15. The two exceptions are copper and chromium.

The expected diagram for chromium would have a filled $4 s$ orbital and four $3 d$ orbitals containing one electron each (the fifth 3d orbital would be empty). The expected diagram for copper would have a filled 4 s orbital, four filled $3 d$ orbitals, and one $3 d$ orbital with one electron.
The actual diagram for chromium has one electron in its $4 s$ orbital and one electron in each of its five $3 d$ orbitals. The actual diagram for copper has one electron in its $4 s$ orbital and five filled $3 d$ orbitals. The discrepancy arises because the predicted electron configurations have a slightly higher energy than the actual electron configurations. When the $3 d$ orbitals are either all half-filled or all completely filled, the configuration becomes stable. The predicted electron configurations are just 1 electron off this energetically more favourable situation.
16. The configuration for the valence electrons, $s^{2}$, indicates that strontium is in Group 2.
17. a. nickel: $[\mathrm{Ar}] 4 s^{2} 3 d^{8}$
b. lanthanum, La: [Xe] $6 s^{2} 5 d^{1}$
18. The value of 4 in $4 s^{2}$ indicates that titanium is in Period 4.
19. The configuration of the valence electrons shows full $s$ and $d$ orbitals, but a half-filled $p$ orbital, so arsenic must be in the $p$ block.
20. a. The electron configuration of the valence electron is $s^{1}$, so francium must be in Group 1 and belong to the $s$ block. Because of the 7 in $7 s^{1}$, francium must be in Period 7 .
b. The electron configuration of the valence electrons is $s^{2} f^{4} d^{4}$, so tungsten must be in Group $6(2+4)$ and belong to the $d$ block. Because of the 6 in $6 s^{2}$, tungsten must be in Period 6.
c. The electron configuration of the valence electrons is $s^{2} d^{10} p^{3}$, so antimony must be in Group $15(2+10+$ 3 ) and belong to the $p$ block. Because of the 5 in $5 s^{2}$, antimony must be in Period 5.

## Answers to Chapter 3 Review Questions

## (Student textbook pages 199-203)

1. b
2. a
3. c
4. a
5. e
6. d
7. a
8. a
9. b
10. d
11.e
11. a
12. c
13. d
14. In both models atoms were spherical and solid (no empty space) and were neutral-particles. As well, in both models, atoms were considered the building blocks of all matter.
Thomson's model contained internal structures and discrete structures that were smaller than the atom itself. Specifically, Thomson's model had negatively charged subatomic particles, now known as electrons, embedded in a continuous mass of positively charged matter.
Dalton's atom was the smallest possible particle of matter. It was indivisible, meaning that it could not be broken down into smaller particles.
15. Thomson's and Rutherford's models of the atom were spherical and contained a sub-atomic particle believed to be the same particle in all matter (the electron). Both models included equal amounts of positively and negatively charged matter to create an overall neutral atom.
In Rutherford's model, the positively charged matter was contained in a very small region at the centre of the atom which he called the nucleus. Most of the matter of the atom was contained within this nucleus. Negatively charged electrons orbited the nucleus. Most of the atoms consisted of empty space.
Thomson's model suggested that the positive and negatively charged matter was evenly distributed and no empty space existed within the atom.
16. In all atomic models, the atom is the building block of all matter, atoms are neutrally charged, and atoms are spherical.
17. The speed of all electromagnetic radiation, in a vacuum, is the same, $3 \times 10^{8} \mathrm{~m} / \mathrm{s}$.
18. Bohr's electron moves in a circular pathway a fixed distance from the nucleus-its position is known. The electron in the quantum mechanical model moves constantly within a region of space that is described by using wave equations. Its exact position cannot be known for certain.
19. By using the basic laws of physics and Planck's concepts, Bohr was able to calculate the energy and radii of the allowed orbits for the hydrogen atom.
20. Groups 1 and 2 metals have low electron affinity and low ionization energies and therefore lose electrons easily, forming positive ions. These positive ions are then attracted to negative ions to form ionic compounds.
21. There are three equal-energy $p$ orbitals. According to Hund's rule, single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. When the three electrons in the $p$ orbitals of nitrogen have occupied the $p$ orbitals, there are no additional electrons to pair with the single electrons in the $p$ orbitals.
22. The exact position of an electron in an atom cannot be known with certainty. The energy of electrons in any electron shell and the relative size of the orbital are known with certainty as defined by the principal quantum number $n$.
23. Quantum numbers provide information about the energy, size, shape, spatial orientation, and the relative direction of the axis of the electron of the orbital.
24. The principle quantum number provides information about the energy and size of the orbital.
25. The electron in a higher energy level would have more energy to overcome the attractive pull of the positive charge of the nucleus and therefore be able to venture farther away from it.
26. The Pauli exclusion principle states that only two electrons of opposite spin could occupy an orbital. Each electron in an electron pair within the same orbital are identical but have opposite spins-the spin quantum number is the same number for any electron, but the sign is either $+($ spin up) or - (spin down).
27. Hund's rule states that the lowest energy state for an atom has the maximum number of unpaired electrons allowed by the Pauli exclusion principle in a given energy sublevel. This rule can also be stated as single electrons with the same spin must occupy each equal-energy orbital before additional electrons with opposite spins can occupy the same orbitals. This rule determines the order of the filling of orbitals.
28. An atom's chemical properties are mainly associated with its ground state electron configuration. As well, there are a tremendous number of possible excited states.
29. The atomic radius decreases across a period (from left to right) and increases down a group. The reason that it decreases across a period is because the amount of charge in the nucleus becomes greater across a period. Thus the attractive force on the electrons becomes greater, pulling the electrons closer to the nucleus. The radius increases down a group because each successive group has one more energy level, or shell, than the previous one and thus electrons are farther from the nucleus.
30. None of the experiments described in the book would have remained the same because the tests depend on the charge carried by the particles.
31. The element described, would be in Period 6, $d$ block, and Group 12.
32. $[\mathrm{Kr}] 5 s^{2} 4 d^{10} 5 p^{3}$
33. $3 s, 3 d, 4 p, 5 s, 5 d$
34. The highest possible value of $l$ is 3 (because there are electrons in a $d$ orbital). Therefore, $m_{l}$ can be anything from -3 to +3 .
35. nickel, Ni
36. aluminum, Al
37. tellurium, Te
38. chlorine, Cl
39. ruthenium, Ru
40. oxygen, O
41. barium, Ba
42. In these cases, the $d$ electrons do not typically participate in chemical reactions. As well, all $p$-block elements in Period 4 and above have $10 d$ electrons.
43. Boron will be larger than fluorine.
44. Magnesium will be larger than silicon.
45. Calcium will be larger than selenium.
46. a. Electron configuration: $[\mathrm{Rn}] 7 s^{2} 5 f^{4} 6 d^{10} 7 p^{6}$
b. Quantum number for last electron: $n=7, l=1$, $m_{l}=1, m_{s}=-\frac{1}{2}$
c. Physical state at room temperature: gas
d. Reactivity: inert, just like all other noble gases
47. a. Na, Si, Ar: First ionization energy tends to increase across a period. Sodium, silicon, and argon are all in Period 3.
b. $\mathrm{Ba}, \mathrm{Ca}, \mathrm{Mg}$ : First ionization energy tends to decrease down a group. Barium, calcium, and magnesium are all in Group 2.
c. $\mathrm{Li}, \mathrm{Be}, \mathrm{He}$ : Helium would have the highest first ionization energy because it is the farthest right in Period 1. Lithium would have the lowest first ionization energy because it is to the left of beryllium in Period 2.
48. a. Oxygen: There are eight electrons in the neutral atom, thus there must be eight protons in the nucleus. Oxygen has the atomic number eight.
b. The electron configuration indicates the values of the quantum numbers $n$ and $l$.
c. The electron configuration does not provide information about the magnetic and spin quantum numbers, $m_{s}$ and $m_{l}$.
49. Distribute BLM G-23 Risk/Benefit Analysis to assist students in organizing their thinking. Answers may be in the form of a chart, a Venn diagram, or risk/ benefit analysis and should include a clear and succinct description of the technology and its use, plus a brief summary of each cost/risk and benefit. Sources of information should also be included.
50. Answers should follow the outline of the question and include descriptions of:

- how X rays can be released in cathode ray tubes
- the precautions that are taken to prevent the unguarded release of X ray radiation from the CRT computer monitor
- any other precautions or warnings that are necessary that extend from the use of a cathode ray tube in a computer monitor.
Ideas that could be further developed include:
- discovery of X rays and early use as an experimental tool
- construction of first CRT, how X rays are produced
- work of J.J. Thomson, Henry Moseley and others
- parts of CRT used in television picture tubes and oscilloscopes
- dangers for those working with CRT, precautions to take
- disposal problems, environmental concerns

Sources of research information should be cited.
52. Venn Diagram:

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

53. Students might suggest such examples as twisting balloons together, balloons of one colour within balloons of a different colour, etc.
54. The condensed electron configuration for bromine is $[\mathrm{Ar}] 4 s^{2} 3 d^{10} 4 p^{5}$.
The [Ar] represents the electron configuration for argon, which would be $1 s^{2} 2 s^{2} 2 p^{6} 3 s^{2} 3 p^{6}$. If you put this electron configuration in place of the [Ar], you would have the complete electron configuration for bromine. Because all of the occupied energy levels in argon and all other noble gases are completely filled, these electron configurations are present in the electron configuration of the elements in the next period. Therefore, for any given period, you can indicate the electron configuration of all of the filled energy levels by writing the symbol for the noble gas in square brackets.
55. The effective nuclear charge is the net positive charge that an electron experiences from the nucleus. It is equal to the nuclear charge less the shielding or screening from the inner core of electrons.
Comparing a Mg atom $(2,8,2)$ with a S atom, $(2,8,6)$, the outer valence electron of Mg experiences a nuclear charge of $12+$ and a shielding from 8 inner electrons. On the other hand, the valence electron in an atom of $S$ experiences a nuclear charge of $16+$ and a shielding of 10 inner electrons. The effective nuclear charge is greater in S and it is a smaller atom than Mg .
Estimates have been made of the shielding effect contributed by electrons in different orbitals that allow for a comparison of effective nuclear charge to be made between different types of atoms. These estimates reasonably agree with the differences in atomic radius between different types of atoms.
56. 


57.

| Scientist | Definition/Description (Formal) | $\begin{array}{c}\text { Key Concept for the quantum } \\ \text { mechanical model }\end{array}$ |
| :--- | :--- | :--- |
| Planck | $\begin{array}{l}\text { suggested that matter, at the atomic level, can absorb or emit only } \\ \text { discrete quantities of energy. Each of these specific quantities is called a } \\ \text { quantum of energy. }\end{array}$ | $\begin{array}{l}\text { the energy of an atom was quantized, } \\ \text { meaning that it can exist only in certain } \\ \text { discrete amounts. }\end{array}$ |
| Einstein | $\begin{array}{l}\text { asserted that light was also quantized. This meant that light occurs as } \\ \text { quanta of electromagnetic energy having particle-like properties. These } \\ \text { particle-like "packets" of energy were later called photons. }\end{array}$ | $\begin{array}{l}\text { photons of light are emitted in gas } \\ \text { discharge tubes to create discrete lines } \\ \text { in atomic spectra }\end{array}$ |
| Heisenberg | $\begin{array}{l}\text { Using mathematics, Heisenberg showed that it is impossible to know } \\ \text { both the position and the momentum of an object beyond a certain } \\ \text { measure of precision. According to this principle, if you can know an } \\ \text { electron's precise position and path around the nucleus, as you would } \\ \text { by defining its orbit, you cannot know with certainty its velocity. } \\ \text { Similarly, if you know its precise velocity, you cannot know with } \\ \text { certainty its position. }\end{array}$ | $\begin{array}{l}\text { Based on the uncertainty principle, } \\ \text { Bohr's atomic model is flawed because } \\ \text { you cannot assign fixed paths (orbits) to } \\ \text { the motion of electrons. }\end{array}$ |
| de Broglie | $\begin{array}{l}\text { as energy has matter-like properties (a photon of light), matter can have } \\ \text { wave-like properties }\end{array}$ | $\begin{array}{l}\text { developed an equation that enabled him } \\ \text { to calculate the wavelength associated } \\ \text { with any object - large, small, or }\end{array}$ |
| microscopic. |  |  |$]$

58. A graphic organizer such as a spider map, fishbone diagram, or a concept map should identify key concepts and relationships among the concepts shown in the Chapter 3 Summary on page 252 of the student textbook. Concepts relate to types of bonds.
59. a. The chemist would approach atomic theory from an experimental point of view, building a model to explain what is observed in an experiment. Mathematicians and physicists would look at the theoretical possibilities for an explanation; the mathematician by trying to quantify results in the form of mathematical equations and the physicist by looking at how the results can be explained by the fundamental laws of the universe.
b. The physicist would specifically look at the laws that describe the interactive forces between particles.
60. The Thomson model of the atom suggests that the atom is made up of matter that has an overall positive charge in which the negatively charged electrons are embedded for an overall neutral-charge atom-there is no discrete particle similar to the electron that contains a certain positive charge. Therefore, the Thomson model has a continuum of positive matter inside it. The Rutherford model claims that all of the positive charge of a nucleus is located in a discrete nuclear core and the negative charges of the atom are found in the discrete particles called electrons moving outside the nucleus. Because the different types of matter are found in discrete structures, the matter can be referred to as "quantized."
61. a. - all the positive charge was concentrated in a central nucleus

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

62. Bohr's model was very successful at predicting, matching, and explaining the spectral lines for hydrogen. However, it was very unsuccessful at predicting, matching, and explaining the spectral lines (especially clustering) in all multi-electron elements.
63. Answers should identify the health-care professional(s) interviewed and include answers to all parts of the question. The answer to question 3 should include some detail to explain it.
64. The electromagnetic spectrum of radiation is a continuum of energy that can be thought of in two ways: a wave of energy in which the frequency and wavelength are inversely related or as packets or photons of energy that are directed related to the frequency of the radiation. Ultraviolet light is one part of this continuum that lies just beyond the violet portion of the visible spectrum. This UV energy can excite electrons in some atoms, which, when they fall back to a lower energy level, emit visible violet light. Canada's new currency has molecules embedded in it that will react in this way under UV radiation. The receptors in the newer bills specifically respond to a narrow band of UV radiation (wavelength of 370 nm ) and emit a very faint white-purple light. This specificity of receptors to respond to a single wavelength makes it difficult to duplicate and therefore the bills are less likely to be counterfeited.
Older Canadian banknotes included dots that glowed blue under ultraviolet light. (These notes are being phased out as polymer notes are being introduced.)
65. 

Firework Colours

| Colour | Compound |
| :--- | :--- |
| Red | strontium salts, lithium salts |
| Orange | calcium salts |
| Green | barium salts, $\mathrm{BaCl}_{2}$ |
| Blue | copper compounds, $\mathrm{CuCl}_{2}$ |
| Purple | mixture of strontium (red) and copper (blue) <br> compounds |

## Source:

http://chemistry.about.com/od/fireworkspyrotechnics/a/ fireworkcolors.htm
66.

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

67. 

| H | He | Li | Be | B |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  |  | C |
| :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: | :---: |
| N | O | F | Ne | Na | Mg |  |  |  |  |  |  |  |  |  |  |  |  |  |  | AI | Si | P | S | Cl | Ar | K | Ca | Sc | Ti |
| V | Cr | Mn | Fe | Co | Ni |  |  |  |  |  |  |  |  |  |  |  |  |  |  | Cu | Zn | Ga | Ge | As | Se | Br | Kr | Rb | Sr |
| Y | Zr | Nb | Mo | Tc | Ru | Rh | Pd | Ag | Cd | In | Sn | Sb | Te | 1 | Xe | Cs | Ba | La | Ce | Pr | Nd | Pm | Sm | Eu | Gd | Tb | Dy | Ho | Er |
| Tm | Yb | Lu | Hf | Ta | W | Re | Os | Ir | Pt | Au | Hg | TI | Pb | Bi | Po | At | Rn | Fr | Ra | Ac | Th | Pa | U |  |  |  |  |  |  |

68. carbon, titanium, strontium, and erbium
69. [Uuo] $8 s^{2} 5 g^{7}$
70. 71. For main group elements, the last numeral of the group number is the same as the number of valence electrons.
1. The $n$ value of the highest occupied energy level is the period number in which the element is found.
2. The number of orbitals in an energy level is equal to $n^{2}$, and the maximum number of electrons allowed in an energy level is equal to $2 n^{2}$.
3. Each block of the periodic table contains twice as many groups as there are orientations of that type of orbital-this corresponds to the maximum number of electrons that can fill all the same type of orbitals within an energy level ( $2 \times 1$ groups in the $s$ block, $2 \times 3$ groups in the $p$ block, $2 \times 5$ groups in the $d$ block, and $2 \times 7$ groups in the $f$ block).
4. Carbon contains no $d$ electrons and lead does. The presence of $d$ electrons makes the electron configuration very complex. In the case of lead, the valence electrons form a "sea of electrons" that causes metallic bonding.
5. a.

b. Li to Na: $\Delta T=83^{\circ} \mathrm{C}$

Na to K: $\Delta T=34^{\circ} \mathrm{C}$
K to $\mathrm{Rb}: \Delta T=25^{\circ} \mathrm{C}$
Rb to Cs: $\Delta T=11^{\circ} \mathrm{C}$
c. The arrow labelled Fr shows where francium would be on the graph. The horizontal arrow shows what the melting temperature probably is. The melting point of francium should be approximately $20^{\circ} \mathrm{C}$.
d. Francium will probably be liquid at $25^{\circ} \mathrm{C}$.
73. Answers should include a brief biography of the chosen subject, including birth, education, and career summary of work and/or achievements, as well as their work in quantum mechanics and quantum chemistry.

## Answers to Chapter 3 Self-Assessment Questions

(Student textbook pages 204-5)

1. c
2. a
3. d
4. e
5. d
6. a
7. e
8. d
9. c
10. a
11. a. When using a cathode ray tube and generating a potential difference across the electrodes that were in a partial vacuum, a particle was emitted from the negative electrode that was drawn to the positive electrode.
b. When the experimenter placed a magnet near the tube, the 'ray' was bent in the direction in which negative particles were known to move.
c. The same results were obtained regardless of which metal was used for the electrodes.
12. Nuclear model: When Rutherford discovered that all of the positive charge and most of the mass was located in a very small region at the centre of the atom, he called it the nucleus.
Planetary model: According to Rutherford's model, the electrons orbited the nucleus much like planets orbit the Sun.
Beehive model: Some people thought that the electrons moving around the nucleus were like bees flying around a hive.
13. A continuous spectrum includes all wavelengths whereas a line spectrum contains several specific wavelengths of electromagnetic waves.
14. The symbol $n$ generally represents an energy level in both the Bohr and quantum mechanical models. In the Bohr model, $n$ refers to a specific allowed energy and orbit with a specific radius. In the quantum mechanical model, $n$ refers to an energy level or shell which has sub-levels or orbitals.
15. $n=7 ; n=5 ; n=4 ; n=2 ; n=1$ : Higher numbers of $n$ are farther from the nucleus and therefore more energy is required to pull the electron away from the nucleus.
16. 



17. When light interacts with matter, it does so in discrete quantities of energy, or photons, in an all-or-nothing manner. However, when it is travelling, it exhibits wavelike properties.
Electrons have a specific amount of mass but when they travel, they exhibit wavelike properties.
18. a. $m_{l}$ can be any integer from -2 to 2 ; five $d$ orbitals
b. $m_{l}$ can be $-1,0$, or +1 ; three $p$ orbitals
c. $m_{l}$ must be 0 ; one $s$ orbital
19. 1. Always fill the lowest energy level and sublevel of energy first before filling ones with higher energies.
2. Place only two electrons maximum in each orbital.
3. Fill each equal-energy orbital (orbitals with the same value of $l$ ) with one electron only, to half fill all of them, before adding a second electron in each one.
20. rhodium, Rh (Note: configuration should read $\left.[\mathrm{Kr}] 5 s^{1} 4 d^{8}\right)$
21. Group 14; Period 3; $p$ block
22. a. main group elements
b. transition metals or transition elements
c. inner transition metals or inner transition elements
23. - Bohr model only: quantized radii for electrons; electrons move in circular orbits; no energy is emitted during orbiting; electrons can absorb energy difference between two energy levels and move to a higher energy level; could explain line spectra of hydrogen atom

- Rutherford model only: electrons orbit randomly in circular orbits; rotating electron is predicted to emit energy and electron should lose energy and be drawn to the nucleus; could not explain line spectra of any atom
- Quantum mechanical model only: electrons have wave-like and matter-like properties; quantum numbers define the characteristics of electrons in region of space called an orbital; exact position and energy of an electron cannot be known simultaneously; explains line spectra of some multielectron atoms.
- Bohr and Quantum model: energy of electron is quantized; most probable position of electron is the same as Bohr's exact prediction of position; explain atomic spectra of hydrogen atom.
- Bohr and Rutherford models: electrons orbit the nucleus
- Rutherford and Quantum: nil
- Rutherford and Bohr and Quantum: Negative charge surrounds a central positive core; atom is a neutral species

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