Name ____________________ Section ________ Date ________

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. PF_3 b. $SiCl_4$ c. H_2S

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. NCl_3 b. Cl_2O c. CF_4

Name ______________________ Section ________ Date _________

Chapter 12: Review Worksheet

1. What does the *bond energy* tell us about the strength of a chemical bond?

2. What experimental evidence do we have for the existence of ionic bonding? In general, what types of substances react to produce compounds having ionic bonding?

3. What circumstance must exist for a bond to be purely covalent? How does a polar covalent bond differ from an ionic bond?

4. How does the polarity of a bond depend on the *difference* in electronegativities of the two atoms participating in the bond? If two atoms have exactly the *same* electronegativity, what type of bond will exist between the atoms? If two atoms have vastly different electronegativities, what type of bond will exist between them?

5. What is the *difference* between a polar bond and a polar molecule (one that has a dipole moment)?

6. When atoms of a metal react with atoms of a nonmetal, what type of electron configurations do the resulting ions attain? Explain how the atoms in a covalently bonded compound can attain noble gas electron configurations.

7. With what general type of structure do ionic compounds occur? Sketch a representation of a general structure for an ionic compound.

8. Why is a cation always smaller and an anion always larger than the respective parent atom?

9. 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. nitrogen b. selenium c. aluminum

10. 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. aluminum and oxygen
    b. magnesium and nitrogen
    c. cesium and sulfur

11. When writing a Lewis structure, explain how we recognize when a molecule must contain double or triple bonds.

12. Write Lewis structures for the following molecules. Draw all resonance structures when appropriate.
    a. C_2H_6 b. H_2SO_4 c. SO_2

13. Determine the geometric shape for each of the following ions.
    a. chlorate ion b. chlorite ion c. perchlorate ion

Name ______________________ Section ________ Date __________

Chapter 12: Challenge Review Worksheet

1. Give an example of a molecule that has polar bonds and a dipole moment. Give an example of a molecule that has polar bonds, but that does *not* have a dipole moment. What are some implications of the fact that water has a dipole moment?

2. How is the attainment of a noble gas electron configuration important to our ideas of how atoms bond to each other?

3. Using an example, describe the bonding in an ionic compound containing polyatomic ions.

4. For three simple molecules of your own choosing, *apply* the rules for writing Lewis structures. Write the discussion as if you are explaining the method to someone who is *not* familiar with Lewis structures.

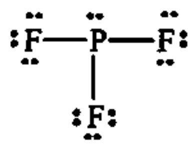
5. Although many simple molecules fulfill the octet rule, some common molecules are exceptions to this rule. Give three examples of molecules whose Lewis structures are exceptions to the octet rule and explain.

6. What do we mean by the *geometric structure* of a molecule? Draw the geometric structures of at least four simple molecules of your choosing and indicate the bond angles in the structures. Explain the main ideas of the *valence shell electron pair repulsion* (VSEPR) *theory*. Using several examples, explain how you would *apply* the VSEPR theory to predict their geometric structures.

7. What bond angle results when there are only two valence electron pairs around an atom? What bond angle results when there are three valence pairs? What bond angle results when there are four pairs of valence electrons around the central atom in a molecule? Give examples of molecules containing these bond angles.

8. How do we predict the geometric structure of a molecule whose Lewis structure indicates that the molecule contains a double or triple bond? Give an example of such a molecule, write its Lewis structure and show how the geometric shape is derived.

13. a.



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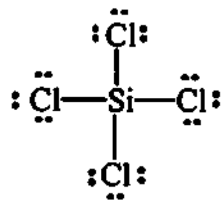

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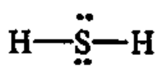


14. A double bond between two atoms represents the atoms sharing two pairs of electrons (4 electrons) between them. A triple bond represents two atoms sharing three pairs (6 electrons) between them.

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

Chapter 12: Review Worksheet

1. The strength of a chemical bond is commonly described in terms of the bond energy, which is the quantity of energy required to break the bond.

2. We know that ionically bonded solids do not conduct electricity in the solid state (since the ions are held tightly in place by all the attractive forces), but such substances are strong electrolytes when melted or when dissolved in water (either of which process sets the ions free to move around). Ionic bonds are formed when a metallic element reacts with a nonmetallic element, with the metallic element losing electrons and forming positive ions and the nonmetallic element gaining electrons and forming negative ions.

3. For a bond to be purely covalent, the electronegativity values of the atoms in the bond must be the same. A polar covalent bond represents a partial transfer of electron density from one atom to another, but with the atoms still held together as a unit. This is in contrast with an ionic bond, in which one atom completely transfers one or more electrons to a more electronegative atom, but with the resulting positive and negative ions able to separate and behave independently of one another (e.g., in solution).

4. In order for a bond to be polar, one of the atoms in the bond must attract the shared electron pair towards itself and away from the other atom of the bond. This can happen only if one atom of the bond is more electronegative than the other (that is, that there is a considerable difference in electronegativities for the two atoms of the bond). If two atoms in a bond have the same

electronegativity, then the two atoms pull the electron pair equally and the bond is nonpolar covalent. If two atoms sharing a pair of electrons have vastly different electronegativities, the electron pair will be pulled so strongly by the more electronegative atom that a negative ion may be formed (as well as a positive ion for the second atom) and ionic bonding will result. If the difference in electronegativities between two atoms sharing an electron pair is somewhere in between these two extremes (equal sharing of the electron pair and formation of ions), then a polar covalent bond results.

5. A distinction must be made between whether or not individual bonds within a molecule are polar and whether the molecule or not individual bonds within a molecule are polar and whether the molecule possesses a net dipole moment. Sometimes the geometric shape of a molecule is such that individual bond dipoles effectively cancel each other out, leaving the molecule nonpolar overall.

6. In general, when atoms of a metal react with atoms of a nonmetal, the metal atoms lose electrons until they have the configuration of the preceding noble gas, and the nonmetal atoms gain electrons until they have the configuration of the following noble gas. Covalently and polar covalently bonded molecules also strive to attain noble gas electronic configurations. For a covalently bonded molecule like F_2, in which neither fluorine atom has a greater tendency than the other to gain or lose electrons completely, each F atom provides one electron of the pair of electrons which constitutes the covalent bond. Each F atom feels also the influence of the other F atom's electron in the shared pair, and each F atom effectively fills its outermost shell. Similarly, in polar covalently bonded molecules like HF or HCl, the shared pair of electrons between the atoms effectively completes the outer electron shell of each atom simultaneously to give each atom a noble gas electronic configuration.

7. Ionic substances in the bulk consist of crystals containing an extended lattice array of positive and negative ions, in a more or less alternating pattern (that is, a given positive ion typically has several negative ions as its nearest neighbors in the crystal). A typical ionic crystal lattice is shown as Figure 12.8 in the text.

8. When an atom forms a positive ion (cation), it sheds its outermost electron shell (the valence electrons), giving a positive ion smaller than the atom from which it was formed. When an atom forms a negative ion, the atoms takes additional electrons into its outermost shell, which causes the outermost shell to increase in size because of the additional repulsive forces. This results in a negative ion which is larger than the atom from which it was formed.

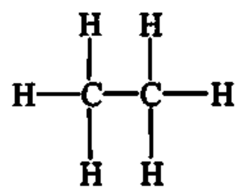
9. a. N: $1s^2 2s^2 2p^3$; N^{3-}: $1s^2 2s^2 2p^6$ or [Ne]
   b. Se: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ or [Ar] $4s^2 3d^{10} 4p^4$
   Se^{2-}: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$ or [Kr]
   c. Al: $1s^2 2s^2 2p^6 3s^2 3p^1$ or [Ne] $3s^2 3p^1$; Al^{3+}: $1s^2 2s^2 2p^6$ or [Ne]

10. a. Al_2O_3
   b. Mg_3N_2
   c. Cs_2S

11. When writing a Lewis structure for a molecule, if (after placing a pair of bonding electrons between each of the atoms to be connected) there are not enough valence electrons remaining to complete the octet (or duet) for each atom, then this strongly suggests that there must be multiple

bonding present in the molecule. For example, if a molecule seems to be two electrons short of enough to complete independently the octet, this suggests that there must be a second bond between two of the atoms (a double bond exists between those atoms). If a molecule seems to be four electrons short, this could mean either that a triple bond exists between two of the atoms, or that there are two double bonds present in the molecule.

12. a.



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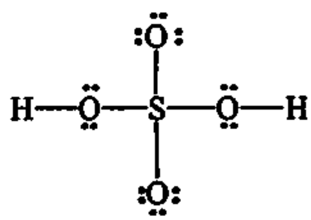

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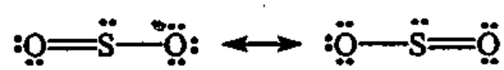


13. a. trigonal pyramid; b. bent or v-shaped; c. tetrahedral

Chapter 12: Challenge Review Worksheet

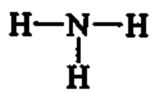
1. For example, compare the two molecules H_2O and CO_2. The water molecule is nonlinear (bent or v-shaped), whereas the CO_2 molecule is linear because of the double bonds. Both the O-H and the C=O bonds are themselves polar (since the atoms involved have different electronegativities). However, since the CO_2 molecule is linear, the two individual bond dipoles lie in opposite directions on the same axis and cancel each other out, leaving CO_2 as a nonpolar molecule overall. Water has polar bonds that lie at an angle of about 105° and thus do not cancel each other out. This means that water is a very polar molecule. Bond dipoles are in actuality vector quantities, and the overall dipole moment of the molecule represents the resultant of all the individual bond dipole vectors. The high polarity of the water molecule, combined with the existence of hydrogen bonding among water molecules, is responsible for the fact that water is a liquid at room temperature.

2. It has been observed over many, many experiments that when an active metal like sodium or magnesium reacts with a nonmetal, the sodium atoms always form Na^+ ions and the magnesium atoms always form Mg^{2+} ions. It has been further observed that aluminum always forms only the Al^{3+} ion, and that when nitrogen, oxygen, or fluorine form simple ions, the ions are always N^{3-}, O^{2-}, and F^-, respectively. Clearly the facts that these elements always form the same ions and that those ions all contain eight electrons in the outermost shell, led scientists to speculate that there must be something very fundamentally stable about a species that has eight electrons in its outermost shell (like the noble gas neon). The repeated observation that so many elements, when

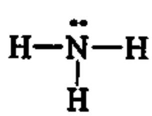
reacting, tend to attain an electronic configuration that is isoelectronic with a noble gas led chemists to speculate that all elements try to attain such a configuration for their outermost shells. However, knowing that species *do* attain a noble gas electron configuration does not explain *why* they do.

3. In the case of ionic compounds involving polyatomic ions, more than one type of bonding force is involved. First of all, ionic bonding exists between the positive and negative ion. However, within the polyatomic ions themselves, the atoms are held together by covalent bonds.

4. Answers will vary. Lets illustrate the method for ammonia, NH_3. First count up the total number of valence electrons available in the molecule (without regard to which atom they originally come from). Remember that for the representative elements, the group in which the element is found on the periodic table indicates the number of valence electrons. For NH_3, since nitrogen is in Group 5, one nitrogen atom would contribute five valence electrons. Since hydrogen atoms only have one electron each, the three hydrogen atoms provide an additional three valence electrons, for a total of eight valence electrons overall. Next write down the symbols for the atoms in the molecule, and use one pair of electrons (represented by a line) to form a bond between each pair of bound atoms.

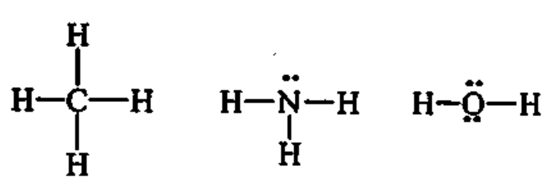


These three bonds use six of the eight valence electrons. Since each hydrogen already has its duet in what we have drawn so far, while the nitrogen atom only has six electrons around it so far, the final two valence electrons must represent a lone pair on the nitrogen.



5. There are several types of exceptions to the octet rule described in the text. The octet rule is really a "rule of thumb" which we apply to molecules. There are some common molecules, which from experimental measurements, we know do not follow the octet rule. Boron and beryllium compounds usually do not fit the octet rule. For example, in BF_3, the boron atom only has six valence electrons, whereas in BeF_2, the beryllium atom only has four valence electrons. Other molecules that are exceptions to the octet rule include any molecule with an odd number of valence electrons (such as NO or NO_2): you can't get an octet (an even number) of electrons around each atom in a molecule with an odd number of valence electrons.

6. The general geometric structure of a molecule is determined by how many electron pairs surround the central atom in the molecule, and how many of those electron pairs are used for bonding to the other atoms of the molecule. As examples, let's show how we would determine the geometric structure of the molecules CH_4, NH_3, and H_2O. First we must draw the Lewis structures for each of these molecules



In each of these structures, the central atom (C, N or O) is surrounded by four pairs (an octet) of electrons. According to the VSEPR theory, the four pairs of electrons repel each other and orient themselves in space as far away from each other as possible. This leads to a tetrahedron orientation of the electron pairs, separated by angles of 109.5°. For CH_4, each of the four pairs of electrons around the C atom is a bonding pair. We therefore say that CH_4 itself has a tetrahedral geometry, with H-C-H bond angles of 109.5°. For NH_3, however, although there are four tetrahedrally arranged electron pairs around the nitrogen atom, only three of these pairs are bonding pairs: there is a lone pair on the nitrogen atom. We describe the geometry of NH_3 as a trigonal pyramid: the three hydrogen atoms lie below the nitrogen atom in space as a result of the presence of the lone pair on nitrogen. The H-N-H bond angles are slightly less than the tetrahedral angle of 109.5°. Finally, only two of these pairs are bonding pairs: there are two lone electron pairs on the oxygen atom. We describe the geometry of H_2O as bent, v-shaped, or nonlinear: the presence of the lone pairs makes the H-O-H bond angle not 180° (linear) but somewhat less than the tetrahedral angle of 109.5°.

7.

Number of valence pairs	Bond angle	Example(s)
2	180°	BeF_2, BeH_2
3	120°	BCl_3
4	109.5°	CH_4, CCl_4, GeF_4

8. In predicting the geometric structure of a molecule, we treat a double (or triple) bond as a single entity (as if it were a single pair of electrons). This approach is reasonable, since all the bonding electrons between two atoms must be present in the same region of space between the atoms (whether one, two, or three electron pairs). For example, if we write the Lewis structure of acetylene, C_2H_2

H: C:::C:H

We realize that each carbon atom in the molecule has effectively only two "things" (termed "repulsive units" in the text) attached to it: a bonding pair of electrons (which bonds the H atom) and a triple bond (which bonds the other C atom). Since there are effectively only two things (repulsive units) attached to each carbon atom, we would expect the bond angles for each carbon atom to be 180°, which makes the molecule linear.

Chapter 13: Basic Review Worksheet

1. Gases have no fixed volume or shape, but take on the shape and volume of the container in which they are confined. This is in contrast to solids and liquids: A sample of solid has its own intrinsic volume and shape, and is very incompressible. A sample of liquid has an intrinsic volume, but does not take on the shape of its container.

2. The SI unit of pressure is the pascal, but this unit is almost never used in everyday situations because it is too small to be practical. Rather, we tend to use units of pressure that are based on the simple instruments used to measure pressures, the mercury barometer and manometer. The mercury barometer, used for measuring the pressure of the atmosphere, consists of a column of mercury that is held in a vertical glass tube by the atmosphere. The pressure of the atmosphere is indicated by the height of the mercury in the long tube (relative to the surface of the mercury in the reservoir). As the atmospheric pressure changes, the height of the mercury column changes. The height of the mercury column is given in radio and TV weather reports in inches in mercury,