GUIDELINES FOR COMPLETING THE ASSIGNMENT

This packet was created to help you succeed in your upcoming AP Chemistry class. Many of the concepts covered in this packet were taught to you in your previous Chemistry class.

For each of the questions make sure you show all relevant work so you can receive full credit. Only HAND WRITTEN work will be accepted.

The packet is due the first day of school on September 7th. For each day the packet is late, your grade will be deducted one grade level.

This packet will count as a 50-point test grade toward your first marking period grade.

The packet will be graded for completeness and correctness. Your teacher will be looking for supporting work to see that you understand each concept.

You will also be given an assessment on the materials within this packet to check for understanding within the first two weeks of school.

If you have questions, please email me during the summer. I expect the packet to be completed using the AP book and internet resources.

Have a great summer!

Supplies needed for your first day of class and every day after:

- Three Ring Binder, notes pages, Scientific Calculator
- Google Drive and apps on your phone

Resources — AP Book- Zumdahl
Khan Academy and other internet sources.
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Atomic Theory

1. What is an isotope?
   Refer to the isotope of Uranium $^{238}\text{U}$
2. How many protons and neutrons are in the nucleus of this isotope.
3. How many electrons are in a single atom of Uranium
4. What is the mass of this isotope of Uranium.
5. Assume silicon has three major isotopes in nature. The average atomic mass of silicon is 28.09 amu. Fill in the missing information in the table.

<table>
<thead>
<tr>
<th>Isotope</th>
<th>Mass (amu)</th>
<th>Abundance</th>
</tr>
</thead>
<tbody>
<tr>
<td>$^{28}\text{Si}$</td>
<td>27.89</td>
<td></td>
</tr>
<tr>
<td>$^{29}\text{Si}$</td>
<td>29.97</td>
<td>4.70%</td>
</tr>
<tr>
<td>$^{30}\text{Si}$</td>
<td>29.97</td>
<td>3.09%</td>
</tr>
</tbody>
</table>

6. Which color of light has the highest frequency, red or green?
7. Which color of light has the longest wavelength, green or violet?
8. Hydrogen emits light with a wavelength of 410 nm, what is the frequency of this light?
9. What is the electron configuration, orbital notation and noble gas notation for phosphorous? For bromine? How many unpaired electrons does phosphorous have? How many unpaired electrons does bromine have?
10. What is the charge for a phosphorous ion? Why does it make this charge?
11. What is the charge for a bromine ion? Why does it make this charge?
12. Which ion has the larger radius in each set?
   a. Be or O
   b. Cu or Br
   c. F or I
   d. O or As
   e. Kr or K
   f. Li or Ba
13. Which of the following sets are isoelectric? (Has the same electron configuration)
   a. Ne and F$^{-1}$
   b. Ca$^{2+}$ and Se$^{-2}$
   c. N and F$^{-1}$
   d. Ba$^{2+}$ and I$^{-1}$
   e. K$^{+}$ and Ca$^{2+}$
14. Which element is the most electronegative in each set?
   a. F or C
   b. Al or Cl
   c. Po or S
   d. Cs or I
   e. Cl or C
Chemical Bonding

15. What charges do Group 1 and Group 2 metals form? Explain your answer using Periodic Trends and Coulombic Forces of Attraction.


17. Draw the size and relative difference for the following ions. Explain the difference in the size using coulombic forces of attraction.
   a. Na⁺ Mg²⁺ Cl⁻¹

18. Which compound makes a covalent bond?
   a. CO or LiF
   b. ZnS or SO₂
   c. BF₃ or Fe₂O₃

19. Which compound forms the bond that is the most ionic in character? Has the greatest electronegativity difference between the atoms in the bond ionic?
   a. Al-O or Na-O
   b. K-Cl or Zn-Cl

20. Which bond is the most polar? (Apply electronegativity)
   a. C-F or C-O
   b. P-O or P-F
   c. As-F or As-S

21. Write the name for the following compounds
   a. FeBr₃  b. CuI  c. CuI₂  d. Li₂SO₄  e. PbS  f. Sn(CO₃)₂

22. Write the formula for the following compounds
   a. chromium (III) hydroxide
   b. magnesium acetate
   c. chromium (IV) oxide
   d. disulfur dichloride
   e. nickel (II) fluoride
   f. ammonia
   g. aluminum nitride
   h. ammonium sulfate
   i. carbon tetraiodide

23. Name the following compounds. Draw the Lewis Structures for the following compounds. For each compound name the geometry and the bond angle based on the geometry.
   a. CO₂
   b. O₃
   c. NH₄⁺¹
   d. CO₃⁻²
   e. HCN
   f. N₂
Quantitative Analysis and Stoichiometry

24. How many moles of sodium carbonate, Na₂CO₃, are contained in 60.0 grams of the compound?

25. How many grams of NaOH are in 0.890 moles of NaOH?

26. How many ethylene molecules are in 15.5 grams of ethylene (C₂H₄)?

27. How many formula units are in 6.30 grams of NaNO₃?

28. How many moles of hydrogen gas can be produced if 0.57 moles of HCl react with excess zinc according to the following equation?

\[ \text{Zn} + \text{HCl} \rightarrow \text{ZnCl₂} + \text{H₂} \]

29. N₂O₅ reacts with water to produce nitric acid. If 3.25 moles of N₂O₅, how many moles of nitric acid are produced?

\[ \_ \text{N₂O₅} + \_ \text{H₂O} \rightarrow \_ \text{HNO₃} \]

30. How many grams of lithium are needed to produce 42.50 grams of lithium nitride according to the following equation?

\[ \_ \text{Li} + \_ \text{N₂} \rightarrow \_ \text{Li₃N} \]

31. Suppose 35.6 grams of Antimony (III) oxide react with excess carbon according to the following equation.

\[ \_ \text{Sb₂O₃} + \_ \text{C} \rightarrow \_ \text{Sb} + \_ \text{CO} \]

a. What mass of antimony will be produced?
b. What mass of CO will be produced?
c. What mass of C is consumed?

32. An 60.0 grams sample of phosphorous reacts with 85 grams of oxygen gas to produce diphosphorous pentoxide according to the following equation.

\[ \_ \text{P} + \_ \text{O₂} \rightarrow \_ \text{P₂O₅} \]

a. Identify the limiting reactant
b. Based on the limiting reactant how many grams of P₂O₅ are produced?
c. If 123.0 grams of P₂O₅ are produced, what is the percent yield?

33. A 8.75 gram sample of silicon dioxide react with 5.50 grams of sodium hydroxide according to the following equation.

\[ \_ \text{SiO₂} + \_ \text{NaOH} \rightarrow \_ \text{Na₂SiO₃} + \_ \text{H₂O} \]

a. What is the limiting reactant if? Justify your answer.
b. How many grams of Na₂SiO₃ can be produced (Theoretical yield)?
c. If 7.24 grams of Na₂SiO₃ are produced (Actual yield), what is the percent yield?
34. A manufacturer produces a vitamin C tablet with a mass of 0.825 grams. The mass of vitamin C in the tablet is only 70 mcg (micrograms) of vitamin C. What is the percent of vitamin C in the tablet? (percent composition by mass).

35. Calculate the percent composition by mass of each element in Na₂SO₄.

36. What is the empirical formula for C₄H₁₀?

37. Is CO₂ an empirical formula? A molecular formula or both? Explain.

38. A compound contains 68.1% carbon, 13.7% hydrogen and, 18.2% oxygen by mass. (Refer to Zumdahl CH 3)
   a. What is the empirical formula for the compound?
   b. If the compound has a molar mass of 176.34 g/mol, what is the molecular formula?

39. A compound was analyzed and found to contain 6.00 grams of carbon and 1.10 g of hydrogen. (refer to Zumdahl CH 3)
   a. Find the empirical formula
   b. If the compound has a molar mass of 142.36 g/mol, what is the molecular formula?

Solutions

40. Each of the following compounds are dissolved in water. Identify the type of compound first – molecular (polar, non-polar, acid) Ionic (base)
   a. Is the compound a strong electrolyte, weak electrolyte, polar molecule or a non-polar molecule?
   b. Which compounds will not dissolve in water? Why
   NaCl, MgNO₃, KOH, HCl, CH₃COOH, HF, C₄H₁₂O₆, CH₃CH₂OH, CCl₄, CH₂Cl₂

41. What piece of equipment do you use to make solutions in a laboratory?

42. What safety precautions do you need to use when working with strong acids?

43. What is the molarity of 0.080 grams of NaOH dissolved in 100 mL solution?

44. How many grams are needed to make a 250 mL solution that is 0.25 M CuCl₂? Describe how you would prepare this solution

45. How do you prepare a dilution of 0.300 M HCl with a volume of 500 mL from a stock that is 12.1 M?

46. How do you prepare 1.00 L of a 0.15 M HNO₃ from a 13.8M stock reagent solution?

47. Which of the following solutions would create a precipitate? Write a molecular equations and a net ionic equation for the reaction that forms a precipitate.
   a. Al(NO₃)₃ + NaCl
   b. K₂CO₃ + MgI₂
   c. CaCl₂ + NaNO₃
48. In a reaction 50.0 mL of 0.15 M CaCl₂ is reacted with 150. mL of 0.25 M of Na₂CO₃.
   a. Predict the products and write a balanced equation.
   b. Write the net ionic equation.
   c. Identify the limiting reactant.
   d. How much solid product is produced?
   e. What is the molarity of NaCl in the final solution?

49. A 25.00 mL of HClO₄ requires 32.0 mL of 0.125 M NaOH for complete neutralization. What is the original concentration of the HClO₄ solution?

50. What volume of 0.25 M KOH will completely react with 50.0 mL of 0.15 M HNO₃?
   a. Write a balanced reaction for the equation.
   b. Why is the called a neutralization reaction?
   c. What is the limiting reactant?
   d. Will the solution remaining after the reaction has occurred be a neutral, acidic or basic pH? How do you know? Explain your answer.
Learn your IONS

“From the Periodic Table “

The ions can be organized into two groups.

1. Their place on the periodic table suggests the charge on the ion, since the neutral atom gains or loses a predictable number of electrons in order to obtain a noble gas electrons configuration.
   a. All Group 1 Metals (alkali metals) lose one electron to form an ion with a +1 charge.
   b. All Group 2 Metals (alkali earth metals) lose two electrons to form an ion with a +2 charge.
   c. All Group 3 Metals like Aluminum lose three electrons to form an ion with a +3 charge.
   d. Al Group 17 Non-metals (halogens) gain one electron to form a -1 charge.
   e. All Group 16 Non-metals gain two electrons to form an ion with a -2 charge.
   f. All Group 15 Non-metals gain three electrons to form an ion with a -3 charge.

Cations keep the name and anions end in “ide”

Polyatomic Ions

Most of the work on learning polyatomic ions comes from a number of patterns in their names.

1. ‘ate’ anions have one more oxygen than the “ite” ion, but the same charge. Learn the ate ion and remember the trend.
   a. Example” SO$_4$^{2-} is named sulfate and SO$_3$^{2-} is the sulfite ion.
   b. Example NO$_3$^{-1} is name nitrate and No2 – is named nitrite.
2. If you had hydrogen to a polyatomic ion, since hydrogen has a +1 charge, the charge of the polyatomic ion decreases by one. Add hydrogen to the ions name.
   a. SO$_4$^{2-} is sulfate HSO$_4$^{-1} is named hydrogen sulfate.
   b. A common household product – baking soda – is NaHCO$_3$ and named sodium hydrogen carbonate.
3. Learn the hypochlorite series (you can apply this to any halogen):
   hypochlorite $\rightarrow$ chlorite $\rightarrow$ chlorate $\rightarrow$ perchlorate
   HClO $\rightarrow$ HClO$_2$ $\rightarrow$ HClO$_3$ $\rightarrow$ HClO$_4$

This series holds true, especially for the halogens,
   hypobromite $\rightarrow$ bromite $\rightarrow$ bromate $\rightarrow$ perbromate
   HBrO $\rightarrow$ HBrO$_2$ $\rightarrow$ HBrO$_3$ $\rightarrow$ HBrO$_4$
## Mastery of Common Ions and Charges

A mastery of common ions, their formulas and charges is essential for AP chemistry. You are expected to know these on the first day of class – and will be tested on them. You should learn them based on the periodic table trends taught in Honors Chemistry. Many of them you already know!

<table>
<thead>
<tr>
<th>Cations</th>
<th>Name</th>
<th>3rd Energy Level Cations</th>
<th>Name</th>
</tr>
</thead>
<tbody>
<tr>
<td>H⁺</td>
<td>Hydrogen</td>
<td>Ag⁺</td>
<td>Silver</td>
</tr>
<tr>
<td>Li⁺</td>
<td>Lithium</td>
<td>Zn²⁺</td>
<td>Zinc</td>
</tr>
<tr>
<td>Na⁺</td>
<td>Sodium</td>
<td>Hg₂⁺²</td>
<td>Mercury (I)</td>
</tr>
<tr>
<td>K⁺</td>
<td>Potassium</td>
<td>NH₄⁺</td>
<td>Ammonium</td>
</tr>
<tr>
<td>Rb⁺</td>
<td>Rubidium</td>
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</tr>
<tr>
<td>Cs⁺</td>
<td>Cesium</td>
<td>Polyatomic Ions</td>
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<tr>
<td></td>
<td></td>
<td>Anions</td>
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</tr>
<tr>
<td>Be⁺²</td>
<td>Beryllium</td>
<td>NO²⁻</td>
<td>Nitrite</td>
</tr>
<tr>
<td>Mg⁺²</td>
<td>Magnesium</td>
<td>NO³⁻</td>
<td>Nitrate</td>
</tr>
<tr>
<td>Ca⁺²</td>
<td>Calcium</td>
<td>SO₃⁻²</td>
<td>Sulfite</td>
</tr>
<tr>
<td>Ba⁺²</td>
<td>Barium</td>
<td>SO₄⁻²</td>
<td>Sulfate</td>
</tr>
<tr>
<td>Sr⁺²</td>
<td>Strontium</td>
<td>HSO₄⁻</td>
<td>Hydrogen sulfate</td>
</tr>
<tr>
<td>Al⁺³</td>
<td>Aluminum</td>
<td>OH⁻</td>
<td>Hydroxide</td>
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<tr>
<td>Anions</td>
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<td>PO₄³⁻</td>
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<td>Fluoride</td>
<td>HPO₄⁻²</td>
<td>Hydrogen phosphate</td>
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<td>Cl⁻</td>
<td>Chloride</td>
<td>H₂PO₄⁻</td>
<td>Dihydrogen phosphate</td>
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<tr>
<td>Br⁻</td>
<td>Bromide</td>
<td>SCN⁻</td>
<td>Thiocyanate</td>
</tr>
<tr>
<td>I⁻</td>
<td>Iodide</td>
<td>CO₂⁻²</td>
<td>Carbonate</td>
</tr>
<tr>
<td>O²⁻</td>
<td>Oxide</td>
<td>HCO₃⁻</td>
<td>Hydrogen carbonate</td>
</tr>
<tr>
<td>Se⁻²</td>
<td>Selenide</td>
<td>ClO⁻ Or BrO⁻ Or IO⁻</td>
<td>Hypochlorite, hypobromite, hypoiodite</td>
</tr>
<tr>
<td>N³⁻</td>
<td>Nitride</td>
<td>ClO₂⁻ Or BrO₂⁻ Or IO₂⁻</td>
<td>Chlorite, bromite, iodite</td>
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<td>P³⁻</td>
<td>Phosphide</td>
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<td>As³⁻</td>
<td>Arsenide</td>
<td>ClO₄⁻ Or BrO₄⁻ Or IO₄⁻</td>
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<tr>
<td>3rd Energy Level Or More Cations</td>
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<td></td>
</tr>
<tr>
<td>Fe³⁺</td>
<td>Iron (III)</td>
<td>MnO₄⁻</td>
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<td>CrO₄⁻²</td>
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<td>Amide</td>
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<tr>
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<td>Lead (II)</td>
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</tr>
<tr>
<td>Hg⁺²</td>
<td>Mercury (II)</td>
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</table>