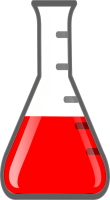
*Unit 5: Bonding, Molecular Geometry, Intermolecular Forces*

Bond

**

Ionic Bond. Taken. Never Shared.

*Important Vocabulary and formulas for chapter 5*

Metal

Nonmetal

Semi-metal (metalloid)

Chemical bond

Difference in electronegativity (ΔEn)

Covalent Bond

Polar covalent bond

Non-polar covalent bond

Covalent compound

Molecule

Ionic Bond

Ionic compound

Formula Unit

Ion

Covalent Network Solid

Metallic bond

Lewis Structures

Resonance Structures

Valence Shell Electron Pair Repulsion Theory

Polarity

Intermolecular Force (Van der Waals interactions)

Hydrogen Bond

Dipole-Dipole interaction

London dispersion force

*Notes: Introduction to Bonds*

The electromagnetic force that hold matter together are called bonds. There are three major types of bonds..

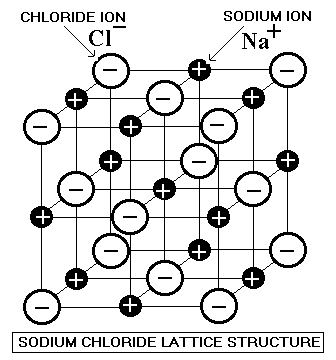
|  |  |  |  |
| --- | --- | --- | --- |
| **I.** | **Ionic Bonds** | **III.** | **Metallic Bonds** |
| **II.** | **Covalent Bonds** |

Ionic Bonds

The ionic bond is formed by the attraction between oppositely charged ions. Ionic bonds are formed between metals and nonmetals. Remember that metal atoms lose one or more valence electrons in order to achieve a stable electron arrangement. When a metal atom loses electrons it forms a positive ion or cation. When nonmetals react they gain one or more electrons to reach a stable electron arrangement. When a nonmetal atom gains one or more electrons it forms a negative ion or anion. The metal cations donate electrons to the nonmetal anions so they stick together in an ionic compound. This means that ionic bonds are formed by the complete transfer of one or more electrons.

A structure with its particles arranged in a regular repeating pattern is called a crystal. Because opposite charges attract and like charges repel, the ions in an ionic compound stack up in a regular repeating pattern called a crystal lattice. The positive ions are pushed away from other positive ions and attracted to negative ions so this produces a regular arrangement of particles where each ion is surrounded by ions of the opposite charge. Each ion in the crystal has a strong electrical attraction to its oppositely charged neighbors so the whole crystal holds together as one giant unit. There are no individual molecules in an ionic compound, just the regular stacking of positive and negative ions. Ionic compounds are represented as formula units – the lowest whole number ratio in which ions combine.

At room temperature ionic compounds are high melting point solids. They are usually white except for compounds of the transition metals that may be colored. They are brittle (break easily). They do not conduct electricity as solids, but do conduct electricity when melted or dissolved in water.

1. Define the following terms:
   1. ionic bond –
   2. cation –
   3. anion –
   4. crystal –
   5. formula unit –
2. What are the smallest units of an ionic bond?
3. List several properties of ionic compounds:
4. Under what conditions do ionic compounds conduct electricity?

Covalent Bonds

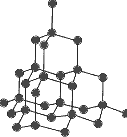
A covalent bond is formed between nonmetal atoms. The nonmetals are connected by a shared pair of valence electrons forming a molecule. A molecule is a neutral particle of two or more atoms bonded to each other. Molecules may contain atoms of the same element such as N2, O2, and Cl2 or they may contain atoms of different elements like H2O, NH3, or C6H12O6.

Covalent substances have low melting points and boiling points compared to ionic compounds or metals. At room temperature, covalent substances are gases, liquids or low melting point solids (think of fats and greases). They do not conduct electricity as solids or when molten and usually do not conduct when dissolved in water.

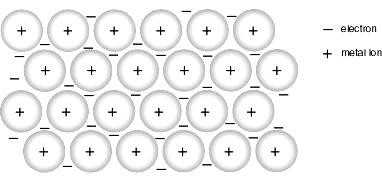
There are several types of covalent bonds. A **single** covalent bond is formed when two atoms share one pair of valence electrons. A **double** covalent bond is when two atoms share two pairs of valence electrons. A **triple** covalent bond is when two atoms share three pairs of valence electrons

1. Define the following terms:
   1. covalent bond –
   2. molecule –
2. List several properties of covalent compounds.
3. Define the following terms:
   1. single covalent –
   2. double covalent –
   3. triple covalent –

There is one last type of covalent bonding—the bonding in network solids (macromolecules). In this type of bonding, atoms share valence electrons but the atoms are arranged in a regular crystalline pattern in which each atom is covalently bonded to its neighbors in all directions. Therefore, you do not have a collection of small molecules that are easy to separate from each other; the whole system is one giant molecule or a macromolecule held together by this network of strong covalent bonds. Network solids are extremely hard, brittle, solids that do not conduct electricity. Diamonds (a form of pure carbon (see figure)), carborundum (silicon carbide) and quartz (silicon dioxide) are examples of macromolecules.

1. What is a network solid?
2. What type of bonding exists in network solids?
3. What are some properties of network solids?
4. What are some examples of network solids?

*Metallic Bonds*

A metallic bond forms between multiple metal atoms. The metallic bond is formed by the mutual attraction of a positively charged cation to “delocalized” valence electrons. Most metal atoms have only one or two valence electrons and these are not tightly bound to the atoms (low ionization energy). In a piece of metal these valence electrons do not seem to belong to any one of the atoms but are able to move freely through the structure from one atom to another. Metals can be thought of as positive ions (the nucleus and inner shells of electrons—all of the atom except the valence electrons) in a “sea” of loosely bound valence electrons.The metal ions line up in a regular repeating

pattern (a crystal lattice) and their loose valence electrons move through this crystal acting as an electron glue (see figure). Each of the ions is strongly attracted to all of the loose electrons surrounding it so the whole metal holds together as a crystal.

These electrical attractions for the electron glue are strong and hard to break so metals are high melting point crystalline solids. Since there are charged particles free to move metals are good conductors of heat and electricity as solids and as liquids. Because the “electron glue” is free to move, if we hammer or pull the cations to new positions the electron glue flows right along with the cations and holds the structure together in the new position. Thus, metals are malleable (can be hammered into sheets) and ductile (can be stretched into wires) and have a high tensile strength (can be stretched without breaking). Metals also tend to have high luster. This is a direct consequence of loosely bound valence electrons.

Metallic bonding is found in elemental metals and in **mixtures of metals called alloys.**

1. What is a metallic bond? Explain how the ions and electrons are arranged.
2. List some properties of metallic bonds.
3. What is an alloy?
4. Identify the following compounds as metallic, ionic or covalent:
   1. NaCl Cl2 Au BrO3-
   2. Mg3N2 Pt Al Ag

*BOND TYPES based on Electronegativity difference.*

Nonpolar covalent-(0 to 5% ionic character) this is a covalent bond in which the bonding electrons are shared equally by the bonded atoms, with a resulting balanced distribution of electrical charge

Polar covalent-(5 to 50% ionic character) this is a covalent bond in which the united atoms have unequal attraction for the shared electrons

Ionic- (50 to 100% ionic character) this is a chemical bond in which the electrons are so unequally shared that the electrons are transferred between the united atoms - a metal atom will give one or more valence electrons to a nonmetal atom

% Ionic Character 100 % 50 % 5 % 0 %

Difference in

Electronegativity (ΔE) 4.0 1.67 0.4 0.0

### **Electronegativities of the Elements**

|  |
| --- |
| H  2.20 |
| Li  0.98 | Be  1.57 |  |  |  |  |  |  |  |  |  |  | B  2.04 | C  2.55 | N  3.04 | O  3.44 | F  3.98 |
| Na  0.93 | Mg  1.31 |  |  |  |  |  |  |  |  |  |  | Al  1.61 | Si  1.90 | P  2.19 | S  2.58 | Cl  3.16 |
| K  0.82 | Ca  1.00 | Sc  1.36 | Ti  1.54 | V  1.63 | Cr  1.66 | Mn  1.55 | Fe  1.83 | Co  1.91 | Ni  1.91 | Cu  1.90 | Zn  1.65 | Ga  1.81 | Ge  2.01 | As  2.18 | Se  2.55 | Br  2.96 |
| Rb  0.82 | Sr  0.95 | Y  1.22 | Zr  1.33 | Nb  1.6 | Mo  2.16 | Tc  1.9 | Ru  2.2 | Rh  2.28 | Pd  2.20 | Ag  1.93 | Cd  1.69 | In  1.78 | Sn  1.96 | Sb  2.05 | Te  2.1 | I  2.66 |
| Cs  0.79 | Ba  0.89 | La  1.10 | Hf  1.3 | Ta  1.5 | W  2.36 | Re  1.9 | Os  2.2 | Ir  2.20 | Pt  2.28 | Au  2.54 | Hg  2.00 | Tl  2.04 | Pb  2.33 | Bi  2.02 | Po  2.0 | At  2.2 |
| Fr  0.7 | Ra  0.9 | Ac  1.1 |  |  |  |  |  |  |  |  |  |  |  |  |  |  |

Notes: Drawing Lewis Structures for Covalent compounds.

Covalent bonds typically occur between two nonmetal atoms. There is not enough of an electronegativity difference between the two atoms for one atom to take valence electrons away from the other atoms. Therefore, in order for each atom to bond according to the octet rule, the two atoms must share their valence electrons. When the valence electrons are shared between two atoms it is called a covalent bond.

A single covalent bond is formed when 2 electrons (1 pair) is shared between two atoms. When 4 electrons (2 pair) are shared between two atoms it is called a double bond. And 6 electrons (3 pair) shared between two atoms forms a triple bond. A pair of valence electrons that is not shared between two atoms is called a lone or unshared pair. When drawing Lewis Dot Structures for covalent molecules we use a dash to represent a bonding pair of electrons and two dots to represent a lone pair of electrons. We do not show the process of bond formation as we do in an ionic bond, but only the finished molecule. The process used to draw a Lewis Dot Structure for a covalent molecule is very different from the process used for ionic compounds.

Important things to know about elements when they form covalent bonds:

* Many elements react to achieve a full valence shell configuration, similar to noble gases. Normally this represents 8 electrons. There are exceptions to this of course.
* The following atoms react to have less than 8 electrons
  + Hydrogen generally reacts so that it has 2 electrons in its valence shell.
  + Beryllium generally reacts so that it has 4 electrons in its valence shell.
  + Boron generally reacts so that it has 6 electrons in its valence shell.
* Elements that have valence electrons in the third energy level or higher may react to have more than 8 electron.
  + Sulfur commonly reacts to form 6 covalent bonds (12 electrons)
  + Phosphorus commonly reacts to form 5 covalent bonds (10 electrons)
  + The non-metals in the 3rd, 4th, 5th and 6th period can have more than 8.

|  |  |  |  |
| --- | --- | --- | --- |
|  | **Ammonia NH3** | **Carbonate CO32-** | **Krypton tetrafluoride KrF4** |
| Step 1: Add up all valence and charge electrons | Nitrogen has 5 and each hydrogen has 1 for a total of 8.  5 + (3×1) = 8 | Carbon has 5, each oxygen has 6 and two extra for the charge.  4 + (3 × 6) + 2 = 24 | Krypton has 8 and each fluorine has 7.  8 + (4 × 7) = 36 |
| Step 2: Draw the skeleton structure with the surrounding atoms connected to the central atom by single bonds. |  |  |  |
| Step 3: Add three pairs of electrons to all surrounding atoms except for hydrogen. Never add electrons to hydrogen |  |  |  |
| Step 4: Count up the total amount of electrons you have shown and subtract from the total number of valence electrons. Add any leftover electrons as pairs to the central atom. | 8 – 6 = 2 leftover electrons. Place on central nitrogen. | 24-24 = 0 leftover electrons. | 36 – 32 = 4 leftover electrons.  Place as pairs on central Kr. |
| Step 5: If all atoms have satisfied valence shells then you are done. If not, move two electrons from a surrounding electron to form a double bond. Repeat if necessary to form a triple bond. |  |  |  |
| Step 6: Check to see if it’s possible to draw resonance structures. This happens when you can put a double bond in more than one place. You must draw all possibilities. |  | | |

All resonance structures are contained within brackets and separated by a double headed arrow. If there is a charge, it is written outside of the brackets.

*Notes: Molecular Geometry (VSEPR theory)*

**VSEPR Theory:**

Many of the properties of molecular compounds depend on the three dimensional shape of the molecule. The shape of a molecule is predicted by using the VSEPR theory:

VSEPR Theory: the electrostatic repulsion between the valence level electron pairs surrounding an atom causes these pairs to be oriented as far apart as possible.

To use VSEPR theory:

1. Draw the Lewis Dot Structure for the molecule.
2. Determine the number of links on the central atom (single, double and triple bonds all count as one link.)
3. Determine the number of lone pairs on the central atom.

|  |  |  |  |
| --- | --- | --- | --- |
| Number of areas of electron density | Number of Links on Central Atom | Number of Lone Pairs  **on Central Atom** | Shape of Molecule |
| 2 | 2 | 0 | Linear |
| 3 | 3 | 0 | Trigonal Planar |
| 3 | 2 | 1 | Bent |
| 4 | 4 | 0 | Tetrahedral |
| 4 | 3 | 1 | Trigonal pyramidal |
| 4 | 2 | 2 | Bent |
| 5 | 5 | 0 | Trigonal bipyramidal |
| 5 | 4 | 1 | see-saw |
| 5 | 3 | 2 | t-shaped |
| 5 | 2 | 3 | linear |
| 6 | 6 | 0 | octahedral |
| 6 | 5 | 1 | square pyramidal |
| 6 | 4 | 2 | square planar |

Polar Molecule vs Nonpolar Molecules

One very important property of molecule or is whether it is polar or nonpolar. If the electrons in a molecule are not evenly distributed, the molecule or can have a negative and positive side. A polar molecule will dissolve in another polar substance (such as water) where a nonpolar substance will not. Similarly, a nonpolar molecule will dissolve in another nonpolar substance (such as carbon tetrachloride) and a polar molecule will not dissolve in a nonpolar substance. We can determine if a molecule is polar or nonpolar by looking at the Lewis Dot Structure and following the rules below:

To determine if a molecule is a polar or nonpolar, look at the central atom.

1. If the central atom has no lone pairs and has all the same types of atoms attached to it, then the molecule is nonpolar.
2. If the central atom has no lone pairs but different atoms attached to it, the molecule is polar.
3. If the central atom has lone pairs, the molecule is polar,

**VSEPR Table**

|  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- |
| **Outer Atoms  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** | **Lone Pairs  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** | **Areas of Electron Density  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** | **Shape  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** | **Ideal Bond Angle  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** | **Example  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** | **Image  [↓](http://en.wikipedia.org/wiki/Molecular_geometry)** |
| 2 | 0 | 2 | linear | 180° | BeCl2 | [50px-Linear-3D-balls](http://en.wikipedia.org/wiki/Image:Linear-3D-balls.png) |
| 3 | 0 | 3 | trigonal planar | 120° | BF3 | [50px-Trigonal-3D-balls](http://en.wikipedia.org/wiki/Image:Trigonal-3D-balls.png) |
| 2 | 1 | 3 | bent | 120° | SO2 | [50px-AX2E1-3D-balls](http://en.wikipedia.org/wiki/Image:AX2E1-3D-balls.png) |
| 4 | 0 | 4 | tetrahedral | 109.5° | CH4 | [50px-AX4E0-3D-balls](http://en.wikipedia.org/wiki/Image:AX4E0-3D-balls.png) |
| 3 | 1 | 4 | trigonal pyramidal | 109.5° | NH3 | [50px-AX3E1-3D-balls](http://en.wikipedia.org/wiki/Image:AX3E1-3D-balls.png) |
| 2 | 2 | 4 | bent | 109.5° | H2O | [50px-AX2E2-3D-balls](http://en.wikipedia.org/wiki/Image:AX2E2-3D-balls.png) |
| 5 | 0 | 5 | trigonal bipyramidal | 90°, 120° | PCl5 | [50px-Trigonal-bipyramidal-3D-balls](http://en.wikipedia.org/wiki/Image:Trigonal-bipyramidal-3D-balls.png) |
| 4 | 1 | 5 | seesaw | 90°, 120° | SF4 | [50px-AX4E1-3D-balls](http://en.wikipedia.org/wiki/Image:AX4E1-3D-balls.png) |
| 3 | 2 | 5 | T-shaped | 90° | ClF3 | [50px-AX3E2-3D-balls](http://en.wikipedia.org/wiki/Image:AX3E2-3D-balls.png) |
| 2 | 3 | 5 | linear | 180° | XeF2 | [50px-AX2E3-3D-balls](http://en.wikipedia.org/wiki/Image:AX2E3-3D-balls.png) |
| 6 | 0 | 6 | octahedral | 90° | SF6 | [50px-AX6E0-3D-balls](http://en.wikipedia.org/wiki/Image:AX6E0-3D-balls.png) |
| 5 | 1 | 6 | square pyramidal | 90° | BrF5 | [50px-AX5E1-3D-balls](http://en.wikipedia.org/wiki/Image:AX5E1-3D-balls.png) |
| 4 | 2 | 6 | square planar | 90° | XeF4 | [50px-Square-planar-3D-balls](http://en.wikipedia.org/wiki/Image:Square-planar-3D-balls.png) |

*Notes: Intermolecular Forces*

**Intermolecular forces** are forces that hold molecules together. Although intermolecular forces (IMF’s) are weaker than either ionic or covalent bonds, they are very important in determining the properties of molecular compounds such as: boiling point, melting point, vapor pressure, solubility and viscosity. There are three types of IMF’s

* 1. **Dispersion forces:** Dispersion forces are caused by the asymmetric distribution of electrons around an atoms nucleus at a given instant of time. The more electrons present in a molecule the greater the dispersion force is. This is particularly well illustrated by the halogens. The two smallest halogens, F and Cl, exist as gases. The middle halogen, Br, exists as a liquid. And I and At, the two largest halogens exist as solids. The more electrons present the stronger the molecules are attracted to each other. All elements/molecules exhibit dispersion forces. Why would that be?
  2. **Dipole-dipole** interactions occur when two polar molecules are attracted to one another. The part of the polar molecule with the highest electron density is attracted to the part of another polar molecule with the lowest electron density. The region of a polar molecule with high electron density is slightly negatively charged and the area of lowest electron density is slightly positively charged. Dipole-dipole interactions are somewhat similar to ionic bonds but much weaker. Only polar molecules exhibit dipole-dipole interactions. What makes molecules polar?
* **Hydrogen bonds** are the strongest of all intermolecular forces. Hydrogen bonds have about 5% of the strength of an average covalent bond. These strong attractions are what gives water it’s unique properties. Hydrogen bonds are formed when a hydrogen bonded to an extremely electronegative atom (N, O, F) is attracted to the lone pairs on an extremely electronegative atom (N, O, F) on another molecule. A hydrogen bond is a very strong dipole-dipole interaction.

A great range of physical properties occurs among covalent molecules. This is due the various intermolecular attractions between the molecules. Typically you will find the properties of molecular compounds are much different the properties of ionic compounds. This is because in a ionic compound all of the particles are connected by the electrostatic attractions we call ionic bonds. These ionic bonds are very strong and must be broken to cause a change in the properties of the ionic compounds. In a molecular compound it is usually only necessary to break the weak intermolecular attractions between the molecules which is easy to do in most cases. A summary of some of the characteristic differences between ionic and covalent compounds is listed below.

|  |  |  |
| --- | --- | --- |
| **Characteristic** | **Ionic Compound** | **Covalent Compound** |
| **Representative Unit** | Formula Unit | Molecule |
| **Bond Formation** | Electrons transferred from on atom to the other atom | Electron pairs are shared between two atoms. |
| **Type of Elements** | Metal + Nonmetal | Nometals |
| **Physical State** | Solid | Solid, Liquid, or Gas |
| **Melting Point** | High (usually above 300°C) | Low (usually below 300°C) |
| **Solubility in water** | Usually high | High to low |
| **Electrical conductivity of aqueous solution** | Good Conductor | Poor to nonconducting |

*Homework #1: Determining Bond Type Homework*

1. *Predict what kind of bond would occur between each set of two elements below using electronegativity differences.*

a. Na and F

b. Mg and O

c. N and Cl

d. Mn and Se

e. K and Br

f. Fr and F

g. Cr and As

h. Rb and P

i. Sr and O

j. Se and I

k. S and Cl

l. Si and O

m. H and C

n. Ba and S

o. H and Cl

p. S and O

q. Li and Br

r. Ni and Te

s. Ag and F

t. P and Br

1. What can you conclude about the following statement? “Metals react with non-metals to form ionic compounds.”
2. In general, metals transfer electrons to nonmetals. In each of the following combinations, use Lewis dot structures to show the transfer of electrons then box your final answer. Use Lewis Dot Structures.

**Ex:** sodium nitride (Na3N)



1. sodium chloride (NaCl) 5. potassium fluoride (KF)
2. barium oxide (BaO) 6. sodium oxide (Na2O)
3. magnesium chloride (MgCl2) 7. aluminum chloride (AlCl3)
4. calcium phosphide (Ca3P2) 8. rubidium oxide (Rb2O)

*Homework #2 – Lewis Structures of Covalent Compounds*

*Draw the Lewis Structure for each compound or ion listed below:*

|  |  |
| --- | --- |
| 1. PH3 | 6. SO4-2 |
| 2. SO2 | 7. OF2 (O in the middle) |
| 3. HOOH | 8. AsF3 |
| 4. HCN | 9. COCl2  (every thing attaches to C) |
| 5. SiH4 | 10. H2S |

*Draw the Lewis Structure for each compound or ion listed below:*

|  |  |
| --- | --- |
| 11. PO4-3 | 16. BH3 (B does not have 8!) |
| 12. SO3-2 | 17. NI3 |
| 13. BrO4-1 | 18. H2CCH2 |
| 14. SeO3 | 19. NCCN (two triple bonds) |
| 15. HCCH | 20. H3CCN (triple bond) |

### Homework #3: Lewis Structures, VSEPR, and Molecule Polarity

*Draw the Lewis Structure for each compound or ion listed below, give its shape, and determine if it’s polar or not:*

|  |  |
| --- | --- |
| 1. HOCl   Shape:  Polar or nonpolar: | 6. HOF  Shape:  Polar or nonpolar: |
| 1. Cl2CO   Shape:  Polar or nonpolar: | 7. CS2  Shape:  Polar or nonpolar: |
| 3. CHClF2  Shape:  Polar or nonpolar: | 8. HBF2  Shape:  Polar or nonpolar: |
| 4. NH2Cl  Shape:  Polar or nonpolar: | 9. SiH4  Shape:  Polar or nonpolar: |
| 5. Cl2O  Shape:  Polar or nonpolar: | 10. SO2  Shape:  Polar or nonpolar: |

**0**

|  |  |
| --- | --- |
| 11. FCl2+  Shape:  Polar or nonpolar: | 16. SO3  Shape:  Polar or nonpolar: |
| 12. CS2  Shape:  Polar or nonpolar: | 17. FO3-  Shape:  Polar or nonpolar: |
| 13. AsO3-3  Shape:  Polar or nonpolar: | 18. N2O (put 1 N in the middle)  Shape:  Polar or nonpolar: |
| 14. IO4-  Shape:  Polar or nonpolar: | 19. H2CO  Shape:  Polar or nonpolar: |
| 15. SeO2  Shape:  Polar or nonpolar: | 20. NO2-  Shape:  Polar or nonpolar: |

*Homework #4: More Lewis Structures and VSEPR*

|  |  |
| --- | --- |
| 1. SbH3  Shape:  Polar or nonpolar: | 6. PCl4+  Shape:  Polar or nonpolar: |
| 2. NF3  Shape:  Polar or nonpolar: | 7. CO3-2  Shape:  Polar or nonpolar: |
| 3. IO2-  Shape:  Polar or nonpolar: | 8. SiO4-4  Shape:  Polar or nonpolar: |
| 4. NO3-  Shape:  Polar or nonpolar: | 9. BF3  Shape:  Polar or nonpolar: |
| 5. CNBr  Shape:  Polar or nonpolar: | 10. CBr4  Shape:  Polar or nonpolar: |

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| 11. HCN |  |  | 16. SeF2 |  |
| Shape:  Polar or Nonpolar: |  |  | Shape:  Polar or Nonpolar: |  |
| 12. PH3 |  |  | 17. POCl3 |  |
| Shape:  Polar or Nonpolar: |  |  | Shape:  Polar or Nonpolar: |  |
| 13. CHCl3 |  |  | 18. XeO4 |  |
| Shape:  Polar or Nonpolar: |  |  | Shape:  Polar or Nonpolar: |  |
| 14. NH4+1 |  |  | 19. PO4-3 |  |
| Shape:  Polar or Nonpolar: |  |  | Shape:  Polar or Nonpolar: |  |
| 15. BF4-1 |  |  | 20. ClO4-1 |  |
| Shape:  Polar or Nonpolar: |  |  | Shape:  Polar or Nonpolar: |

*Homework #5: VSEPR*

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Molecular  Formula | Lewis Dot  Structure | # of Areas of  Electron Density | # of lone pairs | Electronic  Structure | Molecular  Structure | Bond  Angle | Molecular Polarity |
| POCl3 |  |  |  |  |  |  |  |
| SCl2 |  |  |  |  |  |  |  |
| SCl4 |  |  |  |  |  |  |  |
| PCl3 |  |  |  |  |  |  |  |
| BrF5 |  |  |  |  |  |  |  |
| HCN |  |  |  |  |  |  |  |
| KrF4 |  |  |  |  |  |  |  |
| OF2 |  |  |  |  |  |  |  |
| AsF5 |  |  |  |  |  |  |  |
| AsF3 |  |  |  |  |  |  |  |

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| Molecular  Formula | Lewis Dot  Structure | # of Areas of  Electron Density | # of lone pairs | Electronic  Structure | Molecular  Structure | Bond  Angle | Molecular Polarity |
| TeF4 |  |  |  |  |  |  |  |
| TeF6 |  |  |  |  |  |  |  |
| CS2 |  |  |  |  |  |  |  |
| ICl3 |  |  |  |  |  |  |  |
| SeO3 |  |  |  |  |  |  |  |
| BrF3 |  |  |  |  |  |  |  |
| SiCl4 |  |  |  |  |  |  |  |
| HNO2 |  |  |  |  |  |  |  |
| HClO2 |  |  |  |  |  |  |  |
| KrF2 |  |  |  |  |  |  |  |
| BCl3 |  |  |  |  |  |  |  |
| SCl6 |  |  |  |  |  |  |  |

*Periodicity Review Questions*

1. Who first arranged the elements according to nuclear charge?

a. Henry Mosley b. Antoine Lavoisier c. John Dalton d. Dmitri Mendeleev

2. The modem periodic table is arranged in order of increasing atomic \_\_\_\_\_.

a. mass b. charge c. number d. radius

3. Who was responsible for establishing the order in which elements are presented in the modem periodic table?

a. Dmitri Mendeleev b. Ernest Rutherford c. Niels Bohr d. Henry Mosley

4. The periodic law states there is a periodic repetition of physical and chemical properties of the elements \_\_\_\_\_.

1. when they are arranged in order of increasing atomic mass.
2. if only metals are considered.
3. when they are arranged in order of increasing atomic radii.
4. when they are arranged in order of increasing atomic number.

5. The vertical columns of the periodic table are called \_\_\_\_\_.

a. groups b. rows c. periods d. transitions

6. Which of the following elements is in the same period as phosphorus?

a. carbon b. magnesium c. nitrogen d. oxygen

7. Which subatomic particle plays the greatest part in determining the physical/chemical properties of an element?

a. electron b. neutron c. proton d. nucleus

8. What is true of the electron configuration of the noble gases?

1. The outermost s and p sublevels are filled.
2. The outermost s and p subleve1s are partially filled.
3. The outermost s and d sublevels are very close in energy and have electrons in them.
4. The outermost s and f sublevels are very close in energy and have electrons in them.

9. Each period in the periodic table corresponds to \_\_\_\_\_.

a. a principal energy level b. an energy sublevel c. an atomic mass d. an atomic number

10. How does atomic radius change from left to right across a period in the periodic table?

a. it tends to decrease b. it tends to increase

c. it first increases, then decreases d. it first decreases, then increases

11. How does atomic radius change down a group in the periodic table?

a. it tends to decrease b. it tends to increase

c. it first increases, then decreases d. it first decreases, then increases

12. Which of the following factors contributes to the relatively greater atomic size of the higher atomic number elements within a particular family of the periodic table?

a. greater number of energy levels b. larger nuclei

c. greater number of protons d. smaller number of valence electrons

13. Which of the following elements has the smallest atomic radius?

a. chlorine b. sulfur c. selenium d. bromine

14. Which of the following elements has the smallest atomic radius?

a. lithium b. potassium c. oxygen d. sulfur

15. Atomic size generally \_\_\_\_\_

1. increases as you move from left to right across a period.
2. decreases as you move down a group.
3. remains constant within a period.
4. decreases as you move from left to right across a period.

16. The energy required to remove an electron from a sodium atom is called \_.

a. electron affinity b. ionization energy c. electronegativity d. excitation energy

17. Which of the following factors contributes to the lower ionization energy of the higher atomic number elements in a family in the periodic table?

a. greater distance from the nucleus b. larger nuclei

c. greater number of protons in the nuclei d. smaller number of valence electrons

18. Which of the following factors contributes to the greater ionization energy of the elements on the right side of a period in the periodic table?

a. more energy levels b. larger nuclei

c. greater number of protons in the nuclei d. smaller number of valence electrons

19. Which of the following has the smallest first ionization energy?

a. sodium b. calcium c. potassium d. magnesium

20. As you move from left to right across the second period of the periodic table, -

a. the ionization energy increases. b. the atomic radii increase.

c. the electron affinity decreases. d. the atomic mass decreases.

21. Of the following atoms, which one has the smallest first ionization energy?

a. boron b. carbon c. aluminum d. silicon

22. Which of the following's a measure of the energy change occurring when an electron's added to a neutral atom? a. ionization energy b. bonding energy c. electronation energy d. electron affinity

23. Which group of the periodic table has the highest electronegativities?

a. halogens b. alkali metals c. alkaline earth metals d. noble gases

24. Which of the following factors contributes to the relatively lower electronegativites of the higher atomic number elements in any particular family of the periodic table?

a. greater distance from the nuclei b. larger nuclei

c. greater number of protons in the nuclei d. smaller number of valence electrons

25. Which of the following elements has the lowest electron affinity? .

a. sulfur b. selenium c. bromine d. chlorine

26. Electronegativity -

1. is the energy change that accompanies the loss of en electron from an atom.
2. generally decreases as you move down a group.
3. is generally higher for metals than for nonmetals.
4. generally decreases from left to right across a period.

27. Why is the radius of a positive atom always less than the radius of its neutral atom?

1. The nucleus pulls the remaining electrons closer.
2. The number of principal energy levels is always reduced.
3. The atomic orbitals contract all by themselves.
4. Electron speeds are reduced.

28. Why is the radius of a negative ion always greater than the radius of its neutral atom?

1. Because the number of principal energy levels is always reduced.
2. Because atomic orbitals contract all by themselves.
3. Because electron speeds are reduced.
4. Because repulsion between electrons increases.

29. Which of the following factors contributes to the greater ionic radii of the elements at the right of a particular period, relative to the elements at the left of the same period?

a. larger nuclei b. greater number of protons in nuclei

c. smaller number of valence electrons d. more electron repulsion

30. What term describes an atom's tendency to attract electrons to itself when chemically bonded to another atom?

a. electron affinity b. electronegativity c. electrolysis d. electrochemical attraction

31. Which of the following elements is the least electronegative?

a. chlorine b. cesium c. oxygen d. fluorine

32. Compared with the electronegativities of the elements at the left end of a particular period, the electronegativities of the elements at the right side of that same period tend to be \_\_\_\_.

a. lower b. higher c. the same d. unpredictable

33. Which of the following factors is responsible for the fact that higher atomic number elements in any particular family tend to have lower electronegativities than lower atomic number elements in the same family?

a. greater distance from the nuclei b. larger nuclei

c. greater number of protons in the nuclei c. smaller number of valence electrons

34. Which of the following increases with increasing atomic number in the alkaline earth metals?

a. atomic radius b. electron affinity c. first ionization energy d. # of outermost electrons

35. For nonmetals, chemical activity will

1. increase from left to right and decrease from top to bottom.
2. increase from left to right and increase from top to bottom.
3. decrease from left to right and decrease from top to bottom.
4. decrease from left to right and increase from top to bottom.

36. For metals, chemical activity will

1. increase from left to right and decrease from top to bottom.
2. increase from left to right and increase from top to bottom.
3. decrease from left to right and decrease from top to bottom.
4. decrease from left to right and increase from top to bottom.

37. Which of the following is the most reactive?

a. phosphorus b. selenium c. sulfur d. arsenic

38. Which of the following is the least reactive?

a. rubidium b. barium c. cesium d. strontium

39. Which of the following has the most metallic character?

a. rubidium b. barium c. cesium d. strontium

40. Which of the following groups of ions is arranged correctly in order of increasing size?

a. K+1, S-2, Cl-1 b. Cl-1, S-2, K+1 c. Cl-1, K+1, S-2 d. K+1, Cl-1, S-2

41. Which of the following is expected to have the smallest radius

a. N-3 b. O-2 c. F-1 d. Ne

42. Which of the following groups of elements are arranged in order of INCREASING fJrst ionization energy?

a. S, Cl, Se, Br b. Br, Se, Cl, S c. Br, Cl, Se, S d. Se, Br, S, Cl

43. Which of the following would be expected to gain electrons and form negative ions in a reaction?

a. sodium b. strontium c. copper d. sulfur

44. Which of the following would be expected to lose electrons and form positive ions in a reaction?

a. argon b. oxygen c. bromine d. potassium

45. Mendeleev left several holes in his first published periodic table. He felt the reason for these holes was

a. incomplete measurements. b. undiscovered elements

c. nonexistent elements. d. manmade elements

46. Halogens form \_\_\_\_\_ ions by \_\_\_\_\_.

a. negative, gaining 1 electron b. positive, by losing 1 electron

c. negative, gaining 2 electrons d. positive, by losing 2 electrons

47. Alkali metals form \_\_\_\_\_ ions by \_\_\_\_\_.

a. negative, gaining 1 electron b. positive, by losing 1 electron

c. negative, gaining 2 electrons d. positive, by losing 2 electrons

48. Alkaline earth's form \_\_\_\_\_ ions by \_\_\_\_\_.

a. negative, gaining 1 electron b. positive, by losing 1 electron

c. negative, gaining 2 electrons d. positive, by losing 2 electrons

49. Noble gases are mainly known for their

a. extreme chemical reactivity. b. extreme sensitivity to light.

c. inability to react with other elements. d. radioactive nature.

*Bonding Review Questions.*

\_\_\_\_\_ 1. An ionic bond is formed by

a. losing electrons only. b. gaining electrons only.

c. sharing electrons. d. transferring electrons.

\_\_\_\_\_ 2. A covalent bond is formed by

a. losing electrons only. b. gaining electrons only.

c. sharing electrons. d. transferring electrons.

\_\_\_\_\_ 3. An ionic bond usually forms between

a. any two atoms. b. two metals.

c. two nonmetals. d. a metal and nonmetal.

\_\_\_\_\_ 4. A covalent bond usually forms between

a. any two atoms. b. two metals.

c. two nonmetals. d. a metal and nonmetal.

\_\_\_\_\_ 5. If the electronegativity difference between two atoms is between 0.0 and 0.4, the bond is considered to be

a. ionic. b. polar covalent. c. nonpolar covalent. d. nonexistent.

\_\_\_\_\_ 6. If the electronegativity difference between 2 atoms is between 0.4 and 1.67, the bond is considered to be

a. ionic. b. polar covalent. c, nonpolar covalent. d. nonexistent.

\_\_\_\_\_ 7. If the electronegativity difference between two atoms is greater than 1.67, the bond is considered to be

a. ionic. b. polar covalent. c. nonpolar covalent. d. nonexistent.

\_\_\_\_\_ 8. Why do atoms form chemical bonds?

a. they just do. b. to achieve lower potential energy

c. to achieve higher potential energy. d. bonds don’t really exist.

\_\_\_\_\_ 9. When atoms form chemical bonds, what determines the number of electrons they gain, lose, or share?

a. Atoms form bonds in order to attain a noble gas configuration (octet rule).

b. Atoms form bonds in order to have the most valence electrons possible.

c. Atoms form bonds in order to have the least valence electrons possible.

d. Atoms only bond with “friendly atoms,” the number of electrons is irrelevant.

\_\_\_\_\_ 10. When the electrons are shared equally between two atoms, the bond is

a. ionic. b. polar covalent. c. nonpolar covalent. d. imaginary.

\_\_\_\_\_ 11. When the electrons are unequally shared between two atoms, the bond is

a. ionic. b. polar covalent. c. nonpolar covalent. d. imaginary.

\_\_\_\_\_ 12. A magnesium atom will form a \_\_\_ by \_\_\_.

a. anion, losing 2 electrons b. cation, losing 2 electrons

c. anion, gaining 2 electrons d. cation, gaining 2 electrons

\_\_\_\_\_ 13. An iodine atom will form a \_\_\_ by \_\_\_.

a. anion, losing 1 electron b. cation, losing 1 electron

c. anion, gaining 1 electron d. cation, gaining 1 electron

\_\_\_\_\_ 14. Which of the following is characteristic of ionic compounds?

a. solids at room temperature b. do not conduct electricity in aqueous solution

c. low melting points d. representative unit is called a molecule

\_\_\_\_\_ 15. Which of the following is not characteristic of covalent compounds?

a. can be solids, liquids, or gases at room temperature

b. do not conduct electricity in aqueous solution

c. representative particle is called a molecule

d. high melting points

\_\_\_\_\_ 16. What is the name given to the electrons in the highest occupied energy level of an atom?

a. orbital electrons b. valence electrons c. anions d. d. cations

\_\_\_\_\_ 17. In forming chemical bonds, atoms attain \_\_\_\_\_.

a. a state of lower potential energy. c. greater stability.

b. the electron configuration of noble gas atoms. d. all of the above

\_\_\_\_\_ 18. When potassium and sulfur react to form ionic bonds, how many potassium atoms are needed to form a neutral compound?

a. 1 b. 2 c. 3 d. 4

\_\_\_\_\_ 19. When aluminum and oxygen react to form ionic bonds, how many oxygen atoms are needed to form a neutral compound?

a. 1 b. 2 c. 3 d. 4

\_\_\_\_\_20. An ionic compound is \_\_\_\_\_.

a. a salt b. held together by ionic bonds.

c. composed of anions and cations. d. all of the above

\_\_\_\_\_ 21. Under what conditions can potassium bromide conduct electricity?

a. only when melted b. only when dissolved in water

c. only when it is in crystal form d. only when melted or dissolved in water

\_\_\_\_\_ 22. Why do atoms share electrons in covalent bonds?

a. to become ions and attract each other. b. to attain a noble gas electron configuration.

c. to become more polar. d. to increase their atomic numbers

\_\_\_\_\_ 23. Which of the following is the name given to the pairs of valence electrons that do not participate in bonding in diatomic oxygen molecules?

a. unvalenced pair b. outer pair c. inner pair d. lone pair

\_\_\_\_\_ 24. Which elements can form diatomic molecules joined by a single covalent bond?

a. hydrogen only b. halogens only

c. halogens and members of the oxygen group only d. hydrogen and halogens only

\_\_\_\_\_ 25. How many valence electrons does an atom of any halogen have?

a. 1 b. 2 c. 4 d. 7

\_\_\_\_\_ 26. How many electrons are shared in a double covalent bond?

a. 2 b. 4 c. 6 d. 8

\_\_\_\_\_ 27. Which of the following elements can form diatomic molecules held together by triple covalent bonds?

a. hydrogen b. oxygen c. fluorine d. nitrogen

\_\_\_\_\_ 28. How many electrons does a nitrogen atom need to gain to attain a noble gas configuration?

a. 1 b. 3 c. 5 d. 7

\_\_\_\_\_ 29. A molecule with a single covalent bond is \_\_\_\_\_.

a. CO2 b. Cl2 c. CO d. N2

\_\_\_\_\_ 30. How many lone pairs of electrons does the nitrogen atom in ammonia possess?

a. 1 b. 2 c. 3 d. 4

\_\_\_\_\_ 31. How many lone pairs of electrons are there in a molecule of hydrogen iodide?

a. 1 b. 2 c. 3 d. 6

\_\_\_\_\_ 32. How many covalent bonds are there in PCl3?

a. 1 b. 2 c. 3 d. 4

\_\_\_\_\_ 33. The shape of the carbon tetrachloride molecule, CCl4, is called \_\_\_\_\_.

a. tetrahedral b. trigonal pyramidal c. trigonal planar d. bent

\_\_\_\_\_ 34. Which of the following covalent bonds is the most polar?

a. C-C b. C-H c. C-Cl d. C-S

\_\_\_\_\_35. What bonds are present in the Lewis dot structure for ozone, O3?

a. 2 single bonds b. 2 double bonds

c. 1 single and 1 double bond d. 2 ionic bonds

\_\_\_\_\_36. In a polar covalent bond, the electrons \_\_\_\_\_.

1. Are shared equally between the two atoms.
2. Are transferred from the most electronegative atom to the least electronegative atom
3. Are attracted to the most electronegative atom
4. Are attracted to the least electronegative atom

\_\_\_\_\_37. Polyatomic ions are

a. metal atoms with a positive charge b. nonmetal atoms with a negative charge

c. neutral covalent molecules d. covalently bonded atoms with a charge

\_\_\_\_\_38. Which of the following contains no lone pairs of electrons?

a. PH3 b. BH3 c. AsF3 d. BF3

\_\_\_\_\_39. The carbon dioxide molecule, CO2, contains \_\_\_\_\_ and is a \_\_\_\_\_.

a. polar bonds, nonpolar molecule b. nonpolar bonds, nonpolar molecule

c. polar bonds, polar molecule d. nonpolar bonds, polar molecule

\_\_\_\_\_40. If two covalently bonded atoms are identical, the bond is identified as

a. nonpolar covalent b. polar covalent c. nonionic d. ionic

\_\_\_\_\_41. A chemical bond resulting from the electrostatic attraction between positive &negative ions is called a(n)

a. covalent bond b. ionic bond c. charge bond d. dipolar bonds

\_\_\_\_\_42. Most chemical bonds are

a. purely ionic b. purely covalent

c. partly ionic and partly covalent d. metallic

\_\_\_\_\_43. In which of these compounds is the bond between the atoms NOT a nonpolar covalent bond?

a. Cl2 b. H2 c. HCl d. O2

\_\_\_\_\_44. In writing Lewis structures, each nonmetal atom except hydrogen and boron should be surrounded by

a. 2 electrons b. 4 electrons c. 8 electrons d. 10 electrons

\_\_\_\_\_45. In writing a Lewis structure for a polyatomic ion, one electron must be added for each unit of

a. positive charge b. mass c. polarity d. negative charge

\_\_\_\_\_46. In writing a Lewis structure for a polyatomic ion, one electron must be removed for each unit of

a. positive charge b. mass c. polarity d. negative charge

\_\_\_\_\_47. The substance whose Lewis structure indicates three covalent bonds is

a. H2O b. CH2Cl2 c. NH3 d. CCl4

\_\_\_\_\_48. How many electrons must be used in the Lewis structure of the hydroxide ion, OH-1?

a. 1 b. 8 c. 9 d. 10

\_\_\_\_\_49. Methane, CH4, contains how many covalent bonds?

a. 3 b. 4 c. 5 d. 8

\_\_\_\_\_50. The structure of the methane molecule, CH4 is described as

a. linear b. bent c. trigonal pyramidal d. tetrahedral

\_\_\_\_\_51. The methane molecule, CH4, is a

a. polar molecule b. nonpolar molecule c. crazy molecule d. sane molecule

\_\_\_\_\_52. The shape of the hydrogen cyanide molecule, HCN, is described as

a. linear b. bent c. trigonal planar d. trigonal pyramidal

\_\_\_\_\_53. The hydrogen cyanide molecule, HCN is a

a. polar molecule b. nonpolar molecule c. crazy molecule d. sane molecule

\_\_\_\_\_54. The shape of the chlorate ion, ClO3-1, is described as

a. trigonal pyramidal b. trigonal planar c. linear d. tetrahedral

\_\_\_\_\_55. The shape of the sulfur trioxide molecule, SO3, is described as

a. linear b. bent c. trigonal planar d. trigonal pyramidal

\_\_\_\_\_56. The sulfur trioxide molecule, SO3, is a

a. polar molecule b. nonpolar molecule c. crazy molecule d. sane molecule

\_\_\_\_\_57. The shape of the hydrogen sulfide molecule, H2S, is described as

a. trigonal planar b. trigonal pyramidal c. bent d. linear

\_\_\_\_\_58. The H2S molecule is a

a. polar molecule b. nonpolar molecule c. crazy molecule d. sane molecule

\_\_\_\_\_59. The shape of the arsenic trifluoride molecule, AsF3, is described as

a. linear b. tetrahedral c. trigonal planar d. trigonal pyramidal

\_\_\_\_\_60. The AsF3 molecule is a

a. polar molecule b. nonpolar molecule c. crazy molecule d. sane molecule