

Chemistry 116: Laboratory Homework
Experiment 8: Group I Qualitative Analysis

Student Name: _____ Lab Room No. _____ Desk No. _____
Lab Day _____ Lab Start Time _____ Date _____

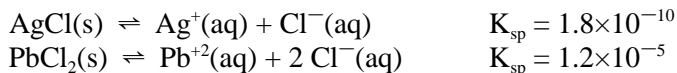
Directions: Answer the following questions. For numerical answers a) show all work to support your final numerical answer(s), b) include correct units and significant figures, and c) circle your final numerical answer. Answer non-numerical questions by a) answering the question, b) fully explaining, and c) using complete sentences.

1. During the qualitative analysis scheme, preliminary separation of cations into five different groups via addition of four different group reagents is based upon what factor?
2. An aqueous solution of a salt (one cation and one anion) is thought to contain the metal cation Ba^{+2} . To test this hypothesis, a student added $(\text{NH}_4)_2\text{CO}_3$ in $\text{NH}_3/\text{NH}_4\text{Cl}$ buffer to 1 mL of the salt solution and obtained a precipitate. The student then concluded that the aqueous solution contained a Group IV metal cation.
 - a. What is wrong with this student's reasoning?
 - b. To what groups could the cation belong?
 - c. What is the proper procedure for testing for the presence of a Group IV metal cation?
3. An aqueous solution was tested for the presence of cations belonging to Groups I-V as outlined in Experiment 8. The first group reagent was added to 1 mL of the aqueous solution and any precipitate was separated from the supernatant liquid. The second group reagent was added to the supernatant liquid and any precipitate was separated, etc. In this way, the four group reagents were sequentially added to the supernatant liquid obtained from the previous separation step. The experimental results are listed below. Cations from which groups are or may be present in this solution? Explain.

	<u>Reagent</u>	<u>Result</u>
1:	dilute HCl	no precipitate
2:	H_2S in 0.3 M HCl	precipitate
3:	H_2S in $\text{NH}_3/\text{NH}_4\text{Cl}$	precipitate
4:	$(\text{NH}_4)_2\text{CO}_3$ in $\text{NH}_3/\text{NH}_4\text{Cl}$	no precipitate

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4. To prove the statement...*PbCl₂ is more soluble than AgCl*...found on page 91 of Experiment 8, calculate the solubilities (**in g/100 mL**) of AgCl and PbCl₂ in pure water. The equilibria are:



5. A solution is 0.00030 M in each of the cations, Ag⁺, Hg₂⁺², and Pb⁺². The K_{sp} values for AgCl, Hg₂Cl₂, and PbCl₂ are 1.8×10⁻¹⁰, 1.4×10⁻¹⁸, and 1.2×10⁻⁵, respectively.
- Calculate the concentration of Cl⁻ needed to begin the precipitation of each cation.
 - As dilute HCl is added and the concentration of Cl⁻ is steadily increased, which chloride salt precipitates first? Which last?
 - What concentration of Cl⁻ is needed to precipitate as much Hg₂⁺² as possible without precipitating the Ag⁺ or Pb⁺²?
 - What concentration of Cl⁻ is needed to precipitate as much Ag⁺ and Hg₂⁺² as possible without precipitating the Pb⁺²?
 - Enough HCl is added so that the equilibrium concentration of Cl⁻ is 0.20 M. Calculate the concentration of Ag⁺ and Pb²⁺ which remain in solution when [Cl⁻]_{eq} = 0.20 M.
 - At this concentration of Cl⁻, which one of the cations (Ag⁺ or Pb⁺²) still remains in solution to a significant extent? Remember, the original solution was 0.00030 M in each cation. Explain.