Answer Key

Unit 1 Matter, Chemical Trends, and Chemical Bonding

Unit Preparation Questions (Assessing Readiness)

(Student textbook pages 4–7)

1. e

2. e

3. Maps should correctly reflect the layout of the room and the location of each safety device listed.

4. Students' scripts should describe the setting, the materials needed, and the dialogue required. The steps for using the extinguisher are: pull the pin, aim the nozzle at the base of the fire, and squeeze the handle slowly while sweeping the spray from side to side.

5. Answers should be written in a persuasive tone and should present the potential hazards that are possible in even a procedure as simple as this. For example, spills that could damage clothes, and chemicals that could splash or drift into eyes.

6. Broken glass in the regular garbage could hurt the custodian. Having a special container for broken glass signals that safety precautions must be followed.

7. a. Material is flammable and combustible.

b. The gas is under pressure. The canister might explode if heated or punctured.

c. Material is poisonous and acts quickly with serious effects.

d. Material is corrosive and can cause mild to severe chemical burns.

8. Sample answer:

a. Keep flammable materials away from heat, sparks, and direct sunlight.

b. Keep the canister away from heat and secure the canister to keep it from being dropped.

c. Use protective equipment and procedures to prevent contact with the material. For example, work under a fume hood and wear gloves, an apron, and safety goggles.

d. Wear protective gloves, an apron, and safety goggles.

9. a. poisonous materials having immediate and serious toxic effects (scull and crossbones)

b. flammable and combustible (flame over a line)

c. corrosive material (hand getting burned)

d. compressed gas (gas tank outline)

10. b

11. a. He **b.** C **c.** Ca **d.** Na **e.** S **f.** O **g.** Ar **h.** F

12. a. phosphorus **e.** potassium

b. aluminum **f.** lithium

c. nitrogen **g.** hydrogen

d. beryllium **h.** neon

13. Sample answer:

- The symbol is the first letter of the element name, such as H for hydrogen.
- The symbol is the first and second letters of the element name, such as He for helium.
- The symbol is the first and third letters of the element name, such as Mg for magnesium.
- The symbol is based on the (original) name for the element in another language, such as W for tungsten (Wolfram).

14. Because each chemical symbol has only one capital letter in it, if two capital letters are written together it indicates two elements that are joined to form a compound. An incorrectly capitalized chemical symbol would be read as a different compound that would have different properties and undergo different reactions than were intended.

15.

Incorrect Symbol	Element Intended	Correct Symbol
Bo	boron	B
Fl	fluorine	F
Po	phosphorus	Po
be	beryllium	Be
Ch	chlorine	Cl
Hy	hydrogen	H
Ma	magnesium	Mg
HE	helium	He
Ni	nitrogen	N
s	sulfur	S
Ox	oxygen	O
Sil	silicon	Si

16. a. carbon and oxygen

b. hydrogen and oxygen

c. lithium and fluorine

d. nitrogen and chlorine

e. magnesium, nitrogen, and oxygen

f. aluminum, sulfur, and oxygen

g. calcium and sulfur

17. Universally recognizable symbols avoid ambiguity and potential confusion with possibly deadly results. They also reduce language barriers for sharing information.

18. The chemical symbol for carbon is C, not Ca, which represents calcium.

19. c

20. a

21. d

22. a. bubbling and colour changes

b. release of energy (light)

23. Sample answer: a burning candle

24. It can sense heat released (warm) or absorbed by (cool) the reaction.

25. Bubbles that were formed because of boiling rather than from gas formation.

26. Sample answer: If two colourless gases react to form a new colourless gas, the chemical change might not show an outward, visible sign.

27. a. Some chemical changes absorb energy. Light energy could cause a chemical change.

b. Hydrogen peroxide is less effective as a disinfectant as light exposure changes it into water and oxygen.

28. a. A white precipitate is forming.

b. Since limewater is an indicator for the presence of carbon dioxide, the carbon dioxide in the student's breath is causing the chemical change.

29. d

30. a

31. b

32. a. positive **b.** negative **c.** negative **d.** positive

33. Electrons are negatively charged. After gaining electrons, an atom has more negatively charged electrons than positively charged protons, so the resulting ion is negatively charged.

34. a. nitrate **b.** sulfate **c.** hydroxide **d.** carbonate

35. A metal and a non-metal form an ionic compound.

36. a. ionic **b.** molecular **c.** ionic **d.** molecular

37. a. ionic; KF **c.** molecular; N_2O

b. molecular; PBr_3 **d.** ionic; $Al(NO_3)_3$

38. Electrons are transferred from one atom to another during the formation of an ionic compound. They are shared between two atoms during the formation of a molecular compound.

39. a. metals **b.** cations (positive ions)

40. The names of ionic compounds include a metal's name and the name of a non-metal ending in –ide, or include the names of polyatomic ions which end in the prefixes –*ate* or –*ite*. The names of molecular compounds use prefixes that indicate the proportions of elements and include the names of non-metals but not metals.

Chapter 1 Elements and the Periodic Table

Learning Check Questions

(Student textbook page 14)

1. (1) In the Thomson model of the atom, the positive charge is spread over the entire atom, whereas in the Rutherford model, the positive charge is contained in a very small volume at the centre of the atom. (2) In the Thomson model, the negative charges are embedded in the large positively charged region. In the Rutherford model, the negative charges orbit the tiny positive charge.

2. See the figure for answer 2 on page 639 of the student textbook.

3. The radius that Bohr calculated for the orbit of the electron in the hydrogen atom is the same as the average distance that Schrödinger calculated for the electron from the nucleus of the hydrogen atom.

4. Models represent an understanding of or idea about an object or concept. Chemists can use models to predict the properties of a substance and design further experiments to test and, if necessary, modify the model.

5. Dalton—All matter consists of tiny particles (called atoms). Atoms of each element are unique. Thomson—Atoms contain negatively charged particles that can be ejected from the atom. Bohr—Electrons exist only in certain allowed energy levels in an atom.

6. $2n^2$: $2 \times 8^2 = 2 \times 64 = 128$

(Student textbook page 26)

7. Mendeleev listed the elements vertically, in order of atomic mass (called atomic weight, at that time). When he came to an element with properties similar to one higher in the list, he started a new column by putting the next element beside the one that had similar properties.

8. When elements are arranged by atomic number, their chemical and physical properties recur periodically. Many elements have similar properties and these properties follow a pattern that repeats itself regularly.

9. Each column in the periodic table constitutes a group. Groups contain elements with similar chemical and physical properties. Each row in the periodic table constitutes a period. The atomic number of the elements increases sequentially across a period. The outermost electron shell that is occupied is the same for each element in a period.

10. Travelling across a period from left to right, the number of electrons in the valence shell increases until the last element in each period has a full valence shell (which indicates that it is a noble gas).

11. The elements in the periodic table are categorized in several different ways. In one case, elements are categorized by whether they are metals, metalloids, or non-metals. In another case, the elements are categorized by very specific chemical and physical properties. Elements are also categorized by dividing the periodic table into blocks.

12. See the flowchart in the selected answers on page 639 of the student textbook.

(Student textbook page 34)

13. The radius of an atom is the radius of a sphere within which electrons spend 90 percent of their time.

14. Electrons exist in a region that can be described as a cloud rather than having defined boundaries. There is currently no way to directly measure the radius within which electrons spend 90 percent of their time.

15. As the charge of a nucleus increases, it exerts a greater force on the electrons. Thus, for electrons in a given energy level, the electrons are drawn closer to the nucleus. As a result, the size of the atom decreases across a period from left to right.

16. Electrons in filled shells reduce the effect of positive charge on the outer electrons. Thus, outer electrons are not as strongly attracted to the nucleus as they would be if the electrons in the lower energy levels were absent. As a result, the size of an atom increases from the top to the bottom of a group.

17. Increasing atomic number: oxygen (8), potassium (19), krypton (36), tin (50). Increasing size (atomic radius): oxygen (73 pm), krypton (112 pm), tin (140 pm), potassium (227 pm). As the atomic number increases going across a period from left to right, the nuclear charge increases, which means there is more pull on the electrons and therefore the atomic radius decreases. Thus, within a period, the progression of the atomic number and size are opposite. Going down a group, however, even though the atomic number increases, the effective nuclear charge is reduced due to shielding; the atomic radius therefore increases. Also, the number of occupied electron shells increases, making the atoms larger.

18. nuclear charge; number of occupied electron shells; shielding; number of valence electrons

Caption Questions

Figure 1.10 (Student textbook page 22): Columns represent periods. The length of the period depends on the number of electrons allowed in the highest energy electron shell. That number increases with period number.

Figure 1.15 (Student textbook page 29):
alkali metals: Group 1
alkaline earth metals: Group 2
actinoids: green ring around the outside of the right arm from Ac to Lr
lanthanoids: orange group inside the actinoid ring and including Lu
halogens: Group 17
noble gases: Group 18

Figure 1.18 (Student textbook page 33): Atomic radii get smaller from left to right across any period. Atomic radii get larger from the top, down in any group. The last three elements of each period have very similar atomic radii. The change in atomic radii across a period shows the most dramatic change between groups 2 and 13 and between groups 15 and 16.

Figure 1.22 (Student textbook page 36): Fluorine has the greatest electronegativity and francium has the smallest electronegativity. They are on diagonally opposed corners of the periodic table.

Figure 1.23 (Student textbook page 37): Nuclei that can get closer to the outer electrons of another atom will attract those electrons with a greater force. Therefore, atoms with smaller radii will have a higher electronegativity compared with atoms with larger radii.

Figure 1.25 (Student textbook page 38): Ionization energy, electron affinity, and electronegativity all follow the same trends. Atomic radius follows trends opposite to the other three.

Section 1.1 Review Questions

(Student textbook page 21)

1. Schrödinger's atom describes electrons as existing in regions of space, represented as electron clouds. His model was a mathematical equation that defined the atom in terms of energy.
2. By looking at the electron cloud model, you cannot determine the number of electrons in the atom or how many are in each energy level.
3. Phosphorous. It has five valence electrons, one electron pair, and three unpaired electrons.
4. See Figure 1.8 on page 15 of the student textbook.
5. Isotopes are atoms with the same number of protons (and are therefore atoms of the same element) but with different numbers of neutrons.
6. When a nucleus is unstable, it can emit a negative particle and one of the neutrons in the nucleus becomes a proton. Since the atomic number changed, it became a different element.
7. Isotopic abundance is the relative amount of an isotope as compared to the total amount of all isotopes of the element.
8. More than one isotope of most stable elements are found on Earth. The reported value is a weighted average of the masses of the naturally occurring isotopes.
9. **a.** bromine-81; $^{81}_{35}Br$; 35; 81; 35; 35; 46
   b. neon-22; $^{22}_{10}Ne$; 10; 22; 10; 10; 12
   c. calcium-44; $^{44}_{20}Ca$; 20; 44; 20; 20; 24
   d. silver-107; $^{107}_{47}Br$; 47; 107; 47; 47; 60
10. 28.08 u
11. Since the average atomic mass of yttrium and the mass of the isotope Y-89 are the same, it can be inferred that yttrium exists in only one isotopic form, Y-89.
12. Atomic number is defined by the number of protons in the nucleus (they are the same thing). Neutrons and protons exist in the nucleus in a ratio that keeps nuclear forces in balance. Since protons have a repulsive force that is far stronger than the attractive force between neutrons, more neutrons than protons are required for balance. The key is that the number of protons increase, and by necessity or definition that means that both the atomic number and ratio of neutrons to protons increase as well.
13. Method four requires only that you know the percentage of each item that is present. Isotopic abundance, given in percentages, is the information you have available when calculating average atomic mass. The other methods require you to know the total number of items.
14. When an organism dies, the amount of C-14 is no longer replenished, so the net ratio of C-14 decreases compared to C-12. Because we know how quickly C-14 decays (is removed from an organism) we can calculate the age of a fossil based on how much C-14 is left in it.

Section 1.2 Review Questions

(Student textbook page 30)

1. Properties of elements fall into a pattern based on atomic number, not on atomic mass as Mendeleev was trying to arrange them.
2. Because they had not yet been discovered
3. Elements in the same period have atoms with the same number of electron shells. Elements in the same group have atoms with the same electron configuration in their valence shells.
4. Sample answer: Sodium must be highly reactive, being an alkali metal found in Group 1. Storing it in oil reduces the likelihood of a violent reaction.
5. Sample answer: Silicon is a brittle metalloid that would break apart if pressure was used to change its shape (the method used for the copper platter). It has a very high melting point (1687 K) and could not be easily moulded into another shape.

6. Main-group elements are the most prevalent elements on Earth. The valence electron configurations of their atoms are completely predictable based on their group (except group 12).

7. The non-metal category includes elements that are gases at room temperature.

8. copper, silicon, iodine

9. 32

10. main-group elements, non-metals, and halogens

11. The name "rare earth metals" refers to the rarity in which they are found in pure form.

12. Strontium is in the same group as calcium, and is slightly more reactive. They are likely to be found in similar places (i.e., bones).

13. a. Sodium (Na) is more reactive. Sodium (Na) is a Group 1 alkali metal while magnesium (Mg) is a Group 2 alkaline earth metal. The alkali metals are the most reactive metals.

    b. Bromine (Br) is more reactive. Bromine is a Group 17 halogen, the most reactive group of non-metals. Krypton (Kr) is a noble gas in Group 18, which are unreactive non-metals.

    c. Hydrogen (H) is more reactive. Though it is in Group 1, hydrogen is a reactive non-metal, not an alkali metal. Helium (He) is a Group 18 noble gas, the most unreactive of the non-metal elements.

14. Answer may be any of the alkali metals.

15. non-metals

16. Sample answer:

Benefits of Various Periodic Tables

Form of Periodic Table	Benefits
Standard	• Used by scientists all over the world • Compact and straightforward to read • Shows many levels of similar properties • Is available in interactive forms online
Circular	• Emphasizes the continuous increase in atomic number • Highlights the repeating nature of properties of elements
Pyramidal	• Shows how elements relate in terms of their electron configurations
Spiral	• Same as the circular periodic table, and visually groups elements according to properties • Is available with detailed interactive software
Cupcake	• Edible

Section 1.3 Review Questions

(Student textbook page 40)

1. Atomic radius is the distance from the centre of an atom to the boundary within which its electrons spend 90 percent of their time. Chemists use X-ray crystallography, neutron diffraction, and electron diffraction to measure the atomic radii of different atoms.

2. O, Sb, Sn, Ba, Cs

3. Within a group, atomic radius increases from the top to the bottom of the periodic table. This is because each row lower down a group, the number of occupied electron shells increases. Even though the positive charge of the nucleus increases, the inner shells shield the outer electrons from the positive charge and the electrons are not attracted as strongly as they would be without the shielding. Therefore, each new occupied shell increases the atomic radius.

4. (A) The sodium atom decreases in size because after losing one electron, there are fewer negative charges repelling but the same number of positive charges attracting, pulling the atom tighter together.

   (B) The chlorine atom increases in size when it gains one electron, because there are more negative charges repelling but the same number of positive charges attracting, thus pushing the atom wider.

5. Ionization energy is the amount of energy required to remove the outermost electron from an atom or ion in the gaseous state. The first ionization energy can be represented by the generalized equation A + energy (1^{st} ionization energy) $\rightarrow A^+ + e^-$, in which A represents any atom.

6. Elements 20 to 30 are all in the same period. They do not include any Group 1 elements (with very low ionization energy) or any Group 18 elements (with very high ionization energy). Also, there are three inner shells shielding the outer electrons for those elements, decreasing the attractive force of the nucleus.

7. $A^{2+} + 3^{rd}$ ionization energy $\rightarrow A^{3+} + e^-$; The third ionization energy is greater than the first or second ionization energy. When an electron is removed, there are fewer negative charges repelling but the same number of positive charges attracting. The force of attraction per electron increases with fewer electrons to hold. Consequently, more energy is required to remove each successive electron.

8. Because a helium atom has only two electrons

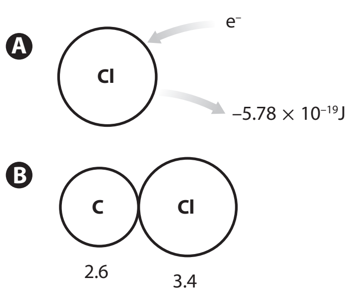
9. cesium, strontium, arsenic, phosphorus, fluorine, helium

10. An element with a positive electron affinity is an element that absorbs energy when an electron is added. Adding energy to an atom makes it more unstable, so elements with positive electron affinities are elements that are unstable when they gain an electron.

11. Halogens have the most negative electron affinities. When they gain an electron, they also gain a filled outer electron shell, which is a very stable configuration.

12. Electron affinity is the energy released or absorbed when an electron is added to a neutral atom (A). In contrast, electronegativity is a relative measure of the ability of an atom to attract shared electrons (B). A key difference is that electron affinity is an actual measured quantity of energy, while electronegativity is measured relative to other elements.

Electrons are attracted more strongly to chlorine in a chlorine-carbon bond (e.g., in carbon tetrachloride) because chlorine has a higher electronegativity. This is because the atoms are not much different in size but an atom of chlorine has a greater effective nuclear charge to attract the electrons.



13. There is an inverse relationship between atomic size and electronegativity. The larger the atom, the more electrons are present to shield the attractive force of the nucleus, and the less the atom is able to attract shared electrons.

14. Lead must be more dense because its atomic number is higher.

Element	Atomic Mass	Atomic Radius
aluminum	26.98 u	1.43×10^{-10} m
lead	207.2 u	1.46×10^{-10} m

15. a. 2.45×10^{-19} J

b. The combination of Na^+ and Cl^- would be more stable because of the electrical attraction between opposite charges.

Quirks and Quarks: Unearthing an Ancient Andean Element Questions

(Student textbook page 20)

1. Sample answer: To protect themselves from inhaling the mercury, the Incan miners could have found another way to extract the mercury that didn't involve heating the cinnabar, or they could have built tall chimneys to send the mercury far away (which would have solved the problem for them, but not for anyone downwind).

2. Sample answer:

Risks: Mercury waste travels through the water cycle and bioaccumulates. Even at relatively low levels it has serious toxic effects on living things including humans.

Benefits: Mercury is useful for many applications including batteries, CFLs, and electronics.

3. Journalists often have a degree in English and may also have an undergraduate or postgraduate degree in journalism or another specialized communications degree or diploma. Those who specialize in science usually have related training. Journalists need to have excellent communication skills, inquisitive minds, and high stamina. As writers, reporters, broadcasters, and presenters, they work for newspapers, news agencies, online news providers, magazines, television stations, and radio stations.

Chemistry Connections: Elements of the Body

(Student textbook page 39)

Student answers should include the following key points:

- Benefits of adding sulfites to food: prolong the shelf life of (preserve) foods
- Quantity of sulfite needed: parts per million (to preserve), such as 1000 ppm in dried fruits
- Risks of adding sulfites to food: allergic or asthmatic reactions
- Natural occurrence of sulfites in food: during fermentation
- Decisions should demonstrate that students have considered the risks and benefits according to their research.

Practice Problems

(Student textbook page 19)

1. 35.45 u

2. 10.81 u

3. 6.94 u

4. 24.31 u
5. 39.9%; 69.72 u
6. 49.31%; 79.91 u
7. 85.47 u
8. The isotopic abundance of nitrogen-14 is very high, with just a trace of nitrogen-15.
9. 186.207 u
10. 192.969 u

Chapter 1 Review

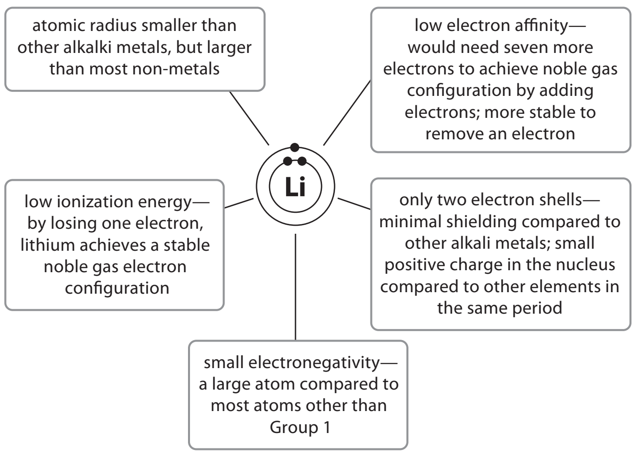
(Student textbook pages 45–7)

1. c
2. a
3. d
4. b
5. b
6. e
7. c
8. b
9. **a.** Should show four unpaired electrons, one at each cardinal point around the symbol
   b. Should show three unpaired electrons, eliminate the paired electrons at the top
   c. Should show only a single electron
10. Mendeleev made a card for each of the 63 elements that were then known. He arranged his cards vertically by atomic mass. When he came to an element that had properties similar to the element in the top row, he started a new column. He left spaces where elements seemed to be missing and occasionally reversed the order of elements according to atomic mass so that side-by-side elements would have the same properties.
11. The elements show periodicity in that, if they are arranged by atomic number, their chemical and physical properties recur over and over again in a regularly recurring (i.e., periodic) way.
12. Atomic radius is the distance from the centre of an atom to the boundary within which the electrons spend 90 percent of their time. This is not the same as the radius of a coin or the distance from the centre to the outer edge of an object because the boundaries are defined by probability, not a fixed size or position.
13. Electron affinity is the energy released or absorbed when an electron is added to an atom. Electronegativity, on the other hand, is an indicator of the relative ability of an atom to attract shared electrons. Electron affinity is a concrete quantity; electronegativity is a relative quantity.
14. Diagrams should reflect categories as shown on page 25 of the student textbook: transition elements are Groups 3–11, inner transition elements are lanthanoids and actinoids, and main-group elements are all the rest.
15. $A^{+} + \text{energy} \rightarrow A^{2+} + e^{-}$
16. If the electron affinity of an element is negative, the resulting ion released energy when it formed and is relatively stable.
17. antimony, tellurium
18. Each of these nuclei is unstable because it has too many neutrons.
19. 121.7580 u
20. **a.** sulfur **b.** radium
    c. Sample answer: The periodic trend is that electronegativity decreases down a group and increases across a period.
21. The missing element will have properties that are most similar to zirconium (e.g., malleable, ductile, and lustrous). It should have an atomic mass that is 1–3 u less than 180.9 u.
22. A neutral potassium atom has four occupied electron shells, with only one electron in its valence shell. That electron is significantly shielded from the nucleus and is able to exist some distance from the nucleus. When the atom forms a cation, it loses its valence electron and is left with only four occupied electron shells with an outer limit that is closer to the nucleus.
23. Halogens have electron configurations that are one electron short of a noble gas configuration. Because the noble gas configuration is so stable, a relatively large quantity of energy is released when a halogen gains an electron and achieves that configuration.
24. **a.** In general, electronegativity decreases down a group and increases across a period.
    b. The trend across a period is fairly clear-cut for Periods 2 and 3. For Periods 4, 5, and 6, the trend is disrupted with elements 30–31, 47–49, and 80–83, which show decreasing electronegativy, followed by an increase. This is due to shielding of the positive charge by the filled sub-shell.
    c. The trend of electronegativity decreasing down a group is fairly well defined for the main-group elements.

d. Electronegativity is closely related to atomic size: the closer the shared electrons can get to the nucleus, the more strongly they are attracted.

25. Both Bohr and Schrödinger tried to describe the location of an electron in relation to the nucleus.

26. Sample answer:



27. Diagrams should look like those on page 11 of the student textbook and indicate that Bohr added detail about energy levels.

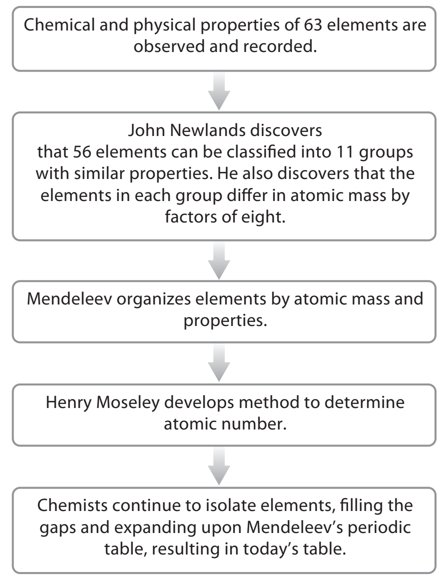
28. Answers should include the idea that different magnesium isotopes have the same number of protons (12) but different numbers of neutrons.

29. Concept maps should explain that effective nuclear charge is the apparent nuclear charge as experienced by the outermost electrons of an atom, which is different from the actual nuclear charge as a result of shielding by the inner-shell electrons. The effective nuclear charge of an atom is smaller than the actual nuclear charge. Within a group, because the number of inner electron shells increases, effective nuclear charge decreases with the result that outer electrons are attracted less strongly and can exist at a greater distance from the nucleus, thus increasing the size of the atom.

30. To determine a simple average of several values, add the values and divide by the number of values. For example, (65 kg + 80 kg + 92 kg) ÷ 3 = 79 kg gives each values are equal importance, or weight. A "weighted average", however, takes into consideration the abundance of each value. For example, if 20% of a group of people have a mass of 65 kg, 10% are 80 kg, and 70% are 92 kg, the "weighted average" is calculated as:

$65 \text{ kg} \times 0.20 + 80 \text{ kg} \times 0.10 + 92 \text{ kg} \times 0.70 = 85.4 \text{ kg}$ (round to 85 kg).

31. The periodic law is a statement that describes the repeating nature of the properties of the elements. Sample flowchart:



32. Presentations may be assess using Assessment Master BLM A-32 Presentation Rubric.

33. Scientists use X-ray crystallography to get a "picture" of a substance such as sodium, allowing them to measure the distance between atoms. The atomic radius is that distance divided by two. Diagrams should look like the figures in Activity 1.2 on page 31 of the student textbook.

34. Sample answer:

Property	Trend Going Down a Group	Trend Going Across a Period
Atomic Radius	increases	decreases
Ionization Energy	decreases	increases
Electron Affinity	decreasingly negative	increasingly negative
Electronegativity	decreases	increases

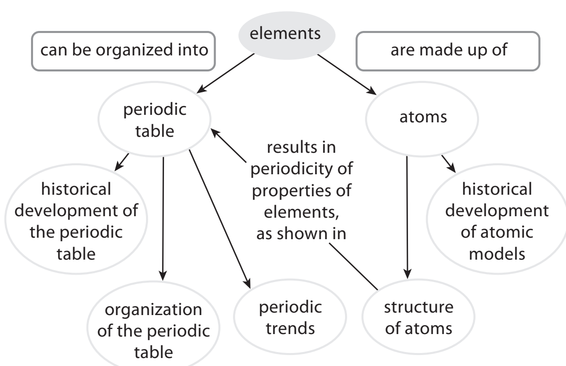
Atomic radius increases down a group because the number of occupied electron shells increases. The resulting shielding of the inner shells means valence electrons are not attracted as strongly to the nucleus. The property decreases across a period because the number of protons increases, which attracts the electrons more strongly to the nucleus. Shielding is not a factor across a period because the increasing number of electrons are all in the same electron shell.

Ionization energy decreases down a group because electrons are farther away and attracted less strongly to the nucleus down a group, as atomic radius increases. The property increases across a period; alkali metals lose electrons relatively easily to achieve the stable noble gas electron configuration, while noble gases, which already possess that stable electron configuration, require a great deal of energy to remove an electron.

Electron affinity increases up to the halogens which have the greatest electron affinity because adding an electron allows them to achieve the stable noble gas electron configuration (expressed as the largest negative values). The trend decreases down a group because the attractive force in the nucleus is decreased by shielding, reducing the stabilizing effect of the nucleus on the anion formed.

Electronegativity has an inverse relationship to atomic radius. The smaller the atom, the closer the shared electrons get to its nucleus and the more strongly they are attracted. Therefore, electronegativity decreases as atomic radius increases, and vice versa.

35. Sample answer:



36. Alkali metals are banned from many classrooms because they are extremely reactive with many substances, including water.

37. Sample answer: Because iodine accumulates in the thyroid gland, the iodine-131 would accumulate there and not harm other tissues. It would kill thyroid cells, which would restrict or destroy the function of an overactive thyroid.

38. Gold is unreactive, malleable, ductile, and conductive. These physical properties make it suitable for use in jewellery, dental work, and as a conductor.

39. Halogens are extremely reactive. It therefore makes sense that they react in a harmful way with the sensitive tissues of the nose, throat, and lungs. A noble gas, however, is extremely unreactive and would not cause any damage by reacting with tissue.

40. Sample answer: To develop smaller and smaller microchips, you would need to know a lot about the physical properties of substances, especially conductivity. Having a strong background in chemistry would give you that knowledge.

Chapter 1 Self-Assessment Questions

(Student textbook pages 48–9)

1. b

2. d

3. d

4. b

5. a

6. d

7. e

8. b

9. a

10. e

11. Isotopic abundance represents the amount of a given isotope of an element that exists in nature, expressed as a percentage of the total amount of the element. Average atomic mass is determined by multiplying the isotopic abundance of each element by its mass.

12. Schrödinger's model of the atom is the model currently accepted as the most accurate. However, it is described by a mathematical equation and visual representations do not give useful information about the atom such as number of electrons, protons, neutrons, or energy levels.

13. 151.96 u

14. The paired electrons are shown at the top, bottom, and right side of the fluorine symbol. The unpaired electron is shown at the left. Paired electrons interact in a unique way that allows them to be close to one another. Unpaired electrons are more likely to participate in bonding with other atoms.

15. Sample answer: Francium is likely the most reactive element, since the alkali metals are highly reactive and their reactivity increases down a group.

16. Rounding off the values, silver-107 has an atomic mass of 107 u, silver-109 has an atomic mass of 109 u, and the average atomic mass is 108 u. Since the average atomic mass is halfway between the masses of the two isotopes, they must be present in roughly equal amounts. In other words, the isotopic abundance of each will be roughly 50%.

17. a. First ionization energy is highest for the atom with atomic number 2 (helium). This value increases across a period, peaking at the noble gas, and then drops at the alkali metal in the next period. The pattern repeats.

b. Atomic radius is highest at the beginning of a period (at the alkali metal) and decreases across a period. With each new period, the highest atomic radius of that period increases compared to the highest atomic radius of the previous period.

c. In general, ionization energy decreases as atomic radius increases.

d. Because significant increases in atomic radius correspond to increased number of occupied electron shells, and since added occupied electron shells correspond to increased shielding of valence electrons from the nucleus, larger atoms tend to require less energy to give up an electron.

18. The periodic law states that when elements are arranged by atomic number, their chemical and physical properties recur periodically (at regular and repeating intervals). The design of the periodic table reflects this law in that the elements are shown in order of increasing atomic number, with a new row of the table corresponding to the beginning of the next period.

19. Chlorine is more volatile than bromine: chlorine is a gas at room temperature and bromine is a liquid. In a hot tub, enough chlorine would not stay dissolved to effectively disinfect (it would evaporate as the heat turned it into a gas). Also, the large amount of gaseous chlorine could be harmful.

20. Atoms of the elements of Period 4 can contain more than 8 electrons. Period 1 elements can contain only two electrons, and elements from Periods 2 and 3 can contain only eight electrons.

21. Atomic radius is the distance from the centre of an atom to the boundary within which the electrons spend 90 percent of their time.

22. Element X is to the right of Y, according to the periodic trend of atomic radii in which atomic radii decrease across a period. This is due to increased attractive force from the added proton in the nucleus with no shielding effects (the added electron is in the same shell).

23. Group 1 elements (alkali metals) have one electron in their valence shells. If an atom of a Group 1 element (except hydrogen) loses an electron, it will achieve the electron configuration of the closest noble gas, which is an extremely stable configuration. Thus, very little energy is required to remove that first electron, corresponding to a small first ionization energy. But a great deal of energy is required to remove a second electron, because doing so disrupts the extremely stable noble gas electron configuration.

24. The halogens (Group 17 elements) have the largest negative electron affinities. Atoms of these elements have one less electron in their shells than the nearest noble gas, meaning that adding one electron gives them an electron configuration of a noble gas. This is an extremely stable configuration, so that a relatively large quantity of energy is released when a halogen atom gains an electron.

25. See Figure 1.25 on page 38 of the student textbook.

Chapter 2 Chemical Bonding

Learning Check Questions

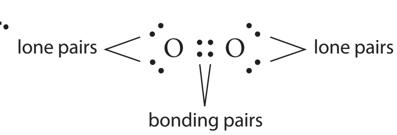
(Student textbook page 59)

1. When bonds form between atoms, the atoms gain, lose, or share electrons in such a way that they create a filled valence shell containing eight electrons. For example, a fluorine atom can gain one electron and become a fluoride ion that has a complete octet of electrons in its valence shell.

2. One calcium atom can donate an electron to each of two bromine atoms. The combination of one calcium ion and two bromide ions results in a neutral compound.

3. Determine the total number of valence electrons that each of the atoms in the compound should have and add them together. Count the number of electrons shown in the Lewis structure. If the numbers are equal, the compound is neutral and is a molecular compound. If the numbers are not equal, the compound carries a charge and is thus a polyatomic ion.

4.



5. Double bonds form when four electrons are shared by two atoms. Triple bonds form when six electrons are shared by two atoms. In some cases, when two atoms share two electrons, neither atom has an octet of electrons in its outer shell. When both atoms contribute another electron, they share four electrons in total to form a double bond. They might then both have an octet of electrons in their outer shells. If not, the atoms may share another electron each to form a triple bond. Double and triple bonds are used to ensure that all the elements in the compound have an octet of electrons in their outer shell.

6. A group of two or more atoms of non-metal elements can share electrons and form covalent bonds, but, as a group, must either lose or gain electrons so that they can all contain a stable octet of electrons. In this case, the atoms have formed a polyatomic ion. This ion can then form an ionic bond with other ions.

(Student textbook page 70)

7. A binary ionic compound is an ionic compound that contains atoms of two and only two different elements.

8. **a.** K_2S, potassium sulfide
   b. MgO, magnesium oxide
   c. $FeCl_2$, iron(II) chloride, or $FeCl_3$, iron(III) chloride
   d. Mg_3N_2, magnesium nitride
   e. HI, hydrogen iodide
   f. $Ca(OH)_2$, calcium hydroxide

9. **a.** chromium(II) bromide
   b. sodium sulfide
   c. mercury(I) chloride
   d. lead(II) iodide
   e. aqueous hydrogen nitrate, or nitric acid
   f. potassium hydroxide

10. **a.** $ZnBr_2$ **d.** $MgCl_2$
    b. Al_2S_3 **e.** H_3N
    c. Cu_3N_2 **f.** $Cu(OH)_2$

11. **a.** hypofluorite **c.** fluorate
    b. fluorite **d.** perfluorate

    When there is a family of compounds that can have 1, 2, 3, or 4 oxygen atoms, the combination of prefixes and suffixes are *hypo*–...–*ite* for one oxygen atom, (none)...–*ite* for two oxygen atoms, (none) ...–*ate* for three oxygen atoms, and *per*–...–*ate* for four oxygen atoms.

12. **a.** $FeSO_4$ **d.** $Mg_3(PO_4)_2$
    b. $NaNO_3$ **e.** H_2CO_3
    c. $CuCrO_4$ **f.** $Al(OH)_3$

(Student textbook page 79)

13. When a substance is melting, the particles (ions or molecules) have gained enough energy to overcome the attractive forces between the particles.

14. The compound with a melting point of 714°C is probably an ionic compound and the one with a melting point of 146°C is probably a molecular compound.

15. Potassium iodide is an ionic compound and, in ionic compounds, each ion is attracted to every oppositely charged ion adjacent to it. There are no combinations of ions that are unique, so they cannot be called molecules.

16. A dipole-dipole force is an attractive force between the slightly positive end of a polar molecule and the slightly negative end of another polar molecule.

17. The forces of attraction among non-polar molecules are very weak. It takes only a small amount of energy for these molecules to pull apart. This means that a relatively low temperature is capable of supplying the small amount of needed energy.

18. Intermolecular forces include dipole-dipole forces and the weak forces among non-polar molecules.

Caption Questions

Figure 2.9 (Student textbook page 56): When you are confident that the Lewis structure is correct and all atoms have an octet of electrons, the number of shared electrons shows the number of bonds. One pair of electrons between two atoms represents one bond.

Figure 2.17 (Student textbook page 60): The electronegativity of chlorine is larger than the electronegativity of carbon, indicating that the chlorine attracts the shared electrons with a greater force than the carbon does.

Figure 2.22 (Student textbook page 74): water

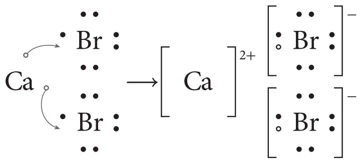
Section 2.1 Review Questions

(Student textbook page 63)

1. The atoms of noble gases are very stable; they do not tend to form chemical bonds with other atoms. Since noble gases have atoms with filled valence shells and since, for most main-group elements, a filled valence shell contains eight electrons, this configuration is called an octet. The octet rule states that when bonds form, the atoms involved share, transfer, or accept electrons to create a noble-gas-like filled outer shell.

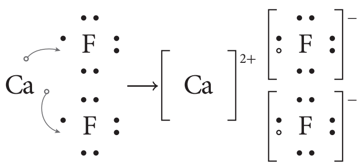
2. Most metals have fewer than four electrons in their valence shells. Since such atoms can lose electrons relatively easily, metals tend to lose electrons to achieve a filled valence shell. In contrast, atoms with four or more electrons in their valence shells (non-metals) gain electrons relatively easily. These atoms tend to gain electrons to achieve a filled valence shell.

3.

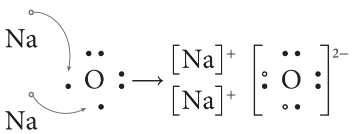


The calcium atom loses its two electrons to achieve a filled valence shell, becoming an ion with a charge of +2. The bromine atoms accept those electrons, becoming ions with charges of –1 and achieving filled valence shells.

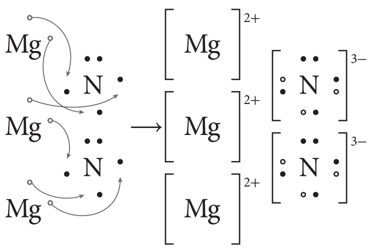
4. a.



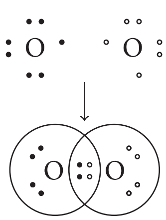
b.



c.

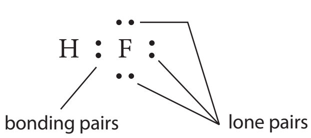


5.



6. 6 electrons (3 pairs of electrons)

7.



8. A neutral nitrogen atom has five valence electrons. A neutral oxygen atom has six valence electrons. Therefore the structure would have 23 electrons if it were a neutral molecule and 24 electrons if it were the nitrate ion with a charge of –1.

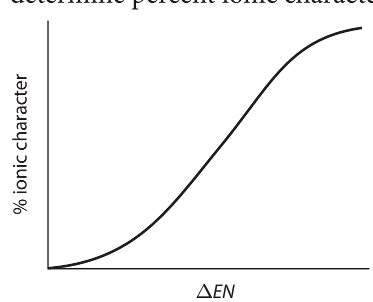
9. Ammonium iodide is considered an ionic compound because it is made up of a cation and an anion. Instead of a metal cation, the compound contains the ammonium cation.

10. a. 1.4; polar covalent
b. 0.4; polar covalent
c. 0.0; non-polar covalent
d. 1.5; polar covalent
e. 0.3; slightly polar covalent
f. 3.1; mostly ionic
g. 1.6; polar covalent
h. 1.8; mostly ionic

11. The partial negative charges are located on the atoms with the higher electronegativity. The partial positive charges are located on the atoms with the lower electronegativity.
a. carbon is slightly positive; fluorine is slightly negative
b. nitrogen is slightly positive; oxygen is slightly negative
c. N/A; non-polar covalent
d. copper is slightly positive; oxygen is slightly negative
e. hydrogen is slightly negative; silicon is slightly positive
f. N/A; mostly ionic
g. iron is slightly positive; oxygen is slightly negative
h. N/A; mostly ionic

12. a. C–S < N–O < H–Cl < Na–Cl
b. N–N < C–Cl < P–O < Mg–Cl

13. It is easy to find the electronegativity difference by looking up the electronegativities of the two atoms in the periodic table and calculating the difference. To find the percent ionic character of a bond, though, you would have to first find the electronegativity difference and then use a graph like the one sketched here to determine percent ionic character.



14. The symbol means that in the nitrogen–oxygen bond shown, the nitrogen has a slightly positive charge and the oxygen has a slightly negative charge.

15. Sample answer: A metal combined with a non-metal to form a solid is an ionic compound. One example found in nature is NaCl(s), salt.

16. The atmosphere consists almost entirely of elements (which have non-polar molecules) and carbon dioxide (which also has non-polar molecules).

Section 2.2 Review Questions

(Student textbook page 75)

1. **a.** Sodium chloride, NaCl(s), is table salt. Dihydrogen monoxide, $H_2O(\ell)$, is water. Sodium hydrogen carbonate, $NaHCO_3(s)$, is baking soda.
   b. sugar
   c. Table salt and sodium hydrogen carbonate are ionic compounds because they are made up of ions (sodium, chloride, and hydrogen carbonate). Sugar and water are molecular compounds. Sugar melts (decomposes, really) at a relatively low temperature, meaning it must be molecular, not ionic. Water exists as a liquid at room temperature, again meaning it must be molecular, not ionic.

2. Prefixes that indicate the numbers of atoms of the different elements are not needed in the names of ionic compounds because in most cases the charges of at least one of the ions will be obvious, and the number of each ion depends on a ratio that will result in a neutral compound. For multivalent metals, the ion charge is indicated after the name of the metal.

3. A polyatomic ion is an ion with more than one atom; essentially a charged molecule.

4. A sulfate ion, SO_4^{2-}, has one more oxygen atom than a sulfite ion, SO_3^{2-}. You would know the difference without looking up names in a table by realizing that the suffix *–ite* means that sulfite has fewer oxygen atoms relative to the ion with the suffix *–ate*, sulfate.

5. **a.** aluminum oxide **b.** mercury(II) iodide **c.** sodium phosphide **d.** potassium phosphate **e.** ammonium chloride **f.** lithium perchlorate **g.** aqueous hydrogen nitrate or nitric acid **h.** lithium hydroxide

6. **a.** ZnO **b.** FeS **c.** KClO **d.** MgI_2 **e.** $CoCl_3$ **f.** NaCN

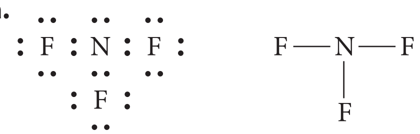
7. You need to use prefixes in the name of a molecular compound to indicate the numbers of atoms of each element present because two elements can bond in multiple ways. For example, nitrogen and oxygen can form six different compounds. (See the answer to #8.)

8. NO, nitrogen monoxide
   NO_2, nitrogen dioxide
   N_2O, dinitrogen monoxide
   N_2O_3, dinitrogen trioxide
   N_2O_4, dinitrogen tetroxide
   N_2O_5, dinitrogen pentoxide

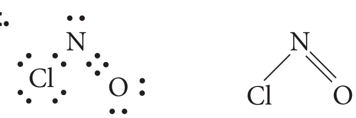
9. **a.** PCl_5 **b.** F_2O **c.** SO_3 **d.** $SiBr_4$ **e.** $Co(OH)_2$ **f.** SF_6

10. **a.** carbon monoxide **b.** boron trichloride **c.** carbon disulfide **d.** carbon tetrachloride **e.** silicon dioxide **f.** phosphorus triiodide **g.** barium hydroxide **h.** trihydrogen borate

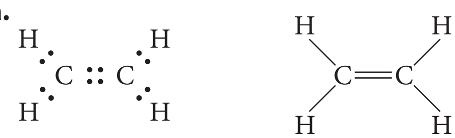
11. Because it is an organic compound, which follow a different naming convention.

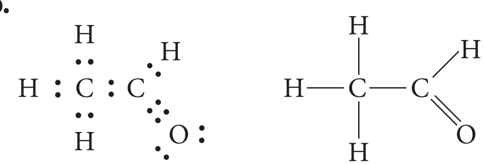
12. **a.**


b.


c.


13. Advantages might include the simplicity of structural formulas, which makes them easier to draw and interpret at a glance. Disadvantages might include the fact that they do not show the lone pairs of electrons or (directly) the number of electrons involved in bonding.

14. **a.**


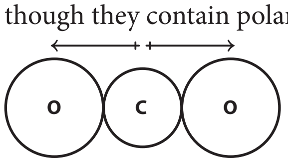
b.


Section 2.3 Review Questions

(Student textbook page 82)

1. The Celsius temperature scale is based on the melting point (0°C) and boiling point (100°C) of water.

2. When a substance is heated, the kinetic energy of its particles increase. As a solid is heated, its particles vibrate more and more energetically until they have enough energy to pull away from the other particles, forming a liquid. If heated further, the particles, which are still close together in a liquid, gain enough energy to bounce off one another rather than sticking together. The liquid forms a gas.
3. The strength of attractive forces between particles determines whether they will pull away from adjacent particles.
4. This compound has a low boiling point because it is a gas at normal temperature and pressure.
5. Polar molecules are attracted to one another by dipole-dipole forces, which are stronger than the weak attractions among non-polar molecules.
6. Compounds that dissolve in water tend to be ionic or polar and have higher melting points that non-polar compounds. The compound that does not dissolve in water is thus expected to have a lower melting point.
7. Although there is no distinct separation of charge, it is still possible for a nucleus in one non-polar molecule to attract the electrons in another non-polar molecule.
8. Compounds having a very high melting and boiling point are likely to be soluble in water. The high melting and boiling points are due to strong forces of attraction between charged particles called ions. The ionic solids in which these particles are held together can be pulled apart from their crystal lattice structure when surrounded by polar water molecules. This is the dissolving process.
9. Sample answer: Carbon dioxide is one example of a compound with molecules that are non-polar, even though they contain polar bonds.



10. For a substance to dissolve in water, the water molecules must be more strongly attracted to particles of water that it is to its own particles.
11. Since the glycerol dissolves, it is likely made up of polar molecules. Glycerol dissolves as polar water molecules surround the polar glycerol molecules. However, an aqueous solution of glycerol will not conduct an electric current. When electrodes are placed in a solution containing polar molecules, the molecules orient themselves so that their positive ends point to the negative electrode and their negative ends to the positive electrode. There is no movement of electrons.

    Since the glycerol molecules are polar, the melting point and boiling point would be expected to be intermediate in value, lower than that of ionic solids but higher than that of molecular, non-polar solids.
12. Ionic compounds can conduct in the liquid state and in an aqueous solution. In both instances, the ions are free to move.
13. Polar molecules cannot conduct electricity in the liquid state or in aqueous solution. In aqueous solution, these molecules orient themselves so that their positive end points to the negative electrode and their negative ends to the positive electrode. There is no movement of electrons. Acids are an exception to this since they react with water to form ions.

    In the liquid state, the polar molecules are not made up of separate positively and negatively charged particles. This is necessary if the solution is to conduct an electric charge. Each molecule has a positive and a negative end; overall the molecule is neutral.
14. Fat molecules are not soluble because they are only slightly polar. Protein molecules are polymers, very large molecules made up of amino acids with many sites that are polar, thus increasing the solubility. Molecules that make up fats can readily adhere to these protein molecules and are carried through the blood.
15. This is a general statement that is often valid. "Like" refers to similarity in structure: ionic compounds with polar solvents; polar molecules with polar solvents; non-polar solids with non-polar solvents. In nature, many factors affect the events we see and often explanations are not clear cut. There are many exceptions to "like dissolves like." For example, the molecules of all alcohols have a polar end but not all alcohols are soluble in water. At some point, the non-polar chain in the alcohol molecule dominates the solubility process.
16. a. The compound having the higher melting point, X, would be predicted to be more polar. Melting point depends upon the electrical attractions between molecules. The more polar the molecule, the greater the attractions and the higher the melting point. Compounds X is polar; compound Y less polar or possibly non-polar.

b. Compound X has a melting point of intermediate value. It is lower than that of ionic solids, but higher than that of molecular, non-polar solids. Compound X would be expected to be made up of polar molecules. This compound will dissolve in water as polar water molecules surround its polar molecules.

Since it has a low melting point, compound Y is likely made up of non-polar (or only slightly polar) molecules with low attractions between molecules. Polar water molecules will not be attracted to these molecules and the solubility would be low.

Practice Problems

(Student textbook page 73)

1. tetraphosphorus heptasulfide

2. lead(II) nitrate

3. $MnCl_4$

4. NI_3

5. copper(I) bromide

6. Fe_2O_3

7. SiO_2

8. selenium hexafluoride

9. calcium oxide

10. $Co(NO_3)_3$

Chapter 2 Review Questions

(Student textbook pages 89–91)

1. c

2. c

3. a

4. d

5. b

6. e

7. b

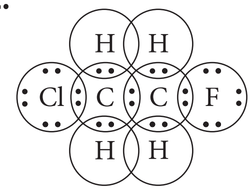
8. d

9. The compounds found in ores are ionic compounds because compounds made of metals combined with non-metals are usually ionic compounds.

10. The compounds would be molecular compounds because they are made up of non-metals. When non-metals bond together, with a few exceptions involving polyatomic ions, they form molecular compounds.

11. Two aluminum ions have a charge of 6+. Three oxide ions have a charge of 6–. A compound with a 2:3 ratio of aluminum and oxygen will be neutral.

12.



The chlorine, carbon, and fluorine atoms each have eight electrons in their valence shells, meaning they have full valence shells. The hydrogen atoms each have two electrons in their valence shells, meaning they have full valence shells.

13. A bond dipole is a bond between atoms that has a slightly positive end and a slightly negative end resulting from an unequal attraction of the atoms for electrons.

14. Boiling point depends on a balance between the kinetic energy of the particles and the strength of the attractive forces among the particles.

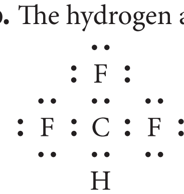
15. Dipole-dipole forces and weak attractive forces make up intermolecular forces.

16. Ionic compounds can conduct electric current in liquid form or when dissolved in water.

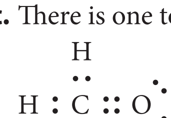
17. a. 1:2 **b.** 1:1 **c.** 1:1 **d.** 1:1

18. a. By itself, an oxygen atom has six electrons in its outer shell. With the atoms, electrons, and bonding shown, the oxygen atom has only six valence electrons—two bonding pairs and one lone pair. If it is sharing two pairs of electrons with carbon, the oxygen atom should have eight valence electrons. This gives it a stable, neutral configuration.

b. The hydrogen atom has too many electrons.



c. There is one too many electron pairs on the oxygen.



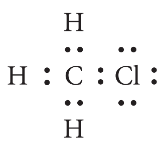
19. a. magnesium chloride

b. sodium oxide

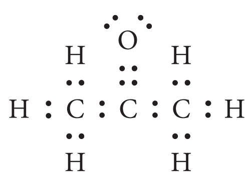
c. iron(III) chloride

d. copper(II) oxide
e. barium hypochlorite
f. ammonium nitrate
g. aqueous hydrogen chromate, or chromic acid
h. trihydrogen phosphate
i. potassium hydroxide
j. cadmium hydroxide

20. a. $AuCl_3$ **f.** $Ca(ClO)_2$
b. MgO **g.** $HCl(aq)$
c. $LiNO_2$ **h.** $H_2SO_4(aq)$
d. Ca_3P_2 **i.** $Co(OH)_2$
e. MnS **j.** $LiOH$

21. a.

c.

b.

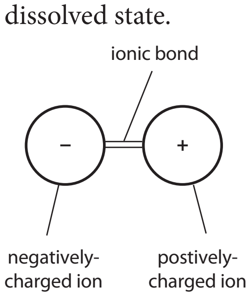
d.


22. a. sulfur dioxide **c.** carbon monoxide
b. dinitrogen tetroxide **d.** dichlorine oxide

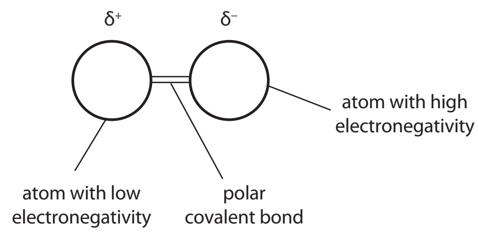
23. a. H_2O **b.** SO_3 **c.** $SiCl_4$

24. a. four formula units of potassium bromide
b. The compound $NaHSO_4(s)$ is sodium hydrogen sulfate.
c. The compound $KNO_2(s)$ is potassium nitrite.

25. Sample answer:
Ionic bonds: Compounds with ionic bonds have very high melting and boiling points. This is because ionic bonds are very strong and need to be broken in order to melt or boil. This requires a great deal of energy. Compounds with ionic bonds conduct electricity when they are in liquid or dissolved form, since they consist of charged particles that can move when in a liquid or dissolved state.

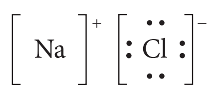


Polar covalent bonds: Compounds that have asymmetrical molecules with polar covalent bonds tend to have melting and boiling points in the intermediate range. The polar covalent bonds result in the molecule having negative and positive poles. In a compound, these poles attract one another. The attractive forces are much weaker than ionic bonds, but stronger than the weak intermolecular forces that attract non-polar molecules. Such compounds also tend to dissolve well in water because their dipolar molecules are attracted to water molecules, which are also dipolar.



26. Sample answer: Hydrochloric acid, $HCl(aq)$, is a highly corrosive and dangerous acid, and sodium hydroxide, $NaOH(aq)$, is a highly corrosive and dangerous base. When these are reacted in equal molar amounts, however, the acid and base are neutralized and form the very safe compounds sodium chloride, $NaCl(aq)$, and water, $H_2O(\ell)$.

27. Sample answer: A sodium atom has one electron in its valence shell, which it loses to other atoms when it forms chemical bonds.

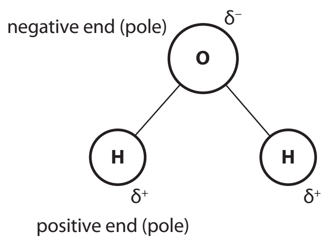


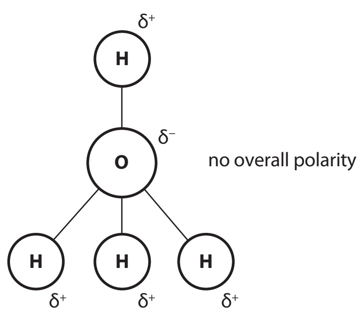
28. a. mostly ionic; $\Delta EN = 3.2 - 1.0 = 2.2$
b. polar covalent; $\Delta EN = 3.4 - 2.6 = 0.8$
c. non-polar covalent; $\Delta EN = 3.0 - 3.0 = 0.0$
d. polar covalent; $\Delta EN = 3.2 - 1.9 = 1.3$

29. Sample answer: I agree. Molecular compounds are made up of molecules. The molecules are held together internally by very strong covalent bonds. But in solids and liquids, they are attracted to one another by intermolecular forces. In gases, the molecules travel in straight lines and collide with one another and experience practically zero attraction for one another. If there were no intermolecular forces, there would be nothing to attract the molecules to one another and all molecular compounds would behave as gases.

30. Boiling points of molecular compounds or elements depend on the sizes of the particles involved and on the strength of the attractions among the particles. In a case like argon and fluorine, when the size of the particles is about the same, the attractive forces become the deciding factor. The boiling points of argon and fluorine are similar, meaning that the forces that attract argon atoms to one another and the forces that attract fluorine molecules to one another are also similar. In both cases they are weak because the molecules of fluorine, just like the argon atoms, have a uniform charge distribution. In other words, neither has an end that is consistently more positive or negative than the other end. The only thing attracting the molecules and atoms is a weak attraction of the electrons in one atom or molecule for the protons in another atom or molecule.

31.

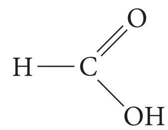




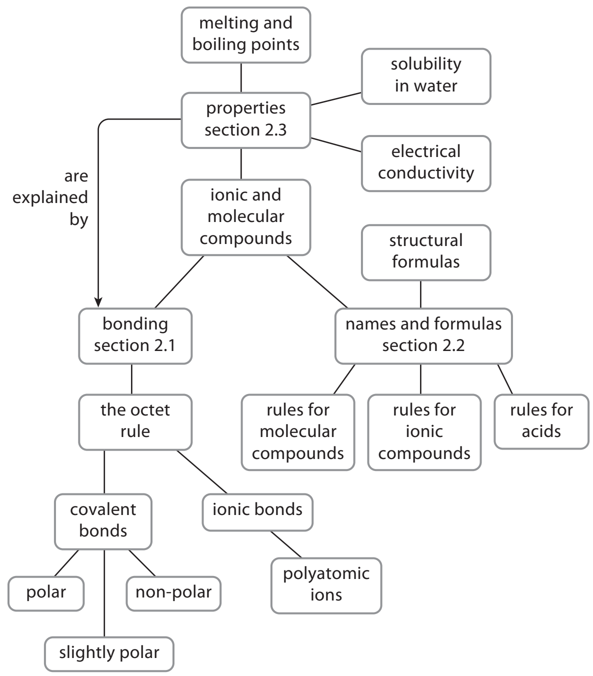
Both methane and water contain polar covalent bonds, but a methane molecule is symmetrical in all directions, while a water molecule is not. Hence, methane is a non-polar molecule, while water has a strong polarity. This means that water molecules are held together by strong dipole-dipole forces (hydrogen bonds, in this case), while methane is held together by weak intermolecular forces. The difference in strength of the intermolecular forces in these two substances accounts for the difference in boiling points.

32. I^-: iodide; IO^-: hypoiodite; IO_2^-: iodite; IO_3^-: iodate; IO_4^-: periodate. Students' naming systems should include suffixes and/or prefixes that can be added or changed to indicate up to four different numbers of oxygen atoms in the ion.

33. Both the Lewis structure and structural formula show the atoms in the molecule and how they are bonded to one another. The lines in the structural formula represent electron pairs involved in bonding. Lone pairs are usually not shown in structural formulas.



34. Sample answer:



35. Designs should include safety precautions and involve the melting points of the two solids. The solid that melts at temperatures achievable in a high school laboratory would be the molecular compound. The ionic compound would not melt under the same conditions.

36. The boiling point of water is 100ºC. The boiling point of methanol is 64.7°C. To separate a mixture of methanol and water, you could heat the mixture to a temperature just above the boiling point of methanol (e.g., 65°C) and gather the gas that evaporates. The water would remain behind in the original vessel.

37. Answers should involve testing the conductivity of the solutions. The solution with the ionic compound can be expected to conduct electricity, while the solution with the molecular compound would not.

38. Sample answer: Challenges: (1) designing a container that could withstand the high temperature needed to melt sodium chloride and (2) dealing with the extremely toxic and corrosive chlorine gas produced in the process. Solution (1) use a heat-resistant material such as graphite or porcelain. Solution (2) equipment could be airtight and the chlorine gas could be collected or reacted with some chemical to form a less harmful compound.

39. The accomplishment was impressive because fluorine is extremely reactive and therefore difficult to isolate—it tends to react with other elements to form compounds. From its position on the periodic table you can see that it has a very high electronegativity. Part of the challenge of working with fluorine was also the danger: fluorine itself is a highly toxic and corrosive gas, and it was isolated from hydrofluoric acid, a dangerously corrosive acid.

40. Sample answer: The detergent molecules surround particles of oil and grease, forming a "blob" in which the interior is non-polar (the oil or grease and non-polar end of the detergent molecule) and the exterior is polar (the polar end of the detergent molecule). Because its exterior surface is polar, the detergent-grease blob dissolves in water and is rinsed away.

Chapter 2 Self-Assessment Questions

(Student textbook pages 92–3)

1. b

2. e

3. b

4. a

5. e

6. d

7. b

8. c

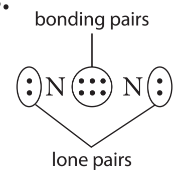
9. b

10. c

11. The octet rule states that when bonds form between atoms, the atoms gain, lose, or share electrons in such a way that they create a filled outer electron shell containing eight electrons. Sulfur has six valence electrons, and therefore needs to gain two electrons to satisfy the octet rule. It can do this by bonding with two sodium atoms, each of which will give up one electron to satisfy the octet rule.

12. The bond between the carbon atoms must be a triple bond because that is the only way to satisfy the octet rule with the given structure. Each carbon atom shares one electron with the hydrogen atom it is bonded to. This leaves each carbon atom in need of three more shared electrons to achieve a stable octet, which it gets by forming a triple bond with the other carbon.
H:C:::C:H

13.



14.

$2(+3) + 3(-2) = 6 - 6 = 0$ $3(+2) + 2(-3) = 6 - 6 = 0$

Sample answer: I would use this formula-writing method when the charges of the ions in an ionic compound are not the same.

15. a. 1.4; polar covalent
b. 0.4; slightly polar covalent
c. 0.0; non-polar covalent
d. 1.8; mostly ionic

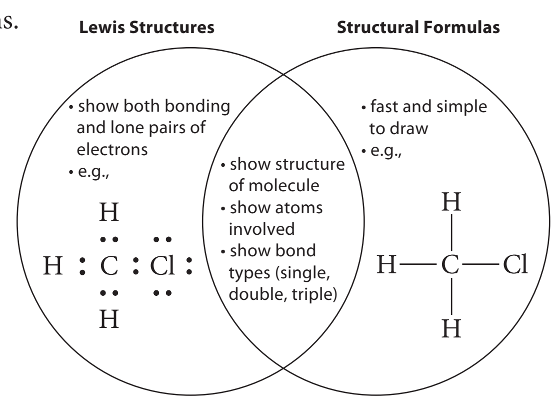
16. a. magnesium phosphate
b. sodium iodate
c. aluminum phosphate
d. sodium hydrogen carbonate

17. a. KSCN **b.** YCl_3 **c.** Fe_2S_3 **d.** SnF_2

18. a. trisilicon tetranitride
b. phosphorus pentachloride
c. sulfur hexafluoride
d. chlorine trifluoride

19. a. SO_3 **b.** CO **c.** Se_2Br_2 **d.** NI_3

20. Sample answer: I would use a structural formula instead of a Lewis structure if I needed a quicker and less cumbersome way to show the connections between atoms.



21. Sample answer: Scandium oxide has an extremely high melting point and is therefore likely to be an ionic compound, meaning its particles are joined by ionic bonds. Nitrogen trichloride has a relatively low melting point. As a solid, its particles are likely joined by dipole-dipole forces or weak intermolecular forces. Ethane has a very low melting point. As a solid, its particles are likely joined by weak intermolecular forces.

Note that these answers are based on melting point alone. If students think about the identity of the substances they will realize that scandium oxide is ionic, nitrogen trichloride (ammonia) is molecular with polar molecules, and ethane is molecular with non-polar molecules.

22. Sample answer: It is true that non-polar molecules have no distinct separation of charge in the same way that dipoles do. But non-polar molecules do have negatively-charged electrons and positively charged protons. It's possible for the electrons in one molecule to be attracted to the protons in the nucleus of an atom of another molecule. This results in weak attractive forces.

23. chlorine > methanol > potassium oxide. The ionic compound (potassium oxide) is highest. The lowest is the one with non-polar molecules (chlorine). Leaving methanol in the middle.

24. The negatively charged pole of one polar molecule is attracted to the positively charged pole of another. This leads to stronger (dipole-dipole) intermolecular forces and gives polar molecules higher boiling points than non-polar ones. Sketches should look like Figure 2.27 on page 79 of the student textbook.

25. A substance can conduct an electric current only if charges (electrons or ions) can move independently of one another.

Unit 1 Review Questions

(Student textbook pages 98–101)

1. d

2. c

3. d

4. d

5. c

6. a

7. e

8. d

9. b

10. c

11. Thomson discovered the presence of electrons in atoms. Rutherford discovered that the atom had a dense central nucleus that was positively charged.

12. Valence electrons are the electrons in the outermost energy level of an atom. The way in which an atom reacts with other atoms is determined by the number of valence electrons. Atoms react according to the octet rule, which states that when bonds form between atoms, the atoms gain, lose, or share electrons in such a way that they create a filled outer shell containing eight electrons.

13. Isotopes of an element have the same number of protons and electrons but have different numbers of neutrons. The isotopes have identical bonding patterns because they have the same electron arrangements.

14. Metals tend to have high electrical conductivities, and non-metals tend to have very low electrical conductivities.

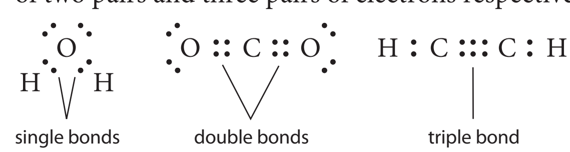
15. O, C, Ge, Ca, Ba

16. Cl, P, Mg, Ca, K

17. Electronegativity increases as effective nuclear charge increases. Electronegativity is a measure of the attraction a nucleus has on shared electrons. As effective nuclear charge increases, the nucleus of the atom would have a stronger attractive force on shared electrons.

18. An oxygen atom can satisfy the octet rule by gaining two electrons to form an ionic bond, by sharing electrons by forming single bonds with two atoms, or by forming a double bond with one other atom.

19. A single covalent bond results from the sharing of one pair of electrons between two atoms. A double covalent bond and a triple covalent bond result from the sharing of two pairs and three pairs of electrons respectively.



20. a. 2.1; mostly ionic

b. 0.8; polar covalent

c. 0.0; non-polar covalent

21. Sample answer: A metal and a non-metal or polyatomic anion, or a polyatomic cation and a metal or non-metal, are bonded together to make an ionic compound. Here are three examples: potassium bromide, $KBr(s)$, magnesium sulfate, $MgSO_4(s)$, and ammonium chloride, $NH_4Cl(s)$.

22. (A) ionic, (B) non-polar covalent, (C) polar covalent

23. Metalloids are located between metals and non-metals on the periodic table. The location reflects the fact that metalloids share some of the properties of both metals and of non-metals.

24. The term *stable octet* refers to an atom or ion having eight electrons in its outermost energy level, a stable state.

25. If the scientist makes water that contains a radioisotope of oxygen, then the energy given off by the radioisotopes could be monitored in the different products that form.

26. The number of electrons in metal atoms decreases and the number of electrons in non-metal atoms increase when the elements form an ionic compound.

27. If a substance has weak intermolecular forces it would likely be a gas at room temperature. As intermolecular forces become stronger, the molecules are held more tightly together, so the substance becomes more likely to be in the liquid or solid state at room temperature.

28. In the solid state, the ions that make up the ionic compound are not free to move past one another, so the charged particles cannot conduct an electric current. However, in the liquid state, the ions can move past one another. Charges that are free to move can conduct an electric current.

29. Particles Contained in Isotopes

Isotope	Atomic Number	Mass Number	Number of Protons	Number of Neutrons
$^{44}_{20}Ca$	20	44	20	24
$^{20}_{10}Ne$	10	20	10	10
$^{14}_{6}C$	6	14	6	8
$^{37}_{17}Cl$	17	37	17	20
$^{28}_{12}Mg$	12	28	12	16
$^{66}_{30}Zn$	30	66	30	36
$^{138}_{56}Ba$	56	138	56	82

30. As atomic number increases within a period, electronegativity increases and atomic radius decreases.

31. **a.** Ca_3N_2 **b.** SI_2 **c.** $PbBr_2$ **d.** $AlPO_3$

32. **a.** nitrogen trichloride **c.** iron(II) oxide
b. potassium carbonate **d.** dinitrogen tetroxide

33. **a.** silicon **b.** sulfur **c.** phosphorus

34. **a.** nitric acid, aqueous hydrogen nitrate
b. aqueous hydrogen iodide, hydroiodic acid
c. aqueous hydrogen oxalate, oxalic acid
d. cobalt(III) hydroxide

35. **a.** HClO(aq) **b.** NH_4OH **c.** HNO_2(aq) **d.** $Mg(OH)_2$

36. **a.** Ionic, since electrons are transferred from Mg to Cl.
b. Mg:2, Cl:7 **c.** Mg:0, Cl:8

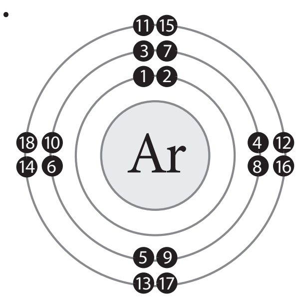
37. The strong force of attraction between the opposite charges of the ions that make up an ionic compound is a result of the charges of the ions. The slightly charged ends of polar water molecules have a force of attraction for the ions. The molecules of a non-polar compound have no permanent dipoles, so there is little attraction between them and the water molecules. As a result, ions are separated and ionic compounds dissolved in water more easily than non-polar molecular compounds are.

38. Sample answer: I would expect NaCl to have the higher melting point because it is an ionic compound and BrCl is a molecular compound. Ionic compounds tend to have very high melting points.

39. **a.** 2+ **b.** 2−; 0 **c.** 1−; 0

40. **a.** Because it is non-conductive in a solid state but conductive in an aqueous solution, it must be an ionic compound.
b. The compound in the solid state shows no electrical conductivity. The compound dissolved in water shows good electrical conductivity.
c. When the ionic compound is dissolved in water, the ions are free to move and are able to conduct electric current, but when the compound is in the solid state, the ions are unable to move and no current can be conducted.

41.



42. The Lewis structures should show six electrons for oxygen, seven electrons for fluorine, eight electrons for neon, and one electron for sodium. The charges of the ions are 2− for oxygen, 1− for fluorine, 0 for neon, and 1+ for sodium. The explanation should point out that the atoms gain or lose electrons to achieve the same number of electrons as the neon atom has.

43. Sample answer: First, I would determine the atomic number of carbon (6) by looking on the periodic table. The atomic number is the number of protons in the atom. It also equals the number of electrons because the atom is neutral. To determine the number of neutrons, I would subtract the atomic number (the number of protons) from the mass number.

44. The sketch should show the following trends:
- increasing from left to right across a period — atomic number, electronegativity, electron affinity (becoming more negative), and ionization energy
- decreasing from left to right across a period — atomic radius
- increasing from top to bottom down a group — atomic number and atomic radius
- decreasing from top to bottom down a group — electronegativity, electron affinity (becoming less negative), and ionization energy

45. The compound formed from sodium and chlorine has ionic bonds, which act to hold the ions tightly together and cause the compound to have a high melting point. The compound formed from bromine and chlorine has covalent bonds. Although the bonds are strong, the forces between the molecules are not as strong as ionic bonds, so the compound has a much lower melting and boiling point and is a gas at room temperature.

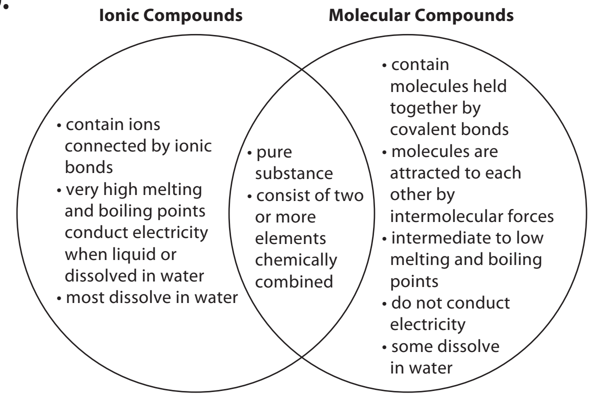
46. The script should clearly demonstrate understanding of the proper use of the product and should clearly communicate safety precautions.

47. The term *covalent* describes putting valence electrons together. In a covalent bond, electrons are shared by two atoms, so the atoms are putting their electrons together in order to make a bond. The two electrons are in the valence shell of both atoms.

48. Sample answer: A compound is composed of elements that join in a specific ratio. If the elements join in a different ratio, then a different compound is formed. The components in a mixture have no specific ratio and can be mixed in any quantities desired.

49. Sketches should show the difference in the pulls of the atoms that share the electrons in each bond. The electrons that are shared in a non-polar covalent bond are pulled equally by both atoms and are shared equally. The electrons in a slightly polar covalent bond are pulled only slightly more by one of the atoms. The electrons in a polar covalent bond are pulled much more strongly by the atom with the higher electronegativity, so the electrons are shared unequally between the atoms.

50.



51. Sample answer: tricycle, *tri*–, 3; octopus, *oct*–, 8; pentagon, *penta*–, 5

52. Blog entries should include several exceptions to the rules given on pages 65, 67, 71, and 72. (Important exceptions are included within the rules.) Examples include:
- Not writing the subscript "1" when writing formulas
- Writing halogens before oxygen in the names and formulas of binary molecular compounds
- Names of organic compounds

53. Sample answer: Statement A is valid because a molecule can be polar only if it has at least one polar bond. The polar molecule HBr is an example of a single polar bond causing a molecule to be polar. However, statement B is not valid because a molecule can be non-polar if it contains only non-polar bonds or if it contains a symmetrical arrangement of polar bonds such that the dipoles cancel one another in the molecule. Carbon dioxide is a non-polar molecule even though each bond between the carbon atom and an oxygen atom is polar.

54. Procedures should describe appropriate safety cautions (wear goggles and apron, electrical safety) and materials. The procedure should describe using a conductivity tester on the solid sample, placing the same amount of each substance into equal volumes of water and stirring to test for solubility, and then testing each solution with the conductivity tester. Evidence for an ionic compound would be non-conductivity in the solid state, high solubility in water, and good conductivity when in solution. Evidence for a molecular compound would be non-conductivity in the solid state, low solubility in water, and non-conductivity for the solution.

55. **a.** B, D, C, A
b. (B) Dalton, (D) Thomson, (C) Rutherford, (A) Bohr
c. Sample answer: (B) The atom is indivisible, and atoms differ in size. (D) The atom is a mass of positive charge with negative electrons embedded in it. The atom is divisible. (C) The atom is composed of a small, dense, massive, positively charged nucleus with electrons in orbit around it. (A) Electrons can exist only in specific energy levels and so are limited in the orbits they can follow.
d. Sample answer: Schrodinger's wave equation model calculated the shapes of electron clouds and describes the distance from the nucleus where electrons spend most of their time.

56. Sample answer: Colourful plastic models of plant and animal cells, with removable parts, helped me to visualize the different organelles and remember their structures, names, and functions.

57. 50.94 u

58. a. the P should have 5 dots around it; add 3 electrons for a 3− charge
b. the O should have 7 dots around it in opposite pairs; add 2 electrons for a 2− charge
c. •Sr•; remove 2 electrons for a 2+ charge
d. Li•; remove 1 electron for a 1+ charge

59. a. one aluminum to three chlorine
b. one aluminum to one nitrogen
c. two aluminum to three oxygen

60. a. molecular compound
b. molecular compound
c. ionic compound

61. NH_3, since N ($EN = 3.0$) has higher electronegativity that P ($EN = 2.2$).

62. Answers should reflect chemical names from three products identified as either ionic or molecular. Formulas should be written for any compounds for which students have learned rules.

63. Sample answer: It is beneficial to convert compounds that are insoluble in water to compounds that are soluble in water because compounds that are insoluble in water tend to be soluble in fat and accumulate in the body. This is a problem if the compound is toxic. Compounds that are soluble in water can be more easily excreted from the body.

64. a. The molecule has a "bent" or "V" shape.
b. An electronegativity difference of 0.85 shows that the bonds are polar covalent bonds. Because the molecule's shape is not symmetrical around the central atom, the polar bonds are not able to cancel one another's pull. As a result, the molecule cannot be non-polar.
c. Liquid: Like the molecule shown, a water molecule is not symmetrical around its central atom and has polar covalent bonds. Therefore the molecule is a dipole and experiences dipole-dipole forces, which tend to correspond to lower boiling and melting points.

65. Ovens and drains are often covered in fatty deposits. Bases react readily with fat, converting them into substances that can be washed away with water. A less corrosive chemical that is still basic is baking soda. A weak base, it would be less effective and require more scrubbing than a strong base.

66. a. The bonds are polar with the carbon atom being slightly positive and the chlorine atom being slightly negative in each bond.
b. The molecule is non-polar because the shape allows the polarity of the bonds to cancel one another.
c. The non-polar nature of the molecules results in weak intermolecular forces, so the compound would likely not be solid at room temperature.

Unit 1 Self-Assessment Questions

(Student textbook pages 102–103)

1. a
2. b
3. a
4. c
5. d
6. e
7. d
8. d
9. b
10. a

11. The model should have a letter *P* in the center with three rings around it. The electrons are represented by dots with two on the innermost ring, eight on the second ring, and five on the outer ring. The outer ring should show a pair and three unpaired electrons. The number of rings represents the energy levels or shells, so phosphorus is in period 3. The number of valence electrons in the outermost ring places phosphorus in Group 15.

12. Each isotope of an element has a different number of neutrons and therefore has a different mass. They react in similar ways in chemical reactions. Isotopes of the same element are unstable and radioactive and will decay and change into a more stable nucleus.

13. 107.90 u

14. a. Each is a gas at room temperature. Each is a non-metal.

b. $\overset{\bullet}{\text{H}}$ $\overset{\bullet\bullet}{\text{He}}$

Helium atoms have a full valence shell with two electrons, so helium is non-reactive. However, hydrogen atoms have one valence electron. Hydrogen is reactive because its atoms tend to lose or gain one electron to empty or complete the first energy level.

15. Diagrams should show that atoms get smaller from left to right across a period. As atomic radius decreases, the nucleus of an atom is closer to the outer electrons of a second atom to which it is bonded. The shorter distance causes a larger force of attraction, which is reflected in a greater electronegativity from left to right across a period.

16. a. Carbon is a non-metal. Silicon and germanium are metalloids. Tin and lead are metals.

b. Atomic radius increases from top to bottom within a group. As the atoms become larger, the valence electrons are farther from the nucleus and the attractive force between them and the nucleus decreases. Because the electrons are held less tightly, they are easier to remove. Thus, the elements develop more metallic character as you move down Group 14.

17. Atoms of metals tend to lose electrons when they react, so the most-reactive metals will give up electrons the most easily. The attractive force on a valence electron of an atom decreases as the atom's radius increases and as shielding effects increase. Atomic radius increases as you move down a group on the periodic table and the additional electrons between the nucleus and the valence electron result in greater shielding effect. The outermost electron experiences less force as a result, so ionization energy decreases as you move down a group on the periodic table. Metals become more reactive as you move down a group because it takes less energy to remove an electron.

18. Each of these nuclei is unstable because it has too many neutrons.

a. lose 2 electrons
b. gain 2 electrons
c. lose 1 electron
d. gain 3 electrons

19. a. ionic; Mg_3N_2
b. covalent; OF_2
c. ionic; $SnBr_2$
d. ionic; $AlPO_4$
e. ionic; $Co_2(SO_3)_3$

20. a. covalent; phosphorus pentachloride
b. ionic; lithium carbonate
c. ionic; copper(II) oxide
d. covalent; dinitrogen trioxide
e. ionic; ammonium nitrite

21. Flowcharts should present a logical sequence of steps to identify the compound as either ionic or molecular, to determine prefixes for a molecular compound, and to identify a multivalent metal and its valence.

22. Prefixes are used in the names of binary molecular compounds to avoid ambiguity because two non-metals often can join in more than one ratio and prefixes are needed to state how many atoms of each element should be represented for a given formula. Roman numerals are used to avoid ambiguity in the names of ionic compounds because multivalent metals can have more than one valence, and without identifying which valence the metal has in a particular compound, the name will not correctly identify the compound.

23. Sample answer:

Properties of Ionic and Molecular Compounds

Substance type	Melting point	Electrical conductivity	State at room temperature	Solubility in water
Ionic	very high	conducts electricity when in liquid state or when dissolved in water	solid	many are soluble
Molecular	low	does not conduct electricity	solid, liquid, or gas	tend to be soluble if polar molecules; tend to be insoluble if non-polar molecules

24. Sample answer: Heat a small amount of each substance on a hot plate. The table salt will show no change because it has a high melting point. Dissolve a small amount of each substance in separate samples of water. Test each mixture using a conductivity tester. The salt dissolved in water will conduct an electric current and the sugar will not.

25. a.

b. The C–H bond has an electronegativity difference of 0.35, so it is slightly polar covalent. The hydrogen atom has partial positive charge, and the carbon atom has partial negative charge. The C–Cl bond has an electronegativity difference of 0.61, so it is polar covalent. The carbon atom has partial positive charge, and the chlorine atom has partial negative charge.

c. The single C–Cl bond causes chloromethane molecules to be polar, while the four identical C–H bonds cause methane molecules to be non-polar. Thus, the dipole-dipole forces that exist between chloromethane molecules are stronger than the dispersion forces that exist between methane molecules. The stronger intermolecular forces cause the melting point and boiling point of chloromethane to be higher than those of methane.