# Periodic Table Unit Review

CP/Honors Chemistry Name: Key

1. Define:
   a. *Period*
   b. *Group (or Family)*
   c. *Alkali Metals*
   d. *Alkaline Earth Metals*
   e. *Halogens*
   f. *Noble gases*
   g. *Transition Elements*
   h. *Metalloids (Semi-Metals)*
   i. *Valence Electrons*
   j. *Electronegativity*
   k. *Ionization Energy*
   l. *Electron Affinity*
   m. *Atomic Radius*
   n. *Isoelectronic*

Groups are related to # of valence electrons
Periods correspond to the valence shell
Blocks indicate the subshells that the electrons are located in.

2. Explain how the structure of the periodic table is related to the location of electrons within the atom.
3. Explain why elements in the same group have similar properties. Elements in the same group have the same # of valence electrons. The # of VE. impact the properties.
4. What happens to the atomic radius from left to right in a period? Why?
5. What happens to the atomic radius from top to bottom in a group? Why?
6. What happens to the ionization energy from left to right in a period? Why?
7. What happens to the ionization energy from top to bottom in a group? Why?
8. What happens to the electronegativity from left to right in a period? Why?
9. What happens to the electronegativity from top to bottom in a group? Why?
10. What happens to the radius of an atom when it gains an electron to become a negative ion? Radius increases
11. What happens to the radius of an atom when it loses an electron to become a positive ion? Radius decreases
12. Study the differences in properties of metals and nonmetals.
13. What are some examples of metalloids? Where are they found on the periodic table? Boron, Silicon, Germanium, Arsenic, Antimony, Tellurium, Astatine. Found along the "staircase" between the metals and nonmetals.
14. What are some examples of metals? Where are they found on the periodic table?
15. What are some examples of nonmetals? Where are they found on the periodic table? C, N, S, He, Xe, etc. Found in the top-right corner (and hydrogen).

14. Li, Ra, Cu, U, etc. Found in the lower-left of the periodic table. (Most elements are metals.)

on other sheet

16. Put these metals in order from most reactive to least reactive: potassium, cesium, sodium, lithium, rubidium Cs, Rb, K, Na, Li

17. Which Group has elements that exist as solids, liquids and gases at standard conditions? Group 17.

18. What is the most common ion for the following elements? How many protons and electrons are in the most common ions of these elements? Which noble gas is isoelectronic with the ion formed?

a. Barium $Ba^{+2}$; 56p, 54e⁻; Xe

b. Nitrogen $N^{-3}$; 7p, 10e⁻; Ne

c. Aluminum $Al^{+3}$; 13p, 10e⁻; Ne

d. Chlorine $Cl^{-}$; 17p, 18e⁻; Ar

19. If barium ions were bonded with chlorine ions, what would the resulting formula be for the ionic compound? $BaCl_2$

20. If aluminum ions were bonded with nitrogen ions, what would the resulting formula be for the ionic compound? AlN

21. Explain why oxygen has a slightly lower first ionization energy than nitrogen.

22. Explain why aluminum has a slightly lower first ionization energy than magnesium.

23. State the trends on the periodic table for electron affinity. across period from left → right; E.A. becomes a larger negative number. down a group; E.A. becomes a smaller negative number.

24. Atom X has an electron affinity of -100 kJ/mol whereas Atom Z has an electron affinity of -75 kJ/mol. Which atom is most likely to gain an electron? X.

21. O: [He] $\underset{2s}{\uparrow\downarrow}$ $\underset{2p}{\uparrow\downarrow\ \uparrow\ \uparrow}$

N: [He] $\underset{2s}{\uparrow\downarrow}$ $\underset{2p}{\uparrow\ \uparrow\ \uparrow}$

The electron being removed from oxygen is sharing an orbital with another electron whereas in nitrogen all the electrons in the 2p subshell have their own orbitals. There are repulsive forces between electrons that share an orbital making it easier to remove the electron.

22. Mg: $1s^2 2s^2 2p^6 3s^2$

Al: $1s^2 2s^2 2p^6 3s^2 3p^1$

The outermost electron in Magnesium is in the 3s sublevel. These electrons are shielded by the 10 electrons in the 1s, 2s, and 2p sublevels. The outermost electron in Aluminum is in the 3p sublevel. This electron is shielded by the 12 electrons in the 1s, 2s, 2p, and 3s sublevels. The greater degree of shielding makes it easier to remove the electron from Al.

# Periodic Table Unit Review

4. A.R. decreases. The more protons in the nucleus, the greater the attractive force pulling electrons closer + decreasing the radius.
5. A.R. increases. As the period # increases, the # of energy levels increase causing the radius to increase.
6. I.E. increases. The more protons in the nucleus, the greater the attractive force holding on to the electrons require more energy to remove them.
7. I.E. decreases. The greater the # of energy levels, the easier it is to remove an electron because the outer energy levels are more highly shielded by the electrons located in the inner shells.
8. EN increases. The greater the # of protons in the nucleus the greater the attraction of the nucleus on the electrons in a chemical bond.
9. EN decreases. The greater the # of energy levels, the more highly shielded the nucleus will be from the electrons in a chemical bond.

# Bonding Review Sheet
# H. Chem

**Terms:**

Ionic bond
covalent bond
metallic bond
crystal lattice
electronegativity
Lewis dot diagram
octet rule
valence electrons
single bond
double bond
triple bond
multiple bond
resonance
bond energy
molecule
network solid
ionic solid
conductivity
solubility
hardness
VSEPR
tetrahedral
trigonal pyramidal
bent
linear
trigonal planar
intermolecular forces
hydrogen bonds
dipole-dipole forces
London Dispersion forces
polar bonds
nonpolar bonds
polar molecules
dipole
nonpolar molecules

For each of the following:

a. Draw the Lewis Dot Diagram. (Show resonance structures when appropriate)
b. Determine the shape.
c. State whether each molecule is polar or nonpolar. (If it is polar, sketch the shape and write a δ+ and δ- showing the polarity.)
d. Determine the type of intermolecular force.
e. Determine the hybridization of the central atom (in #4-12). ← Not needed for Fall 2014

1. $H_2$
2. $N_2$
3. HF
4. $SO_3$
5. $CO_2$
6. $NH_3$
7. HCN
8. $NH_4^+$
9. $CBr_4$
10. $SF_6$
11. $NO_2^-$
12. $PF_5$

See other sheet.

13. What types of atoms form ionic bonds? Covalent bonds? Metallic bonds?

1 metal + 1 nonmetal ← ionic; 2 or more nonmetals ↙ covalent; → metals only

14. What are the rules for using electronegativity to determine bond type?

Generally; If E.N. difference is 0–0.3 = nonpolar covalent
0.3–1.7 = polar covalent
+1.7 = ionic

15. Use the electronegativity rules to determine bond type for each of the following.

a. Mg-O (1.2, 3.5) Ionic
b. Ca-Cl (1.0, 3.0) Ionic
c. H-O (2.1, 3.5) Polar Covalent
d. H-H Nonpolar Covalent
e. S-O (2.5, 3.5) Polar covalent
f. C-Cl (2.5, 3.0) Polar covalent
g. C-S (2.5, 2.5) Nonpolar covalent

16. Which bond from question #15 has the greatest ionic character? Which has the greatest covalent character?

Mg-O
both H-H and C-S

17. List three characteristics of metals and three characteristics of nonmetals.

Shiny (lustrous)
conducts heat + electricity well
malleable
ductile

dull in appearance
brittle
poor conductors

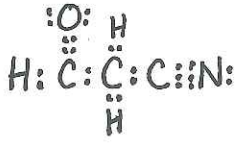
18. Complete the chart below:

| | Covalent (molecular) | Ionic | Metallic | Network covalent |
|---|---|---|---|---|
| Melting point | Relatively Low | Relatively high | Variable | Very high |
| Electrical conductivity as a solid | Poor | Poor | Very Good | Poor |
| Solubility in water | Sometimes Soluble | Sometimes Soluble | Not Soluble | Not Soluble |
| Electrical conductivity when dissolved in water | Poor | Very Good | ↓ | ↓ |
| hardness | Relatively Soft | hard | Variable | Very hard |

Not covered Fall 2014

19. Describe the difference between a sigma bond and a pi bond.

20. In the following structure, how many σ-bonds are present? π-bonds?



# Bonding Review Sheet – H. Chem

1. a) H:H  c) nonpolar
   b) linear  d) London Dispersion

2. a) :N⋮⋮N:  c) nonpolar
   b) linear  d) London Dispersion

3. a) H:F:  c) polar  $\overset{\delta+}{H}-\overset{\delta-}{F}$
   b) linear  d) Hydrogen Bonding

4. a) $\left[:\ddot{O}:S::\ddot{O}: \atop :\ddot{O}:\right] \longleftrightarrow \left[\ddot{O}::S:\ddot{O}: \atop :\ddot{O}:\right] \longleftrightarrow \left[:\ddot{O}:S:\ddot{O}: \atop :\ddot{O}:\right]$
   b) trigonal planar
   c) nonpolar  d) London Dispersion

5. a) $\ddot{O}::C::\ddot{O}$  c) nonpolar
   b) linear  d) London Dispersion

6. a) $H:\ddot{N}:H$ (H below)  c) polar $\overset{\delta-}{N}(\delta+)(\delta+)$
   b) trigonal pyramidal  d) Hydrogen Bonding

7. a) H:C⋮⋮N:  c) Polar  $\overset{\delta+}{O}-C-\overset{\delta-}{N}$
   b) linear  d) Dipole-Dipole

8. a) $\left[\begin{matrix} H \\ H:\ddot{N}:H \\ H \end{matrix}\right]^{+}$  c) Nonpolar
   b) tetrahedral  d) London Dispersion

9. a) $:\ddot{Br}: \\ :\ddot{Br}:C:\ddot{Br}: \\ :\ddot{Br}:$  c) Nonpolar
   b) tetrahedral  d) London Dispersion

10. a) $:\ddot{F}: \quad :\ddot{F}: \\ :\ddot{F}-S-\ddot{F}: \\ :\ddot{F}: \quad :\ddot{F}:$  b) octahedral
    c) Nonpolar
    d) London Dispersion

11. a) $\left[:\ddot{O}:N::\ddot{O}:\right]^{-} \longleftrightarrow \left[\ddot{O}::N:\ddot{O}:\right]^{-}$
    b) bent
    c) Polar  d) Dipole-Dipole
    $\overset{\delta-}{O}-\overset{\delta+}{N}-\overset{\delta-}{O}$

12. a) $:\ddot{F}: \\ :\ddot{F}-P-\ddot{F}: \\ :\ddot{F}: \quad :\ddot{F}:$  b) trigonal bipyramidal
    c) Nonpolar
    d) London Dispersion.