Answers to Unit 2 Review Questions

(Student textbook pages 264-7)

1. b

2. a

3. d

4. e

5. e

6. a

7. a

8. c

9. c

10. b

11. e

12. e

13. a

14. b

15. d

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

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

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

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

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

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

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

20. a. ultraviolet, X ray, gamma rays

b. violet

c. infrared, microwaves, radio waves

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

e. red

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

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

23. Answers should include any three of the following:

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

ii. Electrons can exist only in allowed orbits.

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

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

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

25. a. solid but malleable and ductile

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

c. solid and extremely hard, difficult to ignite

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

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

28. (electron group arrangement; molecular shape)

a. octahedral; square pyramidal

b. trigonal bipyramidal; seesaw

c. trigonal bipyramidal; linear

d. tetrahedral; trigonal pyramidal

e. octahedral; square planar

f. tetrahedral; tetrahedral

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

30. **a.** Bader: developed AIM theory (atoms in molecules); links quantum mechanics to the atoms and bonds in a molecule
  b. LeRoy: developed the "LeRoy radius," which is related to the distance between bonded atoms that are about to split into individual atoms
  c. Gillespie: developed the concepts of VSEPR

31. **a.** B: [He]$2s^2 2p^1$; B^{3+}: $1s^2$ does not form
  b. Mg: [Ne]$3s^2$; Mg^{2+}: [He]$2s^2 2p^6$
  c. S: [Ne]$3s^2 3p^4$; S^{2-}: [Ne]$3s^2 3p^6$
  d. K: [Ar]$4s^1$; K^+: [Ne]$s^2 3p^6$
  e. H: $1s^1$; H^+: no electrons

32. **i.** c
  ii. a
  iii. b
  iv. c
  v. b

33. **a.** boron
  b. fluorine
  c. manganese
  d. calcium
  e. hydrogen
  f. argon

34. carbon

35. iodine

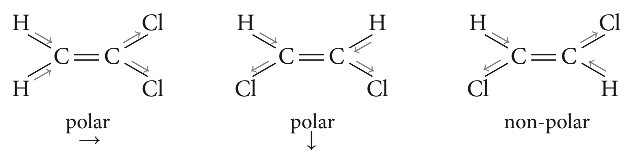
36. **a.** Group 1
  b. i. [Uuo 118]$8s^1$
    ii. $n = 8, l = 0, m_l = 0, m_s = +\frac{1}{2}$ or $-\frac{1}{2}$
  c. very soft solid or a liquid
  d. reacts explosively with water

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

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

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

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



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

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

a. B: 2s [↑↓] 2p [↑][ ][ ] (boron ions do not exist)

b. Mg: 3s [↑↓] 3p [ ][ ][ ]; Mg^{2+}: 3s [ ] 3p [ ][ ][ ]

c. S: 3s [↑↓] 3p [↑↓][↑][↑]; S^{2-} 3s [↑↓] 3p [↑↓][↑↓][↑↓]

d. K: 4s [↑] 3d [ ][ ][ ][ ][ ] 4p [ ][ ][ ];
K^+ 4s [ ] 3d [ ][ ][ ][ ][ ] 4p [ ][ ][ ]

e. H: 1s [↑]; H^+ 1s [ ]

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

43. **a.** polar covalent; $\Delta EN = 1.3$
  b. mostly covalent; $\Delta EN = 0.0$

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

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

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

b. *sample answer*: lard

45.

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

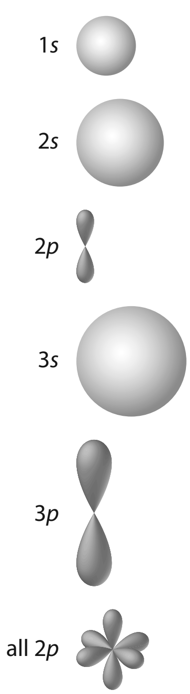
46. a. H_2O: H_2O forms hydrogen bonds whereas H_2S does not. Therefore, the intermolecular forces holding H_2O molecules are stronger.

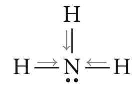
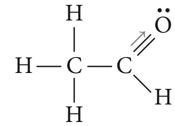
b. $SiCl_4$: Sulfur tetrachloride is unstable and decomposes before it can boil.

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

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

48.

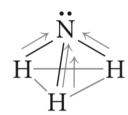


49. a.  ; 

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

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

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



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

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

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

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

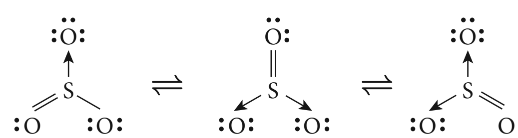
Polar covalent bonds and mostly covalent bonds: One electron contributed by each atom.

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

Polar covalent bonds: Always polar

Mostly covalent bonds: Always non-polar

51.



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

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

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

53.

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

54. Discussions should include the following points.

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

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

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

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

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

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

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

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

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

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

SiF_4 forms a tetrahedron; XeF_4 forms a square plane

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

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

61. a. iron

b. Group: 8; period: 4

62. a. oxygen

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

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

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

Answers to Unit 2 Self-Assessment Questions

(Student textbook pages 268-9)

1. b

2. d

3. e

4. b

5. c

6. a

7. e

8. d

9. c

10. d

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

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

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

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

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

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

c. i. *The Bohr model*

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

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

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

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

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

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

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

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

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

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

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

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

14. Co

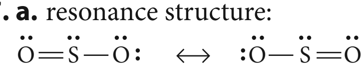
15. The electron configuration shows only the principle (n) and orbital shape (l) quantum numbers. It gives no information about the magnetic (m_l) or spin (m_s) quantum numbers.

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

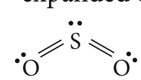
b. 5p

c. 3

17. a. resonance structure:



expanded octet:

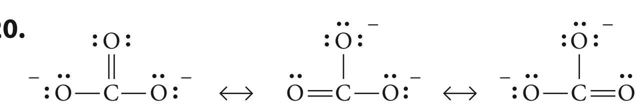


b. The experimentally measured bond lengths of the S-O bonds are shorter than would be predicted by the resonance structure (between single and double bond), but instead, are in agreement with the length of a double bond. Therefore, the expanded octet, or expanded valence, is the most likely structure.

18. Group 13; Period 5; p block

19. $H-C\equiv N:$

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. $\delta+\, \mathrm{C} - \mathrm{O}\, \delta-$ not polar $\mathrm{C} - \mathrm{S}$

b. $\delta+\, \mathrm{C} - \mathrm{F}\, \delta-$ $\delta+\, \mathrm{C} - \mathrm{N}\, \delta-$

c. not polar $\mathrm{P} - \mathrm{H}$ $\delta+\, \mathrm{P} - \mathrm{Cl}\, \delta-$

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

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

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