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.


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:

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
Chapter 12: Basic Review Worksheet

1. In general, what do we mean by a chemical bond? Name the principal types of chemical bonds.

2. What do we mean by ionic bonding? Give an example of a substance whose particles are held together by ionic bonding.

3. What do we mean by covalent bonding and polar covalent bonding? How are these two bonding types similar, and how do they differ?

4. Define electronegativity.

5. What does it mean to say that a molecule has a dipole moment?

6. Give evidence that ionic bonds are very strong. Does an ionic substance contain discrete molecules?

7. Write the electron configuration for each of the following atoms and for the simple ion that the element most commonly forms. In each case, indicate which noble gas has the same electron configuration as the ion.
   a. sodium  b. iodine  c. calcium

8. On the basis of their electron configurations, predict the formula of the simple binary ionic compound likely to form when the following pairs of elements react with each other:
   a. barium and chlorine
   b. sodium and fluorine
   c. potassium and oxygen

9. What is the most important factor for the formation of a stable compound? How do we use this requirement when writing Lewis structures?

10. In writing Lewis structures for molecules, what is meant by the duet rule? To which element does the duet rule apply?

11. In writing Lewis structures for molecules, what do we mean by the octet rule? Why is attaining an octet of electrons important for an atom when it forms bonds to other atoms?

12. What is a bonding pair of electrons? What is nonbonding (or lone) pair of electrons?
13. Write Lewis structures for the following molecules:
   a. PF₃  b. SiCl₄  c. H₂S

14. What does a *double* bond between two atoms represent in terms of the number of electrons shared? What does a *triple* bond represent?

15. What are resonance structures?

16. Determine the geometric shape for each of the following molecules:
   a. NCl₃  b. Cl₂O  c. CF₄
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 “counterclockwise” 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, 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: $1s^22s^22p^63s^1$
   b. N: $1s^22s^22p^3$
   c. Be: $1s^22s^2$
   d. Sr: $1s^22s^22p^63s^23p^64s^23d^04p^65s^2$ or [Kr]5s$^2$

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 (that is, 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: Standard 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 $\lambda$ 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 $\nu$ of electromagnetic radiation represents how many complete waves pass a given point in space per second and is measured in waves per second (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 $3 \times 10^8$ m/s in a vacuum. For electromagnetic radiation, these three properties are related by the formula $\lambda \times \nu = c.$
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 polar covalent. In a nonpolar covalent bond, two atoms of the same electronegativity (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 toward 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 the 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.

4. Electronegativity represents the relative ability of an atom in a molecule to attract shared electrons toward itself.

5. A molecule is said to possess an overall dipole moment if the centers of positive and negative charge in the molecule do not coincide.

6. The properties of typical ionic substances (hardness, rigidity, high melting and boiling points, etc.) suggest that the ionic bond is a very strong one. The formulas we write for ionic substances are really only empirical formulas, showing the relative numbers of each type of atom present in the substance. For example, when we write CaCl₂ as the formula for calcium chloride, we are only saying that there are two chloride ions for each calcium ion in the substance and not that there are distinct molecules of CaCl₂.

7. a. Na: 1s²2s²2p⁶3s¹ or [Ne]3s¹ or Na⁺: 1s²2s²2p⁶ or [Ne]
   b. I: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁵5s²4d¹⁰5p⁵ or [Kr]5s²4d¹⁰5p⁵
   c. Ca: 1s²2s²2p⁶3s²3p⁶4s²3d¹⁰4p⁵5s²4d¹⁰5p⁶ or [Xe]
   d. BaCl₂
   e. NaF
   f. K₂O

8. Bonding between atoms to form a molecule involves only the valence electrons of the atoms [not the inner (core) electrons]. Thus, when we draw the Lewis structure of a molecule, we show only the valence electrons (both bonding valence electrons and nonbonding valence electrons). The most important requisite for the formation of a stable compound (and which we try to demonstrate when we write Lewis structures) is that each atom of a molecule attains a noble gas electron configuration. When we write Lewis...
structures, we arrange the bonding and nonbonding valence electrons to try to complete the octet (or duet) for as many atoms as is possible.

10. The “duet” rule applies only for the element hydrogen: When a hydrogen atom forms a bond to another atom, the single electron of the hydrogen atom pairs up with an electron from the other atom to give the shared electron pair (that is, two electrons—a duet) that constitutes the covalent bond. By sharing this additional electron, the hydrogen atom attains effectively the 1s² configuration of the noble gas helium.

11. The “octet” rule applies for the representative elements other than hydrogen. When one of these atoms forms bonds to other atoms, the atom shares enough electrons to end up with the ns²np⁶ (that is, eight electrons—an octet) electron configuration of a noble gas.

12. Bonding electrons are those electrons used in forming a covalent bond between atoms. The bonding electrons in a molecule represent the electrons that are shared between atoms. Nonbonding (or lone pair) electrons are those valence electrons which are not used in covalent bonding and which “belong” exclusively to one atom in a molecule. For example, in the molecule ammonia (NH₃), the Lewis structure shows that there are three pairs of bonding electrons (each pair constitutes a covalent bond between the nitrogen atom and one of the hydrogen atoms) as well as one nonbonding pair of electrons that belongs exclusively to the nitrogen atom. The presence of nonbonding pairs of electrons on an atom in a molecule has a significant effect on the geometric shape of the molecule and on the molecule’s properties.

13. a.

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:F─P─F:

:F:
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b.

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:Cl─Si─Cl:

:Cl:
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c.

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H─S─H
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14. A double bond between two atoms represents the atoms sharing two pairs of electrons (four electrons) between them. A triple bond represents two atoms sharing three pairs (six electrons) between them.

15. If it is possible to draw more than one valid Lewis structure for a molecule, differing only in the location of the double bonds, we say that the molecule exhibits resonance.

16. a. trigonal pyramid  b. bent or V-shaped  c. tetrahedral