Name ____________________________ Section ________ Date _________

Chapter 11: Basic Review Worksheet

1. What is *electromagnetic radiation*?
2. Sketch a representation of a wave and indicate on your drawing one wavelength of the wave.
3. Explain what it means for an atom to be in an *excited state* and what it means for an atom to be in its *ground state.*
4. What is a *photon*?
5. Describe Bohr's model of the hydrogen atom.
6. Explain what is meant by the term *orbital.*
7. What is the symbol for the lowest-energy hydrogen orbital?
8. Give the symbols for each of the orbitals that constitute the third and fourth principal energy levels of hydrogen.
9. Describe *electron spin.*
10. What does the *Pauli exclusion principle* tell us about electrons?
11. List the *order* in which the orbitals are filled as the atoms beyond hydrogen are built up.
12. How many electrons can be placed in a given *s* subshell? In a given *p* subshell? In a specific *p* orbital?
13. Define the *valence electrons* and the *core electrons* in an atom.
14. Sketch the overall shape of the periodic table and indicate the *general regions* of the table that represent the various *s, p, d,* and *f* orbitals being filled.
15. Write the electron configurations for the following atoms:
    a. Na b. N c. Be d. Sr
16. What are the *representative elements*? In what region(s) of the periodic table are these elements found? In what general area of the periodic table are the *metallic* elements found? In what general area of the table are the *nonmetals* found? Where in the table are the *metalloids* located?
17. Define the terms *ionization energy* and *atomic radius.*
18. How do the *ionization energies* and *atomic sizes* of elements vary, both within a vertical group (family) of the periodic table and within a horizontal row (period)?
19. Arrange the following atoms from largest to smallest atomic radius, and from highest to lowest ionization energy.
    a. Na, K, Rb b. C, O, F c. Na, Si, O

Name ______________________________ Section ________ Date __________

Chapter 11: Review Worksheet

1. Give some examples of electromagnetic radiation.

2. Explain what the *wavelength* (λ) and *frequency* (ν) of electromagnetic radiation represent.

3. At what speed does electromagnetic radiation move through space? How is this speed related to λ and ν?

4. How does an excited atom *return* to its ground state?

5. How is the wavelength (color) of light related to the energy of the photons being emitted by an atom? How is the energy of the photons being *emitted* by an atom related to the energy changes taking place *within* the atom?

6. How did Bohr envision the relationship between the electron and the nucleus of the hydrogen atom? How did Bohr's model explain the emission of only discrete wavelengths of light by excited hydrogen atoms? Why did Bohr's model not stand up as more experiments were performed using elements other than hydrogen?

7. How does the wave mechanical picture of the atom fundamentally differ from the Bohr model?

8. How do wave mechanical *orbitals* differ from Bohr's *orbits*? What does it mean to say that an orbital represents a probability map for an electron?

9. Describe the general characteristics of the first (lowest-energy) hydrogen atomic orbital. Does this orbital have a sharp "edge"? Does the orbital represent a surface upon which the electron travels at all times?

10. What do the *principal energy levels* and their sublevels represent for a hydrogen atom? How do we designate specific principal energy levels and sublevels in hydrogen?

11. Describe the sublevels and orbitals that constitute the third and fourth principal energy levels of hydrogen. What are the general shapes of their probability maps?

12. How does electron spin affect the total number of electrons that can be accommodated in a given orbital?

13. Why do we place unpaired electrons in the 2*p* orbitals of carbon, nitrogen, and oxygen?

14. How many electrons overall can be accommodated in the first and second principal energy levels?

15. Why are the valence electrons more important to the atom's chemical properties than the core electrons? How is the number of valence electrons in an atom related to the atom's position on the periodic table?

16. Explain how the valence-electron configuration of most of the elements can be written just by knowing the relative *location* of the element on the table. Give specific examples.

17. Write the electron configurations for the following atoms:
    a. P
    b. Se
    c. Zr
    d. Ce

18. Arrange the following atoms from largest to smallest atomic radius, and from highest to lowest ionization energy.
    a. Na, K, P
    b. Rb, N, Al
    c. Cs, I, O

Name ______________________________ Section _________ Date __________

Chapter 11: Challenge Review Worksheet

1. Do atoms in excited states emit radiation randomly, at any wavelength? Explain.

2. What does it mean to say that the hydrogen atom has only certain *discrete energy levels* available? How do we know this?

3. Why was the quantization of energy levels surprising to scientists when it was first discovered?

4. Schrodinger and de Broglie suggested a "wave-particle duality" for small particles – that is, if electromagnetic radiation showed some particle-like properties, then perhaps small particles might exhibit some wave-like properties. Explain this duality.

5. Use the wave mechanical picture of the hydrogen atom to describe what happens when the atom absorbs energy and moves to an "excited" state.

6. Summarize the postulates of the wave mechanical model of the atom.

7. Explain *why* the ionization energies and atomic sizes of the elements are related to the position on the periodic table.

8. Write an electron configuration for the following atoms that corresponds to an excited state.
    a. C
    b. Cr
    c. Br
    d. Os

14. One driving force is "energy spread", which occurs in exothermic processes as energy goes from more to less concentrated. However, if an endothermic process occurs, this must be because of "matter spread" (for example, when a gas is produced from a solid, matter is "spread").

Chapter 10: Challenge Review Worksheet

1. We have to consider both "matter spread" and "energy spread" as driving forces. While water turning to steam is favored due to matter spread, it is disfavored by energy spread. This is because energy must flow via heat into water in order to make the conversion to steam. This is one example of many processes in which the two driving forces are counter to each other, therefore making spontaneity temperature dependent.

2. We will assume the following:
   a. All of the heat lost by the iron is gained by the water.
   b. The iron has an initial temperature of 100.0°C.
   c. The density of water is 1.00 g 1mL over the temperature.
   d. The iron and water achieve the same final temperature.

   To make both signs positive, we will define the change in temperature of the iron as 100.0°C - t_f and the change in temperature of the water as t_f - 25.0°C.

   heat lost by iron = heat gained by water
   (5.00) (0.45 J/g°C) (100.0 - t_f) = (125.0g) (4.184J/g°C) (t_f - 25.0)
   225 - 2.25t_f = 523 t_f – 13075
   13300 = 525.25 t_f
   25.3°C = t_f

3. The first law of thermodynamics tells us that energy is conserved. Thus, when a ball rolls down a hill, potential energy is converted to kinetic energy due to the motion of the ball, and to frictional heating of the ball and hill. There is nothing in the first law, however, that states that the energy that has been "spread" cannot "re-concentrate" to the ball and cause the ball to roll back up the hill. The second law of thermodynamics tells us that entropy (or disorder) is constantly increasing. Thus, energy spread as the ball rolls down the hill is a driving force, and energy will not spontaneously "re-concentrate".

Chapter 11: Basic Review Worksheet

1. Electromagnetic radiation represents the propagation of energy through space in the form of waves.

2. A representative wave is depicted in Figure 11.3 in the text.

3. An atom is said to be in its ground state when it is in its lowest possible energy state. When an atom possesses more energy than in its ground state, the atom is said to be in an excited state.

4. Photons are discrete quantities of radiation. Atoms do not gain or emit radiation randomly, but rather do so only in discrete bundles of radiation called photons.

5. Bohr pictured the electron moving in only certain circular orbits around the nucleus. Each particular orbit (corresponding to a particular distance from the nucleus) had associated with it a particular energy (resulting from the attraction between the nucleus and the electron).

6. When we draw a picture of a given orbital, we are saying that there is a 90% probability of finding the electron within the region indicated in the drawing.

7. The lowest energy hydrogen atomic orbital is called the 1s orbital.

8. The third principal energy level of hydrogen is divided into three sublevels, the 3s, 3p, and 3d sublevels. The fourth principal energy level of hydrogen is divided into four sublevels, the 4s, 4p, 4d, and 4f orbitals.

9. In simple terms, in addition to moving around the nucleus of the atom, we picture the electron rotating (spinning) on its own axis. There are only two ways a body spinning on its own axis can rotate (often described as "clockwise" and "counter-clockwise" motions) to the right or to the left (for example, the earth is rotating on its axis to the right).

10. The Pauli exclusion principle summarizes our theory about intrinsic electron spin: a given atomic orbital can hold a maximum of two electrons, and those two electrons must have opposite spins.

11. The order in which the orbitals fill with electrons is indicated by the periodic table according to atomic number and the "block". Thus, we get 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, etc. It is not intuitively obvious why electrons fill the 4s orbitals before the 3d orbitals, but this is indicated by the known form of the periodic table.

12. A particular s subshell can hold two electrons. Since any p subshell consists of three orbitals, a given p subshell can hold a maximum of six electrons. A particular p orbital (like any orbital) can hold only two electrons (of opposite spin).

13. The valence electrons of an atom are the electrons in the outermost shell of the atom. The core electrons are those in principal energy levels closer to the nucleus than the outermost shell (the core electrons are the electrons that are not valence electrons).

14. This general periodic table is shown in the text as Figure 11.31. Students should not have to memorize of exactly where each element is found in the table, but they should know that the table is arranged in terms of the electronic structure of the atoms. For example, the first horizontal row of the table corresponds to the n = 1 shell, which consists of only the 1s orbital, and so there are only two elements in the row. However, the second row of the table contains eight elements, since the n = 2 shell contains a total of four orbitals (the 2s and the three 2p orbitals).

15. a. Na: $1s^2 2s^2 2p^6 3s^1$
    b. N: $1s^2 2s^2 2p^3$
    c. Be: $1s^2 2s^2$
    d. Sr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$ or [Kr] $5s^2$

35.2°C?" The specific heat capacity of water is 1.000 cal/g°C. This is the quantity of energy required to raise the temperature of only one gram of water by only one Celsius degree. In this example, we are raising the temperature by (35.2 – 15.1) = 20.1°C. So, using the specific heat capacity as a conversion factor, we can say

$$\text{energy required} = \left(\frac{1.00\ \text{cal}}{\text{g}\ ^\circ\text{C}}\right) \times (25.0\ \text{g}) \times (20.1^\circ\text{C}) = 503\ \text{cal}$$

Notice how the units of g and °C cancel, leaving the answer in energy units only. This sort of calculation of energy change can be done for any substance, using the substance's own specific heat capacity.

8. a. 56°F = 13.3°C; 75 °F = 23.9°C
   temperature change = 24 -13 = 11°C
   mass = 1.0×10^3 J/(0.13 J/g°C) (11°C) = 7.3×10^2 g
   b. temperature change = 385K – 289K = 96K = 96°C
   mass = 1.0×10^3 J/[(0.24 J/g°C) (96°C)] = 43 g
   c. 85.0°F = 29.4°C
   temperature change = 29.4°C – (-10.0°C) = 39.4°C
   mass = 1.0×10^3 J/[(0.89 J/g°C) (39.4°C)] = 29 g

9. E = q + w
   a. (-23 J) + 12 J = -11 J
   b. 14 J + (-5 cal) = 14 J + (-21 J) = -7 J
   c. 87 J + (-32 J) = 55 J

10. a. $$100.0\text{g } C_3H_8 \times \frac{1\,\text{mol } C_3H_8}{44.094\text{g } C_3H_8} \times \frac{-2221\,\text{kJ}}{1\,\text{mol } C_3H_8} = -5037\ \text{kJ}$$

    b. $$100.0\text{g } CO_2 \times \frac{1\,\text{mol } CO_2}{44.01\text{g } CO_2} \times \frac{-2221\,\text{kJ}}{3\,\text{mol } CO_2} = -1682\ \text{kJ}$$

11.

	ΔH (kJ/mol)
$I_2(s) \rightarrow I_2(g)$	62.8
$I_2(g) \rightarrow 2I(g)$	151.0
$Cl_2(g) \rightarrow 2Cl(g)$	242.3
$2I(g) + 2Cl(g) \rightarrow 2ICl(g)$	-2(211.3)
$I_2(s) + Cl_2(g) \rightarrow 2ICl(g)$	33.5

12. Yes, enthalpy is a state function. The definition of enthalpy (H) is H = E + PV. Since E, P, and V are all state functions, H is a state function as well.

13. The energy crisis has to do with the fact that the quality of the energy is decreasing (due to entropy, or more specifically "energy spread"). The amount of energy is constant, but the amount of useful work that can be done with this energy is decreasing.

16. The representative elements are the elements in Groups 1-8 of the periodic table whose valence electrons are in s and p subshells. There are two groups of representative elements at the left side of the periodic table (corresponding to the s subshells) and six groups of representative elements at the right side of the periodic table (corresponding to the p subshells). Metallic character is largest at the left-hand end of any horizontal row of the periodic table, and largest at the bottom of any vertical group. Overall, these two trends mean that the most metallic elements are at the lower left of the periodic table, and the least metallic (i.e., the nonmetals) are at the top right of the periodic table. The metalloids, which have both metallic and nonmetallic properties are found in the "stairstep" region indicated on most periodic tables.

17. The ionization energy of an atom represents the energy required to remove an electron from the atom. The atomic radius is the distance from the center of the nucleus to the outer "edge" of the valence electrons.

18. As one goes from top to bottom in a vertical group in the periodic table, the ionization energies decrease (it becomes easier to remove an electron). Within a given vertical group, the atoms get progressively larger (increase in atomic radius) when going from the top of the group to the bottom. In going from left to right within a horizontal row of the elements in the periodic table, the atoms get progressively smaller.

19. a. atomic radius: Rb > K > Na; ionization energy: Na > K > Rb
    b. atomic radius: C > O > F; ionization energy: F > O > C
    c. atomic radius: Na > Si > O; ionization energy: O > Si > Na

Chapter 11: Review Worksheet

1. Visible light, radio and television transmissions, microwaves, x-rays, and radiant heat are all examples of electromagnetic radiation.

2. The waves by which electromagnetic radiation is propagated have several characteristic properties. The wavelength (λ) represents the distance between two corresponding points (peaks or troughs) on consecutive waves, and is measured in units of length (meters, centimeters, etc.). The frequency (ν) of electromagnetic radiation represents how many complete waves pass a given point in space per second and is measured in waves/sec (hertz).

3. The speed (c) or propagation velocity of electromagnetic radiation represents how fast a given wave itself moves through space, and is equal to 3×10^8 m/sec in a vacuum. For electromagnetic radiation, these three properties are related by the formula $\lambda \times \nu = c$.

4. An atom is promoted from its ground state to an excited state by absorbing energy. When the atom returns from an excited state to its ground state, it emits the excess energy as electromagnetic radiation.

5. The photons of radiation emitted by atoms are characterized by the wavelength (color) of the radiation: longer wavelength photons carry less energy than shorter wavelength photons. The energy of a photon emitted by an atom corresponds exactly to the difference in energy between two allowed energy states in an atom. Thus, we can use an observable phenomenon (emission of light by excited atoms), to gain insight into the energy changes taking place within the atom.

6. For Bohr, when an atom absorbs energy, the electron moves from its ground state in the orbit closest to the nucleus (n = 1) to an orbit father away from the nucleus (n = 2, 3, 4,…). When an excited atom returns to its ground state, corresponding to the electron moving form an outer orbit to the orbit nearest the nucleus, the atom emits the excess energy as radiation. Since the Bohr orbits correspond to fixed distances form the nucleus and from each other, when an electron moves form one fixed orbit to another, the energy change is of a definite amount. This corresponds to a photon being emitted of a particular characteristic wavelength and energy. The original Bohr theory worked very well for hydrogen: Bohr even predicted emission wavelengths for hydrogen which had not yet been see, which were subsequently found at the exact wavelengths Bohr had calculated. However, when the simple Bohr model for the atom was applied to the emission spectra of other elements, the theory could not predict or explain the observed emission spectra.

7. The wave-mechanical theory for atomic structure developed by Schrodinger provided a model for the structure of the hydrogen atom that was consistent with all the observed properties of the hydrogen atom. Unlike Bohr's theory, however, which failed completely for atoms other than hydrogen, the wave-mechanical model was able to be extended to describe other atoms with considerable success. Rather than the fixed "orbits" which Bohr had postulated, the wave-mechanical model for the atom pictures the electrons of an atom as being distributed in regions of space called orbitals.

8. The wave-mechanical model for the atom does not describe in classical terms the exact motion or trajectory of an electron as it moves around the nucleus, but rather predicts the probability of finding the electron in a particular location within the atom. The orbitals that constitute the solutions to the mathematical formulation of the wave-mechanical model for the atom represent probability contour maps for finding the electrons. When we draw a particular picture of a given orbital, we are saying that there is a 90% probability of finding the electron within the region indicated in the drawing.

9. The 1s orbital is spherical in shape (that is, the electron density around the nucleus is uniform in all directions from the nucleus). The 1s orbital represents a probability map of electron density around the nucleus for the first principal energy level. The orbital does not have a sharp edge (it appears fuzzy) since the probability of finding the electron does not ever go to zero with distance from the nucleus. The orbital does not represent just a spherical surface on which the electron moves (this would be similar to Bohr's original theory): when we draw a picture to represent the 1s orbital we are indicating that the probability of finding the electron within this region of space is about 90%. We know that the likelihood of finding the electron within this orbital is very high, but we still don't know exactly where in this region the electron is at a given instant in time.

10. The principal energy levels (for hydrogen) correspond fairly well with the "orbits" of the Bohr theory, and are designated by an integer, n, called the "principal quantum number (n = 1,2,3,…). In the wave-mechanical model, however, these principal energy levels are further subdivided into sets of equivalent orbitals called subshells. For example, the n = 2 principal level of hydrogen is further subdivided into an s and a p subshell (indicated as the 2s and 2p subshells, respectively). These subshells, in turn, consist of the individual orbitals in which the electrons reside. The 2s subshell consists of the single, spherically-shaped 2s orbital, while the 2p subshell consists of a set of three equivalent, dumbbell-shaped 2p orbitals (which are often designated as

$2p_x$, $2p_y$, and $2p_z$ to indicate their orientation in space). Similarly, the n = 3 principal energy level of hydrogen is subdivided into three subshells: the 3s subshell (1 orbital), the 3p subshell (a set of three orbitals), and the 3d subshell (a set of five orbitals).

11. The 3s subshell consists of the single 3s orbital: like the other s orbitals, the 3s orbital is spherical in shape. The 3p subshell consists of a set of three equal-energy 3p orbitals: each of these 3p orbitals has the same shape ("dumbbell"), but each of the 3p orbitals is oriented in a different direction in space. The 3d subshell consists of a set of five 3d orbitals: the 3d orbitals have the shapes indicated in Figure 11.23, and are oriented in different directions around the nucleus (students sometimes say that four of the 3d orbitals have the shape of a 4-leaf clover). The 4s subshell consists of the single 4s orbital. The 4p subshell consists of a set ot three 4p orbitals. The 4d subshell consists of a set of five 4d orbitals. The shapes of the 4s, 4p, and 4d orbitals are the same as the shapes of the orbitals of the third principal energy level (the orbitals of the fourth principal energy level are larger and further from the nucleus than the orbitals of the third level, however). The fourth principal energy level also contains a 4f subshell, consisting of seven 4f orbitals (the shapes of the 4f orbitals are beyond the scope of this text).

12. Since there are only two possible orientations for an electron's intrinsic spin, this results in a given orbital being only able to accommodate two electrons (one spinning in each direction).

13. Since the three orbitals within a given p subshell are of exactly the same energy, and differ only in their orientation in space, when we write the electron configuration of an element like N or O which has a partially-filled p-subshell, we place the electrons in separate p-orbitals to minimize the interelectronic repulsion. Thus, the configuration of nitrogen could be written as $1s^2 2s^2 2p_x^{\,1} 2p_y^{\,1} 2p_z^{\,1}$ to emphasize this.

14. Since the first principal level consists only of the 1s orbital, the n = 1 level can contain only two electrons. Since the second principal level consists of the 2s oribital and the set of three 2p orbitals, the n = 2 level can hold a maximum of [2 + 3(2)] = 8 electrons.

15. Since the valence electrons are those in the outermost filled shell of the atom, it is these electrons that are affected by the presence of other atoms, and which are gained, lost, or shared with other atoms. The periodic table is basically arranged in terms of the valence electronic configurations of the elements (elements in the same vertical group have similar configurations): for example, all the elements in Group 1 have one valence electron.

16. Just from the column and row location of an element, one can determine the expected valence shell electronic configuration. For example, the element in the third horizontal row, in the second vertical column, has a $3s^2$ for an atom valence configuration. We know that the valence electrons are in the n = 3 shell because the element is in the third horizontal row. We know that the valence electrons are s electrons because the first two electrons in a horizontal row are always in an s subshell. We know that there are two electrons because the element is the second element in the horizontal row. As an additional example, the element in the seventh vertical column of the second horizontal row in the periodic table has the valence configuration $2s^2 2p^5$.

17. a. P: $1s^2 2s^2 2p^6 3s^2 3p^3$
    b. Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ or [Ar] $4s^2 3d^{10} 4p^4$
    c. Zr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$ or [Kr] $5s^2 4d^2$
    d. Ce: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^1 4f^1$ or [Xe] $6s^2 5d^1 4f^1$

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18. a. atomic radius: K > Na > P; ionization energy: P > Na > K
    b. atomic radius: Rb > Al > N; ionization energy: N > Al > Rb
    c. atomic radius: Cs > I > O; ionization energy: O > I > Cs

Chapter 11: Challenge Review Worksheet

1. Excited atoms definitely do not emit their excess energy in a random or continuous manner. Rather, an excited atom of a given element emits only discrete photons of characteristic wavelength and energy when going back to its ground state.

2. Since an excited atom of a given element always emits photons of exactly the same energy, we take this to mean that the internal structure of the atom is such that there are only certain discrete allowed energy states for the electrons in atoms, and that the wavelengths of radiation emitted by an atom correspond to the exact energy differences between these allowed energy states. For example, excited hydrogen atoms always display the visible spectrum shown in Figure 11.11 of the text. We take this spectrum as evidence for the existence of only certain discrete energy levels within the hydrogen atom, and we describe this by saying that the energy levels of hydrogen are quantized.

3. Previously, scientists had thought that atoms emitted energy continuously.

4. De Broglie and Schrodinger took the new idea that electromagnetic radiation behaved as if it were a steam of small particles (photons), as well as a traditional wave, and basically reversed the premise. That is, if something which had previously been considered to be entirely wave-like also had a particle-like nature, then perhaps small particles also have a wave-like nature under some circumstances.

5. When an atom (hydrogen, for example) absorbs energy, the electron moves to a higher energy state, which corresponds in the wave-mechanical model to a different type of orbital. The orbitals are arranged in a hierarchy of principal energy levels, as well as sub-levels of these principal levels.

6. Atoms have a series of principal energy levels symbolized by the letter n. The n = 1 level is the closest to the nucleus, and the energies of the levels increase as the value of n increases going out from the nucleus. Each principal energy level is divided into a set of sublevels of different characteristic shapes (designated by the letters, s, p, d, and f). Each sublevel is further subdivided into a set of orbitals: each s subshell consists of a single s orbital; each p subshell consists of a set of three p orbitals; each d subshell consists of a set of five d orbitals; etc. A given orbital can be empty or it can contain one or two electrons, but never more than two electrons (if an orbital contains two electrons, then the electrons must have opposite intrinsic spins). The shape we picture for an orbital represents only a probability map for finding electrons: the shape does not represent a trajectory or pathway for electron movements.

7. Students often find it easier to make sense of the atomic radius trends and to use these to explain the trends in ionization energy. The relative sizes of atoms vary systematically with the location of an element in the periodic table. Within a given vertical group, the atoms get progressively larger when going from the top of the group to the bottom: the valence electrons of the atoms

6. For Bohr, when an atom absorbs energy, the electron moves from its ground state in the orbit closest to the nucleus (n = 1) to an orbit father away from the nucleus (n = 2, 3, 4,…). When an excited atom returns to its ground state, corresponding to the electron moving form an outer orbit to the orbit nearest the nucleus, the atom emits the excess energy as radiation. Since the Bohr orbits correspond to fixed distances form the nucleus and from each other, when an electron moves form one fixed orbit to another, the energy change is of a definite amount. This corresponds to a photon being emitted of a particular characteristic wavelength and energy. The original Bohr theory worked very well for hydrogen: Bohr even predicted emission wavelengths for hydrogen which had not yet been see, which were subsequently found at the exact wavelengths Bohr had calculated. However, when the simple Bohr model for the atom was applied to the emission spectra of other elements, the theory could not predict or explain the observed emission spectra.

7. The wave-mechanical theory for atomic structure developed by Schrodinger provided a model for the structure of the hydrogen atom that was consistent with all the observed properties of the hydrogen atom. Unlike Bohr's theory, however, which failed completely for atoms other than hydrogen, the wave-mechanical model was able to be extended to describe other atoms with considerable success. Rather than the fixed "orbits" which Bohr had postulated, the wave-mechanical model for the atom pictures the electrons of an atom as being distributed in regions of space called orbitals.

8. The wave-mechanical model for the atom does not describe in classical terms the exact motion or trajectory of an electron as it moves around the nucleus, but rather predicts the probability of finding the electron in a particular location within the atom. The orbitals that constitute the solutions to the mathematical formulation of the wave-mechanical model for the atom represent probability contour maps for finding the electrons. When we draw a particular picture of a given orbital, we are saying that there is a 90% probability of finding the electron within the region indicated in the drawing.

9. The 1s orbital is spherical in shape (that is, the electron density around the nucleus is uniform in all directions from the nucleus). The 1s orbital represents a probability map of electron density around the nucleus for the first principal energy level. The orbital does not have a sharp edge (it appears fuzzy) since the probability of finding the electron does not ever go to zero with distance from the nucleus. The orbital does not represent just a spherical surface on which the electron moves (this would be similar to Bohr's original theory): when we draw a picture to represent the 1s orbital we are indicating that the probability of finding the electron within this region of space is about 90%. We know that the likelihood of finding the electron within this orbital is very high, but we still don't know exactly where in this region the electron is at a given instant in time.

10. The principal energy levels (for hydrogen) correspond fairly well with the "orbits" of the Bohr theory, and are designated by an integer, n, called the "principal quantum number (n = 1,2,3,…). In the wave-mechanical model, however, these principal energy levels are further subdivided into sets of equivalent orbitals called subshells. For example, the n = 2 principal level of hydrogen is further subdivided into an s and a p subshell (indicated as the 2s and 2p subshells, respectively). These subshells, in turn, consist of the individual orbitals in which the electrons reside. The 2s subshell consists of the single, spherically-shaped 2s orbital, while the 2p subshell consists of a set of three equivalent, dumbbell-shaped 2p orbitals (which are often designated as

are in progressively higher principal energy shells (and are progressively further from the nucleus) as we go down a group. In going from left to right within a horizontal row of the elements in the periodic table, the atoms get progressively smaller. Although in a given row all valence electrons in the same principal energy shell, the nuclear charge is progressively increasing from left to right, making the given valence shell progressively smaller as the electrons are drawn more closely to the nucleus. The ionization energy of an atom represents the energy required to remove an electron from the atom. As one goes from top to bottom in a vertical group in the periodic table, the ionization energies decrease (it becomes easier to remove an electron). As one goes down a group, the valence electrons are farther and farther from the nucleus (larger atomic radius) and are less tightly held. The ionization energies increase when going from left to right within a horizontal row of the periodic table. The left-hand side of the periodic table is where the metallic elements are found which lose electrons relatively easily. The right-hand side of the periodic table is where the nonmetallic elements are found: rather than losing electrons, these elements tend to gain electrons. Within a given horizontal row in the periodic table, the valence electrons are all in the same principal energy shell. However, in going from left to right in the horizontal row, the nuclear charge which holds onto the electrons is increasing one unit with each successive element, making it that much more difficult to remove an electron (and the atomic radius is smaller).

8. Answers will vary. The excited state electron configuration must have the same number of electrons as the ground state electron configuration. However, the highest energy electron is not in the ground state. The ground state electron configurations are provided below.
   a. C: $1s^2 2s^2 2p^2$
   b. Cr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d^5$ or [Ar] $4s^1 3d^5$
   c. Br: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$ or [Ar] $4s^2 3d^{10} 4p^5$
   d. Os: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^6$ or [Xe] $6s^2 4f^{14} 5d^6$

Chapter 12: Basic Review Worksheet

1. A chemical bond is a force that holds two or more atoms together and makes them function as a unit. The principal types of chemical bonding are ionic bonding, pure covalent bonding and polar covalent bonding.

2. Ionic bonding results when elements of very different electronegativities react with each other. Sodium chloride, NaCl, is an example of a typical ionic compound.

3. A bond between two atoms is, in general, covalent if the atoms share a pair of electrons in mutually completing their valence electron shells. Covalent bonds can be subclassed as to whether they are pure (nonpolar) covalent or are polar covalent. In a nonpolar covalent bond, two atoms of the same elecronegativity (often this means the same type of atom) equally share a pair of electrons: the electron cloud of the bond is symmetrically distributed along the bond axis. In a polar covalent bond, one of the atoms of the bond has a higher electronegativity than the other atom, and draws the shared pair of electrons more closely towards itself (pulling the electron cloud along the bond axis closer to the more electronegative atom). Because the more electronegative atom of a polar covalent bond ends up with a higher electron density than normal, there is a center of partial negative charge at this end of he bond. Conversely, because the less electronegative element of a polar covalent bond has some of its valence electrons pulled partly away from the atom, a center of positive charge develops at this end of the bond.

$2p_x$, $2p_y$, and $2p_z$ to indicate their orientation in space). Similarly, the n = 3 principal energy level of hydrogen is subdivided into three subshells: the 3s subshell (1 orbital), the 3p subshell (a set of three orbitals), and the 3d subshell (a set of five orbitals).

11. The 3s subshell consists of the single 3s orbital: like the other s orbitals, the 3s orbital is spherical in shape. The 3p subshell consists of a set of three equal-energy 3p orbitals: each of these 3p orbitals has the same shape ("dumbbell"), but each of the 3p orbitals is oriented in a different direction in space. The 3d subshell consists of a set of five 3d orbitals: the 3d orbitals have the shapes indicated in Figure 11.23, and are oriented in different directions around the nucleus (students sometimes say that four of the 3d orbitals have the shape of a 4-leaf clover). The 4s subshell consists of the single 4s orbital. The 4p subshell consists of a set ot three 4p orbitals. The 4d subshell consists of a set of five 4d orbitals. The shapes of the 4s, 4p, and 4d orbitals are the same as the shapes of the orbitals of the third principal energy level (the orbitals of the fourth principal energy level are larger and further from the nucleus than the orbitals of the third level, however). The fourth principal energy level also contains a 4f subshell, consisting of seven 4f orbitals (the shapes of the 4f orbitals are beyond the scope of this text).

12. Since there are only two possible orientations for an electron's intrinsic spin, this results in a given orbital being only able to accommodate two electrons (one spinning in each direction).

13. Since the three orbitals within a given p subshell are of exactly the same energy, and differ only in their orientation in space, when we write the electron configuration of an element like N or O which has a partially-filled p-subshell, we place the electrons in separate p-orbitals to minimize the interelectronic repulsion. Thus, the configuration of nitrogen could be written as $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$ to emphasize this.

14. Since the first principal level consists only of the 1s orbital, the n = 1 level can contain only two electrons. Since the second principal level consists of the 2s oribital and the set of three 2p orbitals, the n = 2 level can hold a maximum of [2 + 3(2)] = 8 electrons.

15. Since the valence electrons are those in the outermost filled shell of the atom, it is these electrons that are affected by the presence of other atoms, and which are gained, lost, or shared with other atoms. The periodic table is basically arranged in terms of the valence electronic configurations of the elements (elements in the same vertical group have similar configurations): for example, all the elements in Group 1 have one valence electron.

16. Just from the column and row location of an element, one can determine the expected valence shell electronic configuration. For example, the element in the third horizontal row, in the second vertical column, has a $3s^2$ for an atom valence configuration. We know that the valence electrons are in the n = 3 shell because the element is in the third horizontal row. We know that the valence electrons are s electrons because the first two electrons in a horizontal row are always in an s subshell. We know that there are two electrons because the element is the second element in the horizontal row. As an additional example, the element in the seventh vertical column of the second horizontal row in the periodic table has the valence configuration $2s^2 2p^5$.

17. a. P: $1s^2 2s^2 2p^6 3s^2 3p^3$
    b. Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ or [Ar] $4s^2 3d^{10} 4p^4$
    c. Zr: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$ or [Kr] $5s^2 4d^2$
    d. Ce: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 5d^1 4f^1$ or [Xe] $6s^2 5d^1 4f^1$