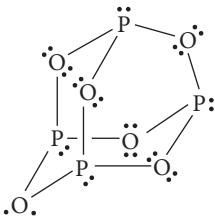
P_4O_{10}, P_4O_6, SO_3, SO_2, Cl_2O_7, Cl_2O are molecular compounds.
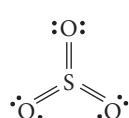
P_4O_{10}: The phosphorus atoms have expanded valences.
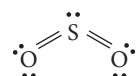
P_4O_6: The phosphorus atoms have expanded valences.
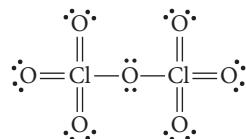
SO_3: The sulfur atom has an expanded valence.
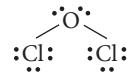
SO_2: The sulfur atom has an expanded valence.

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: $[He]2s^22p^1$; B^{3+}: $1s^2$ does not form

b. Mg: $[Ne]3s^2$; Mg^{2+}: $[He]2s^22p^6$

c. S: $[Ne]3s^23p^4$; S^{2-}: $[Ne]3s^23p^6$

d. K: $[Ar]4s^1$; K^+: $[Ne]s^23p^6$

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
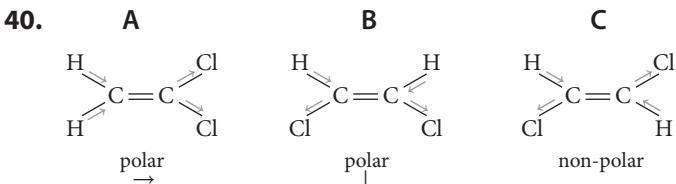
35. iodine

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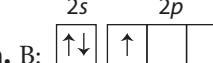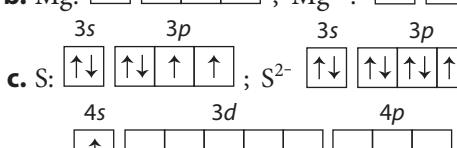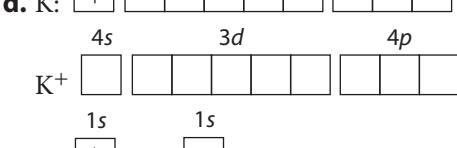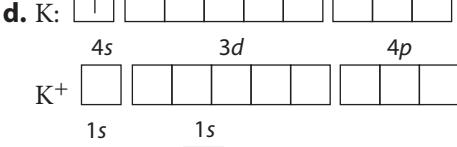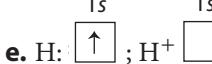
**38.** The attractions among the sodium ions and iodide ions are weaker than those among sodium and iodide ions for reasons given in the answer to question 37. Therefore, it is easier for the dipoles of the water molecules to pull apart the ions in the NaI crystals than in the NaF crystals.

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.

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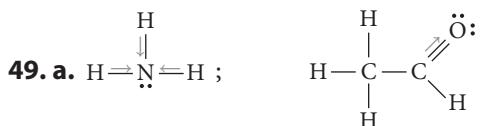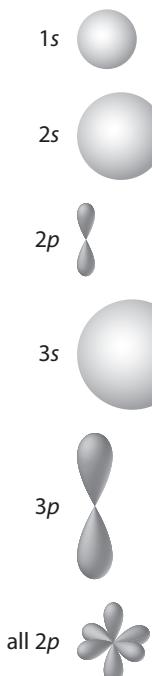
46. a. H_2O: H_2O forms hydrogen bonds whereas H_2S does not. Therefore, the intermolecular forces holding H_2O molecules are stronger.

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.

1s

2s

2p

3s

3p

all 2p

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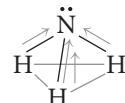
b. Count the number of bonding pairs and lone pairs around the central atom. Total number of groups (BP + LP) determines electron group arrangements. Number of BP and LP determines molecular shape.

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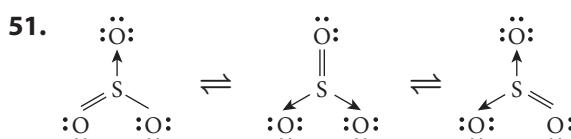
Mostly covalent bonds and co-ordinate covalent bonds: can be non-polar

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 $5p$ shell is empty and the $5s$ shell is filled.

- Te can form Te^{6+} by losing 6 electrons, resulting in the configuration $[Kr]4d^{10}$. The $5p$ and $5s$ shells are both empty leaving the $4d$ 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 90º 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
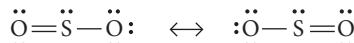
15. The electron configuration shows only the principle (n) and orbital shape (l) quantum numbers. It gives no information about the magnetic (m_l) or spin (m_s) quantum numbers.

16. a. $m_l = -1, 0, +1$
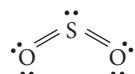
b. $5p$

c. 3

17. a. resonance structure:

$$\ddot{O}=\ddot{S}-\ddot{O}: \quad \leftrightarrow \quad :\ddot{O}-\ddot{S}=\ddot{O}$$
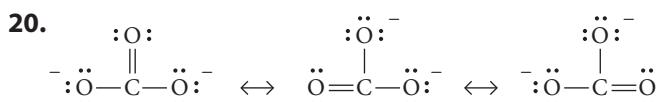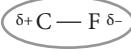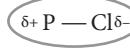
expanded octet:

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. $H-C\equiv N:$

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.

$\delta+ C — O \delta-$ (circled)   not polar   C — S

b.

$\delta+ C — F \delta-$ (circled) $\delta+ C — N \delta-$

c.

not polar   P — H $\delta+ P — Cl\delta-$ (circled)

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.