AP Chemistry Problems

I. Physical and Chemical Changes, Matter and Energy Uncertainty, Calculations, Dimensional Analysis

1. Use appropriate metric prefixes to write the following measurements without use of exponents: (a) 6.5×10^{-9}cm; (b) 12.5×10^{-8}s ; (c) 3.5×10^{3}L.

2. Convert: (a) 2.52×10^{3}kg to g ; (b) 0.0023 mm to nm ; (c) 6.25×10^{-4}s to ms.

3. What type of quantity (for example length, volume, density) the following units indicate: (a) mL; (b) cm^2; (c) mm^3; (d) mg/L; (e) ps; (f) nm; (g) K?

4. (a) A cube of ruthenium metal 1.5 cm on a side has a mass of 42.0 g. what is its density in g/cm^3? Will ruthenium metal float on water? (Materials that are less dense than water will float.) (b) The density of bismuth metal is 9.8 g/cm^3. What is the mass of a sample of bismuth that displaces 65.8 mL of water? (c) The density of a piece of metal alloy is 4.35 g/cm^3. What is the volume of 53.7 kg of this metal?

5. (a) The temperature on a warm summer day is 93°F. What is the temperature in °C? (b) The melting point of potassium iodide (a salt) is 681°C. What is the temperature in kelvins? (c) Carbon tetrachloride freezes at 250.2 K. What is its freezing point in degrees Fahrenheit?

6. Indicate the number of significant figures in each of the following measured quantities: (a) 1.689×10^{-3}km; (b) 0.0234 m^2; (c) 7,194,300 cm; (d) 435.983 K; (e) 204.080 g.

7. Round each of the following numbers to three significant figures, and express the result in standard exponential notation: (a) 143700; (b) 0.09750; (c) 890,000; (d) 6.764×10^{4}; (e) 33,987.22; (f) –6.5559.

8. Carry out the following operations, and express the answer with the appropriate number of significant figures:

   (a) 320.55 – (6104.5/2.3) (b) $[(285.3 \times 10^{5}) - (1.200 \times 10^{3})] \times 2.8954$ (c) (0.0045 x 20,000.0) + (2813 x 12) (d)

   863 x [1255 – (3.45 x 108)]

9. Carry out the following conversions: (a) 16.2 ft to m; (b) 5.44 qt to mL; (c) 45.7 in/s to km/hr; (d) 23.5 yd^3 to m^3

10. Read the following description of the element zinc, and indicate which are physical properties and which are chemical properties. Zinc is a silver-gray colored metal, which melts at 420°C. When zinc granules are added to dilute sulfuric acid, hydrogen is given off and the metal dissolves. Zinc has a hardness on the Mohs scale of 2.5 and a density of 7.13 g/cm^3 at 25°C. It reacts slowly with oxygen gas at elevated temperatures to form zinc oxide, ZnO.

11. A match is lit and held under a cold piece of metal. The following observations are made: (a) The match burns. (b) The metal gets warmer. (c) Water condenses on the metal. Which of these occurrences are due to physical changes, and which are due to chemical changes?

II. Subatomic particles, isotopes, ions, average atomic mass

1. Hydrogen sulfide is composed of two elements: hydrogen and sulfur. In an experiment, 6.500 g of hydrogen sulfide is fully decomposed into its elements. (a) If 0.384 g of hydrogen are obtained in this experiment, how many grams of sulfur must be obtained?

2. How many protons, neutrons and electrons are in the following atoms: (a) ^{40}Ar (b) ^{55}Mn (c) ^{65}Zn

3. Fill in the gaps in the following table assuming each column represents a neutral atom:

Symbol	^{46}Ti		
Protons		45	
Neutrons		58	18
Electrons			16
Atomic no.			
Mass no.			

4. Give the number of protons, electrons and neutrons for the following isotopes:

a. Oxygen-16 b. Bismuth-209 c. Carbon-12

5. Give the number of protons and electrons for the following ions:

a) Potassium Ion (cation) K^+

b) Ion (cation) Al^{+3}

c) Sulfur Ion (anion) S^{2-}

6. Which symbol correctly represents an element (D) whose atom contains 15 protons and 20 neutrons?
a) $^{20}_{15}D$ b) $^{35}_{15}D$ c) $^{15}_{35}D$ d) $^{22}_{15}D$

7. Rb^{+1}, Se^{-2}, Kr, Br^{-1} all have the same number of ______________.

8. How many protons, neutrons, and electrons are in the $^{140}_{58}Ce^{+3}$ ion?
Protons = _____ electrons = ______ neutrons = ______

9. If an atom has 15 protons, 17 electrons, and 20 neutrons, the correct symbol would be ______.
a) $^{35}_{15}P^{-2}$ b) $^{32}_{15}P^{+2}$ c) $^{35}_{15}P^{-5}$ d) $^{35}_{17}Cl^{-2}$ e) $^{35}_{15}P^{+5}$

10. Write the correct symbol, with both superscript and subscript, for each of the following (use the list of elements on the front inside cover): (a) the isotope of sodium with mass 23 (b) the nuclide of vanadium that contains 28 neutrons (c) the isotope of chlorine with mass 37 (d) the nuclide of magnesium that has an equal number of protons and neutrons.

11. ^{235}U and ^{238}U is an example of isotopes. Explain why

12. The element magnesium consists of three natural occurring isotopes with masses 23.98504, 24.98584, and 25.98259 amu. The relative abundances of these three isotopes are 78.70, 10.13, and 11.17 percent, respectively. From these data calculate the average atomic mass of magnesium.

13. There are three stable isotopes of a hypothetical element: one isotope has a mass of 27.977amu and is present 92.21%; another isotope has a mass of 28.976 amu and is present 4.70%, and the other isotope has a mass of 29.974 amu. What is the average atomic mass of this element?

14. Silver has two isotopes, one with 60 neutrons and the other with 62 neutrons. Give the chemical notation for each of these isotopes and calculate the relative abundance for each isotope given that the average atomic mass for silver is 107.87 amu.

15. Locate each of the following elements in the periodic table; indicate whether it is a metal, metalloid, or nonmetal; and give the name of the element: (a) Li (b) Sc (c) Ge (d) Yb (e) Mn (f) Au (g) Te.
16. For each of the following elements, write its chemical symbol, determine the name of the group to which it belongs, and indicate whether it is a metal, metalloid, or nonmetal: (a) potassium (b) iodine (c) magnesium (d) argon.

III. Electron Configuration & Orbital Diagrams

1. Write the orbital notation for each of the following elements.

Example: N (7 e-) __ __ __ __ __

a. Magnesium ______________________________

b. Oxygen ______________________________

c. Aluminum ______________________________

2. Write the complete and abbreviated (noble gas) electron configurations of the following elements. State how many electrons are unpaired and how many valence electrons the element has.

a. Oxygen ____

b. Chlorine ____

c. Sodium ____

d. Aluminum ____

3.Determine what elements are denoted by the following electron configurations:

____**a.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

____**b.** $1s^2 2s^2 2p^6 3s^2 3p^2$–

___ **c.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^3$

___ **d.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

____**e.** $1s^2 2s^2 2p^6$

___ f. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^1$

___**g.** $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^6$

4. Which of the following configurations is an impossible and explain why

____a. $1s^2 2s^2 2p^6 3s^2 3p^8 s^2$

____b. $1s^2 2s^2 2p^6 3s^2 3p^2$

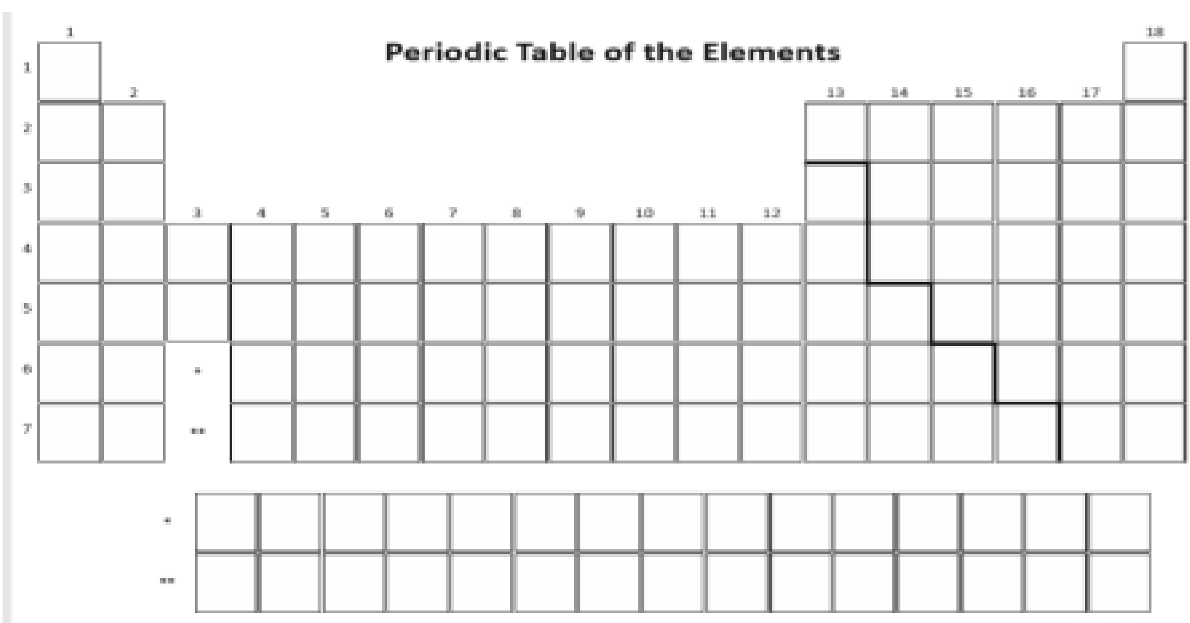
5. $1s^2 2s^2 2p^6 3s^2 3p^6 6s^1$ represents an excited atom. Explain why.

6. Define the following: Hund's rule, Aufbau principle, Pauli Exclusion.

IV. Waves and Energy

1. A. Draw a *wave* of light with high frequency. B. Draw a *wave* of light with low frequency.
2. How many Joules of energy are there in one photon of orange light whose wavelength is 630nm?
3. If an X-ray machine emits E.M.R with a wavelength of 1.00×10^{-10} meters, what is the frequency?
4. A photon has a frequency (ν) of 2.68×10^6 Hz. Calculate its energy.
5. Calculate the energy (E) and wavelength (λ) of a photon of light with a frequency (ν) of 6.165×10^{14} Hz.
6. A certain light has a frequency of $6.26 \times 10^{14}\ s^{-1}$. What is its wavelength?
7. What is the wavelength of radiation whose frequency is 103.5MHz?
8. Rank the parts of the electromagnetic spectrum from lowest frequency (a) to highest (g):
9. Rank the parts of the electromagnetic spectrum from shortest wavelength (A) to longest (G):
10. What is the relationship between frequency and wavelength?
11. What is the relationship between frequency and energy?

V. Periodic trends



1. On the blank periodic table, color and label:

a. Alkali metals, alkaline metals, transition metals, nonmetals, metalloids, halogens, noble gases, inner transition metals b.

The "s" block, the "p" block, the "d" block, the "f" block

c. Draw arrows to show the following periodic trends acrros each period and down each group. Be sure to label which was trend is increasing and which is decreasing. A. atomic radius, ionization energy, electronegativity and electron affinity.

2. How many valence electrons do each of the following groups have?
IA ______ IIA_____ IIIA ______ VA ______ VIIA ______ VIIIA _______

3. The Successive ionization energies for a given element are listed below:
1st ionization 5000 kJ 2nd ionization 6890 kJ 3rd ionization 10500 kJ 4rd ionization 12000 kJ To what group does the element belong to? _________________

4. The Successive ionization energies for a given element are listed below:
1st ionization 1000 kJ 2nd ionization 2890 kJ 3rd ionization 3500 kJ 4rd ionization 4000 kJ To what group does the element belong to? _________________

5. Cations are ______than their neutral atoms while anion are _________than their neutral atoms.

6. The __________ states that atoms will gain, lose or share valence electrons in order to fulfill their outer layer.

7. Phosphorus, sulfur, and selenium are located near each other on the periodic table. Which of these elements is (a) the largest atom? (b) the atom with the highest ionization energy? (a) ________(b)________

8. Scandium, nickel, and lanthanum are located near each other on the periodic table. Which of these elements is (a) the largest atom? (b) the atom with the smallest ionization energy? (a) __________ (b) ___________

9. Which of the following atoms is smallest: (a) vanadium, chromium, or tungsten? ___________
(b) Which of these atoms has the highest electron affinity? _____________________________

10. Which of the following atoms is smallest: (a) nitrogen, phosphorus, or arsenic? ___________
(b) Which of these atoms has the smallest electronegativity? ___________________________

11. Which of the following is the largest: a potassium atom, a potassium ion with a charge of 1+, or a rubidium atom?
____

12. Which of the following is the largest: a chlorine atom, a chlorine ion with a charge of 1-, or a bromine atom?
____

13. Which of the following is the smallest: a lithium atom, a lithium ion with a charge of 1+, or a sodium atom?
____

14. Aluminum, silicon, and phosphorus are located near each other on the periodic table. Which of these elements is (a) the largest atom? (b) the atom with the highest ionization energy? (a) ________(b)________

15. Which atom in each pair has the largest atomic radius?
______ a. Li or K _____ b. Ca or Ni _____ c. Ga or B_____ d. O or C _____ e. Cl or Br
______ f. Be or Ba _____ g. Si or S _____ h. Fe or Au _____ i. S or Br _____ j. B or Cl

16. Which ion in each pair has the smaller atomic radius?
______a. Al^{+3} or P^{-3} ______ b. F^{-1} or S^{-2} _______ c. N^{-3}or O^{-2} _______ d. Na^{+1} or Ca^{+2}

_____e. O^{-2} or Cl^{-1} _____f. K^{+1} or Ga^{+3} ______ g. Fe^{+2} or Fe^{+3} ______ h. O^{-2} or F^{-1}

17. Which atom in each pair has the largest electron affinity?
_____a. Be or Ba ____b. Si or S ____ c. Fe or Au ____ d. S or Br ____ e. B or Cl

18. Elements in group IA form ions of ____charge; in group IIA form ions of ____charge; in group IIIA ___ ; in group VA form ions of ____charge; in group VIA form ions of ___charge; in group VIIA form ions of ___ charge.

19. $1s^2 2s^2 2p^6 3s^1$ has __ valence electrons.

20. $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^4$ has ____ valence electrons

VI. Ionic vs Molecular compounds
Naming Ionic and Covalent compounds

1. How many hydrogen atoms are in each of the following chemical formula: (a) carbon atoms in $C_2H_5COOCH_3$; (b) oxygen atoms in $Ca(ClO_3)_2$; (c) hydrogen atoms in $(NH_4)HPO_4$?

2. From the following list, find the groups of compounds that have the same empirical formula:

C_2H_2 , N_2O_4, C_2H_2, C_6H_6, NO_2, C_3H_6, C_4H_8

3. Write the empirical formula corresponding to each of the following molecular formulas: (a) S_4N_4 ; (b) C_7H_{15} ; (c) $C_6H_{10}O_2$; (d) P_4O_6 ; (e) $C_6H_{10}F_8$ (f) Si_3O_9.
4. Using the periodic table, predict the charges of the ions of the following elements: (a) Ba ; (b) La (c) Ga (d) S (e) Br.

5. Predict the chemical formulas of the compounds formed by the following pairs of ions: (a) Mg^{2+} and NO_3^- (b) Na^+ and CO_3^{2-} (c) Ba^{2+} and OH^- (d) NH_4 and PO_4^{3-} (e) Hg_2^{2+} and ClO_{3^-}

6. Which of the following are ionic, and which are molecular? (a) Sc_2O_3 (b) NaI (c) SCl_2 (d) $Ca(NO_3)_2$ (e) $FeCl_3$ (f) LaP (g) $CoCO_3$ (f) $(NH_4)_2SO_4$

7. Name the following ionic compounds: (a) AlF_3 (b) $Cu(NO_3)_2$ (c) $Mg(ClO_3)_2$ (d) $SrSO_3$ (e) $CoBr_2$ (f) SnI_2 (g) $Cr(NO_3)_3$ (h) $ZnHPO_4$ (i) $AgClO_4$ (j) $(NH_4)_2Cr_2O_7$

8. Give the chemical formula for each of the following ionic compounds: (a) Magnesium Nitride (b) Iron (II) Sulfite (c) Chromium (III) Carbonate (d) Calcium Hydride (e) Magnesium Bicarbonate (f) Potassium Hypochlorite (g) Copper (II) Acetate

9.The oxides of nitrogen are very important ingredients in determining urban air pollution. Name each of the following components: (a) N_2O (b) NO (c) NO_2 (d) N_2O_5 (e) N_2O_4

VII. VSEPR

For each of the following molecules:
1. draw structural formula 2. give molecular shape 3. give angles 4. state polarity 5. hybridization

a. OI_2 f. BeH_2
b NF_3 g. IBr_3
c. SeF_4 h. ArF_4
d. $AsCl_5$ i. HSiP
e. CH_3Cl j. CO_2

VIII. Mole Conversion

1. How many moles are 1.20×10^{25} atoms of phosphorous?

2. How many atoms are in 0.750 moles of zinc?

3. Find the number of moles of argon in 452 g of argon.

4 . Find the mass in 2.6 mol of lithium bromide.

5. What is the volume of 0.05 mol of neon gas at STP?

6. Determine the volume in liters occupied by 14 g of nitrogen gas at STP.

7. Find the mass, in grams, of 1.00×10^{23} molecules of N_2.

8. How many moles are in 75 grams of Iodine

9. How many grams of sodium are in 5.48×10^{23} atoms of sodium?

10. How many formula units are in 22 grams of calcium phosphate?

11. How many atoms of oxygen are in 5.00 moles of calcium hydroxide?

12. How many ions of sulfate are in 32.0 grams of aluminum sulfate?

IX. Percent by mass, Empirical,Molecular Formula and formula of a hydrate

1. Calculate the percentage by mass of the indicated element in the following compounds: (a) hydrogen in ammonium sulfate, $(NH_4)_2SO_4$, a substance used as nitrogen fertilizer, (b) platinum in $PtCl_2(NH_3)_2$. a chemotherapy agent called cisplatin.

2. (a) What is the difference between an empirical formula and a molecular formula? (b) Styrene, a hydrocarbon used to make Styrofoam cups and insulation, has an empirical formula of CH and a molar mass of 104 g/mol. What is its molecular formula?

3. What is the percentage of each element in lithium carbonate?

4. What is the empirical formula of a compound containing 2.16g Al, 3.85g S, and 7.68g of O?

5. a) Nicotine is a stimulant and an addictive chemical found in tobacco. An analysis of nicotine produces the following percent composition: 74.03% carbon, 17.27% nitrogen, and 8.70% hydrogen. What is the empirical formula of nicotine? b) Further tests show that the molar mass of nicotine is 162.23 g/mol. Given this information, what is the molecular formula of nicotine?

6. Washing soda, a compound used to prepare hard water for washing laundry, is a hydrate, which means that a certain number of water molecules are included in the solid structure. Its formula can be written as $Na_2CO_3 \cdot xH_2O$ per mole of Na_2CO_3. When a 2.558-g sample of washing soda is heated at 125°C, all the water of hydration is lost, leaving 0.948 g of Na_2CO_3. What is the value of x?

X. Chemical Equations

1. Convert these descriptions into balanced equations: (a) When sulfur trioxide gas reacts with water, a solution of sulfuric acid forms. (b) Boron sulfide, $B_2S_3(s)$, reacts violently with water to form dissolved boric acid, H_3BO_3. and hydrogen sulfide gas. (c) Phosphine, PH_3 (g), combusts in oxygen gas to form gaseous water and solid tetraphosphorus decoxide. (d) When solid mercury(II)nitrate, is heated, it decomposes to form solid mercury(II)oxide, gaseous nitrogen dioxide, and oxygen. (e) Copper metal reacts with hot concentrated sulfuric acid solution to form aqueous copper(II)sulfate, sulfur dioxide gas, and water.

2. Write a balanced chemical equation for the reaction that occurs when (a) styrene, $C_8H_8(l)$ burns in air, (b) lithium metal is added to water; (c) $SrCO_3(s)$ is decomposed upon heating.

3. Balance the following equations, and indicate whether they are combustion, composition (synthesis), or decomposition

reactions: (a) $Al(s) + Cl_2(g) \rightarrow AlCl_3(s)$ (b) $C_2H_4(g) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

(c) $Li(s) + N_2(g) \rightarrow Li_3N(s)$ (d) $PbCO_3(s) \rightarrow PbO(s) + CO_2(g)$

(e) $C_7H_8O_2(l) + O_2(g) \rightarrow CO_2(g) + H_2O(l)$

4. Directions: Write equations, predict products, balance, give type, name the compound when asked,

a. Iron(II) and calcium sulfide type of reaction: ________________ b. Nickel (III) hydroxide type of reaction: ________________ c. Fluorine and sodium chloride type of reaction: ________________ d. Copper (II) chlorate type of reaction:_________________
e. Zinc (II) sulfate and cobalt (III) hydroxide type of reaction:_________________ f. potassium and iodine type of reaction :________________ g. Lead (IV) carbonate type of reaction:_________________
h. Magnesium nitrate and ammonium fluoride type of reaction:________________ i. Manganese (III) sulfate type of reaction _________________ j. silver (I) and oxygen type of reaction ________________

XI. Stoichiometry, percent yield, limiting and excess

1. Hydrofluoric acid, HF (aq), cannot be stored in glass bottles because compounds called silicates in the glass are attacked by the HF(aq). For example, sodium silicate, Na_2SiO_3, reacts in the following way:

$$Na_2SiO_3(s) + 8HF(aq) \rightarrow H_2SiF_6(aq) + 2NaF(aq) + 3H_2O(l)$$

How many moles of HF are required to dissolve 0.50 mol of Na_2SiO_3? (b) How many grams of NaF form when 0.300 mol of HF reacts in this way? (c) How many grams of Na_2SiO_3 can be dissolved by 0.300 g of HF?

2. Calcium hydride reacts with water to form calcium hydroxide and hydrogen gas. (a) Write a balanced chemical equation for the reaction. (b) How many grams of calcium hydride are needed to form 5.0 g of hydrogen?

3. The complete combustion of octane, C_8H_{18}, a component of gasoline, proceeds as follows:

$$2\ C_8H_{18}(l) + 25\ O_2(g) \rightarrow 16CO_2(g) + 18\ H_2O(l)$$

How many moles are needed of O_2 to burn 1.50 mol of C_8H_{18}? (b) How many grams of O_2 are needed to burn 1.50 g of C_8H_{18}? (c) Octane has a density of 0.692 g/mL at 20°C. How many grams of O_2 are required to burn 1.00 gal. of C_8H_{18}?

4. One of the steps in the commercial process for converting ammonia to nitric acid involves the conversion of NH_3 to NO:

$$4NH_3(g) + 5O_2(g) \rightarrow 4NO(g) + 6H_2O(g)$$

In a certain experiment 2.50 g of NH_3 reacts with 2.85 g of O_2. (a) Which reactant is the limiting reactant? (b) How many grams of NO form? (c) How much of the excess reactant remains after the limiting reactant is completely consumed?

5. 16.48 grams of sulfur trioxide and 12.48 grams of water react to form H_2SO_4. What mass of sulfuric acids is produced? Identify the limiting reagent and the excess? How many grams of the excess is left over.

6. Potassium reacts with oxygen gas to produce potassium oxide. Determine the percent yield for the reaction between 8.92 grams of potassium and 3.28 grams of oxygen gas if 6.36 grams of potassium oxide are produced?

7. Identify the limiting reactant when 11.5 L of H_2S at STP is bubbled through a solution containing 22.0 grams of potassium hydroxide to form potassium sulfide and water. How many grams of the excess is left over? How many grams of the water were produced? How many grams of the potassium sulfide were produced?

8. Determine the percent yield for the reaction between 14.0 grams of nitrogen gas and 13.0 grams of hydrogen gas if 11.5 grams of NH_3 is produced?

XII. Stoichiometry with Molarity

Using the definition of molarity, the given balanced equations, and stoichiometry, solve the following problems. 1. . $Ca(OH)_2(aq) + H_2SO_4(aq) \rightarrow CaSO_4(s) + 2H_2O(l)$

How many L of 0.35 M $Ca(OH)_2(aq)$ are needed in order to have 7.5 mol of $Ca(OH)_2$?

2. . $CaCO_3(aq) + 2HCl(aq) \rightarrow CO_2(g) + CaCl_2(aq) + H_2O(l)$

How many mol of $CaCO_3(s)$ are needed to make 165 mL of 1.6 M HCl(aq) solution?

3. $2NH_4Cl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(aq) + 2NH_3(g) + 2H_2O(l)$

a. How many g of NH_3 will be formed from this reaction when 3.0 L of a 1.75 M solution of $Ca(OH)_2(aq)$ react?

4. $Zn(s) + 2HCl(aq) \rightarrow ZnCl_2(aq) + H_2(g)$

At STP how many L of H_2 are produced when 33.0 mL of 0.28 M HCl(aq) react?

5. $AgNO_{3\,(aq)} + K_3PO_{4\,(aq)} \rightarrow Ag_3PO_4(s) + 3\ KNO_3(aq)$

A 60.7 mL solution of 1.15 M $AgNO_3$ is mixed with a 24.57 g of solid K_3PO_4. When mixed together, a new solid forms. Determine what mass $Ag_3PO_4(s)$ will form. Calculate the percent yield if only 18.045 g $Ag_3PO_4(s)$ was formed in lab.

XIII. Dilutions

1) If I dilute 350 mL of 0.75 M lithium acetate solution to a volume of 850 mL, what will the concentration of this solution be? 2)

How much volume of a 0.10 M HCl solution will be made by diluting 350 mL of 10 M HCl?

3) If 62.5 ml of HCl stock solution is used to make 440.0 ml of a 0.680 M HCl dilution, what is the molarity of the stock solution?

4) How much water would I need to add to 750 mL of a 2.5 M KCl solution to make a 2.0 M solution?

5) If a 13M HBr stock solution is used to make 750.0 ml of a 0.80 M HCl dilution, what volume of the stock solution was used?

6) How much 0.05 M HCl solution can be made by diluting 250 mL of 10 M HCl?

7) If I add 35 mL of water to 120 mL of a 0.20 M NaOH solution, what will the molarity of the diluted solution be?

XIV. Acids and Bases, pH, pOH

1) Using your knowledge of the Brønsted-Lowry theory of acids and bases, write equations for the following acid-base reactions: a) $HNO_3 + OH^{-1} \rightarrow$

b) $CH_3NH_2 + H_2O \rightarrow$

c) $OH^{-1} + HPO_4^{-2} \rightarrow$

2) Write the names for the following acids and bases:

a) FrOH
b) HI
c) H_2CO_3
d) $Ni(OH)_2$
e) H_2CO_3

3) Write the formula for the following acids and bases:

a) chloric acid

b) sulfurous acid

c) nitric acid

d) phosphorous acid

e) acetic acid

4) Name the 7 strong acids

5) Name the strong bases

6) What is the pH and pOH of a 1.4×10^{-3} HCl solution?

7) What is the pH and pOH of a 1.24×10^{-5} KOH solution?

8) What is the pH and pOH of a solution made by adding water to 23 grams of hydrobromic acid until the volume of the solution is 2400 mL?

9) What is the pH and pOH of a solution made by adding water to 38 grams of LiOH solution until the volume of the solution is 1480 mL? 10)

What is the pH of 0.15 M HNO_2 solution Ka for $HNO_2 = 4.93 \times 10^{-10}$

11) What is the pH of a 0.15 M NH_3 solution Kb for $NH_3 = 1.8 \times 10^{-5}$

12) What is the Ka of a 0.20 M HA solution with a pH of 5.21

13) What is the Kb of 0.20M $C_6H_5NH_2$ solution with a pH of 8.25

XV. Galvanic Cells and oxidation numbers

For the following two problems answer the following:

(a) Write the balanced net ionic chemical equation for the overall spontaneous cell reaction that occurs when the circuit is complete. Calculate the standard voltage, *E*, for this cell reaction.

1) $Zn^{2+} + 2e^- \rightarrow Zn$ $E_{red} = -0.76$ v

$Ag^{+} + 1e^- \rightarrow Ag$ $E_{red} = +0.80$ v

2) $Cd^{2+} + 2e^- \rightarrow Cd$ $E_{red} = -0.40$ v

$Sn^{2+} + 2e^- \rightarrow Sn$ $E_{red} = -0.14$ V

3. Assign oxidation numbers to the elements in the following compounds.

a. H_2S b. $H_2C_2O_4$ c. $H_2PO_4^{-1}$ d. IO_3^{-1} e. H_3O^{+1} f. Na_2SO_4

g. $FeCl_3$ h. NH_4^{+1} i. $K_2Cr_2O_7$ j. Ni_2O_3 k. CO_3^{-2} l. N_2H_{5+1}

4. Identify the elements being reduced and oxidized.

a. $2KClO_3 \rightarrow 2KCl + 3\ O_2$
b. $2AgNO_3 + Ni \rightarrow Ni(NO_3)_2 + 2Ag$
c. $H_2 + Cl_2 \rightarrow 2HCl$
d. $2\ HBr + Cl_2 \rightarrow 2HCl + Br_2$
e. $H_2 + O_2 \rightarrow 2H_2 0$
f. $NaI + Br_2 \rightarrow NaBr + I_2$

XVI. Net Ionic Equations

1. Solution of silver (I) acetate is added to a solution of lithium phosphate
2. Solution of barium hydroxide is added to a solution of magnesium bromide.
3. Solutions of potassium fluoride and dilute hydrochloric acid are mixed.
4. Magnesium turnings are added to a solution of iron (III) chloride.
5. Chlorine gas is bubbled into a solution of potassium iodide
6. Aluminum metal is added to a solution of nickel (II) fluoride.