**Ms. Perry’s HPS** – Step by Step Guided Review

Key is online at <http://shakerscience.weebly.com> under our last unit – Unit 14 Stars and Solutions

Reference sheet is at the end, will look almost identical on the final. Record your answers on a separate sheet of paper, so you can reuse this as many times as you need!

Unit 8 Atomic Theory and Structure

Section 1 – *Scientists and Thinkers with their models*

* Explain the difference between Democritus and John Dalton.
* What did J.J. Thompson find? What was his model like?
* What did Ernest Rutherford discover? How?
* What did Chadwick find?
* Why was Bohr’s model so important?
* What does the contemporary Electron Cloud model say about atomic structure?

Section 2 – *Subatomic particles*

* Define isotope, cation, and anion.
* An atom has 6 protons, 6 neutrons, and 5 electrons. What term describes this type of particle? Which element is it? What is its mass? What is its charge?
* The atom from the previous question loses a neutron and an electron. What about it changes? What about it DOES NOT change?
* The atom from the previous question undergoes alpha decay and loses an alpha particle. What is an alpha particle made of? What about the atom changes?
* Write the isotope potassium-41 in nuclear notation. (Check your notes or Google it if you don’t remember!)
* A neutral oxygen atom has a mass of 17. What are the subatomic particles counts of this isotope?

Section 3 – *Average Atomic Mass*

* Why are the masses listed on the periodic table (generally) not whole numbers?
* What does it mean that an isotope has a higher “percent abundance” than another?
* Calculate the average atomic mass of chlorine if the following is true: 35Cl is 75.77% abundant while 37Cl is 24.23% abundant.

Section 4 – *Electron configuration*

* Write the electron configuration of bromine
* Draw the orbital diagram for bromine.
* Copper and chromium are “exceptions”. How?
* Draw the “noble gas” orbital diagrams for copper and chromium.

Section 5 – *The Bohr Model* *and the Jumping Electrons*

* What happens when an electron “jumps up” from energy level 1 to 2? (No calculations yet!)
* What happens when an electron falls back down to ground state?
* Calculate the energy associated of a photon that is emitted when an electron falls from energy level 6 to energy level 2.
* See the previous – what is the frequency of that light?
* See the previous – what is the wavelength of that light?

Unit 9 Classification of Matter and Periodic Trends

Section 1 – *Physical and Chemical*

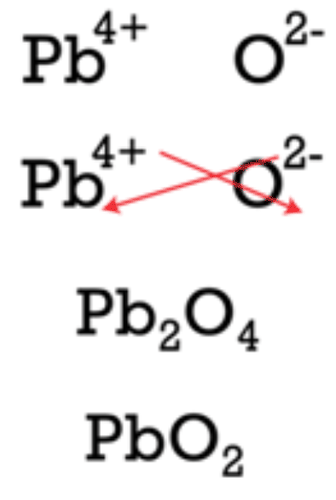
* You boil water to make some pasta. Is this a physical or a chemical change? Why?
* You burn a piece of chemistry homework because you are salty. Is this a physical or a chemical change? Why?
* You dissolve a salt (because you are still salty) in your coffee. Is this a physical or a chemical change? Why?
* You observe your foil has turned black and is bubbling after you wrapped your tomato with it. Is this a physical or a chemical property you are observing? Justify.
* Describe the difference between homogeneous and heterogeneous mixtures.
* Describe the differences between solutions, suspensions, and colloids.

Section 2 – *The Trends*

* Describe the octet rule.
* Metals lie to the left of the periodic table, nonmetals to the right, and metalloids along a staircase line. With one exception, which is? Additionally, when metals and nonmetals become charged, what types of ions do they generally become? Justify why.
* How many valence electrons does magnesium have? Iodine? Helium? Carbon? Sulfur?
* Why does atomic radius increase as you go down a group? Why does it decrease as you go across?
* Why does ionization decrease as you go down a group? Why does it increase as you go across?
* Explain the differences in reactivity of metals in a group versus the halogens in group 7A.
* What do periods tell you about the elements in them?
* Draw Lewis dot diagrams for magnesium, aluminum, oxygen, and neon.
* Draw Lewis dot diagrams for the ions of sodium, nitrogen, and barium.

Unit 10 Chemical Bonding

lead (IV) oxide

Section 1 – *Ionic Bonding and Nomenclature*

* Describe what happens in an ionic bond .
* List several characteristics of ionic compounds.
* What is a polyatomic ion?
* Name the following compounds: Na2O, MgCl2, CuSO4, Ba3(PO4)2, KOH, AlN, AlF3
* Write the formulas for the following compounds BUT WAIT. Follow these steps if you forgot:
  + Find the charge of the cation (usually the metal, named first. If transition, Roman numeral is the charge)
  + Find the charge of the anion (ends in -ide, monatomic meaning its one atom, ends in anything else, find it on the polyatomic sheet)
  + Write down the ions with their charges
  + Flip the charges as subscripts, reducing to simplest whole number ratio. See example. If you have a polyatomic ion and the subscript for it is not one, put it in parentheses.

Iron (III) oxide, lead (II) carbonate, sodium sulfide, aluminum nitrate, potassium chlorate

Section 2 – *Covalent Bonding and Nomenclature*

* Describe what happens in a covalent bond.
* List several characteristics of covalent compounds.
* What is a double or triple covalent bond?
* What makes a covalent bond polar? Give an example.
* Name the following compounds: P4O10, SO3, NO2, CO, N2O, SF6, PI5
* Write the formulas for the following compounds BUT WAIT. Follow these steps if you forgot:
  + Keep the whole name of the first element, adding a prefix for the number of atoms present UNLESS there is only one.
  + Keep the stem of the name of the second element, adding -ide as the ending, and adding a prefix for the number of atoms present.

Sulfur monoxide, boron trifluoride, tetracarbon decahydride, dinitrogen pentoxide

Section 3 – *Lewis Dot structures Level 2 – Compounds and Molecular geometry*

* Draw the Lewis structure for the following compounds. BUT WAIT. Follow these steps if you forgot:
  + Count the total valence electrons present in the compound by adding up the valence electrons of all atoms present
  + Join the atoms, with the least electronegative at the center (excluding hydrogen—it’s never central)
  + Check everyone’s octet (everyone should have 8… except for who???)
  + Add lone pairs around any atom that still needs an octet.
  + Make sure you are checking you don’t use electrons you don’t have or that you use too few. Run out of electrons? Make some double or triple bonds!

CO2, CH3F, H2O, CN, O2, C2H4

* Refer to each of the Lewis structure you just drew. What are their molecular shapes?
* What about for NH3? Draw the structure, state the shape, AND say whether it is polar or not. Justify your answer.

Section 4 – *Metallic Bonding*

* Describe what happens in a metallic bond.
* Explain why metals have certain properties by relating it to their bonding type.

Unit 11 Chemical Reactions and Equations

Section 1 – *Percent composition with Empirical and Molecular Formulas*

* Distinguish between empirical and molecular formulas.
* If a compound is 7.19% phosphorus, and 92.81% bromine, what is the empirical formula?
* What is the molecular formula if the previous compound has a “true” or molecular mass of 431 g/mol?
* What is the mass percent of a compound with a formula of AlPO4?

Section 2 – *Molar Masses and Using them*

* What is the molar mass of a compound with a formula of AlPO4?
* How many moles are present if you have 2.3 x 1025formula units of AlPO4?
* How many moles are present if you have 16 g of AlPO4?
* How many grams are present if you have 0.035 mol of AlPO4?

Section 3 – *Types of Reactions*

* Give a real example of a synthesis (combination) reaction.
* Give a real example of a decomposition reaction.
* Give a real example of a cation single displacement reaction.
* Give a real example of an anion single displacement reaction.
* Give a real example of a double displacement reaction.
* Give a real example of a combustion reaction.
* Balance the following equation:

Al2(CO3)3 + HNO3 🡪 Al(NO3)3 + CO2 + H2O

Unit 12 Stoichiometric Analysis

Section 1 – *Proudly* *predicting products*

* Will a reaction happen between Aluminum metal and sodium chloride? Why or why not? Write the balanced equation if so.
* Will a reaction happen between chlorine gas and sodium iodide? Why or why not? Write the balanced equation if so.
* Will a reaction happen between zinc metal and copper (II) chloride? Why or why not? Write the balanced equation if so.
* Will a precipitate form from the reaction between iron (III) chloride and potassium hydroxide? Write the balanced equation with state symbols.

Section 2 – *Quantitative analysis of equations*

* Using the balanced equation between sodium hydroxide and iron (II) fluoride, how many grams would be present if you wanted to react 10.0 g of sodium hydroxide completely?
* If you are given 15.5g of iron (II) fluoride to work with in the lab, which is the limiting reactant?
* How much of the excess is used and leftover?
* How much of each product can be produced?

Unit 13 Stars and Solutions

Section 1 – *Life cycle of a star and the HR diagram*

* List the different types of stars.
* What is happening in a star? Why is it luminous?
* What is different about massive stars and stars like our sun?
* What is peculiar about the graphing of an HR diagram?
* What can you tell about a star when plotting it on the HR diagram?

Section 2 – *Making Solutions with* *Molarity*

* What is molarity? What does it mean to be concentrated versus dilute?
* What is a solute? What is a solvent?
* What is a stock solution?
* What is the molarity of a solution that was made from 50 mL of water and 3.1 g of potassium iodide.
* What mass is needed of potassium nitrate to make a 0.1 M solution that has a volume of 0.300 L?

Section 3 – *pH, Acids, and Bases*

* What is the pH of a solution that has a concentration of 0.01 M hydronium ions?
* Describe the pH scale and gives examples of acids and bases.
* What is the auto-ionization of water?
* Give some other important properties of water.

Section 4 – *Solubility and Solutions*

* Describe what “like dissolves like” means.
* Why is water considered the “universal solvent”?
* How does soap work?
* Distinguish between unsaturated, saturated, and supersaturated solutions.
* What is an “equilibrium”?
* What is a fishkill?
* How is solubility graphed typically? How does is change for “salts” (ionic compounds)?