Name: ___________________________

Section: ________

Lab Worksheet for "**Chemical Equilibrium and Le Chatelier's Principle**"

General Instructions:

- Complete Part A, Part B Steps 1a-1e (skip 1f) and Steps 2a-2e (skip 2f-2i). Follow the procedure in the lab manual and record your data on this worksheet.
- As your laboratory report, turn in to your TA this worksheet along with the appropriate pages from your laboratory notebook (introduction/ procedure and your answers to the worksheet questions).

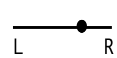
PART A: Colors of Coordination Compounds (Complex Ions)

Step	aqueous complex ion	coordination compound (s)	ionic precipitate	Observations
1a		$CoCl_2 \cdot 6H_2O$		
	$[Co(H_2O)_6]^{2+}$			
1b	$[CoCl_4]^{2-}$			
2a		$CuSO_4 \cdot 5H_2O$		
	$[Cu(H_2O)_6]^{2+}$			
2b	$[Cu(NH_3)_4(H_2O)_2]^{2+}$			
2c			$Cu(OH)_2$	
2d	$[CuBr_4]^{2-}$			

PART B: Shifting Chemical Equilibria

Record the following information in the tables for the cobalt and copper equilibrium systems:

- Your observation about the color of the solution at each step of the procedure.
- Your estimate, based on color, of the relative amounts of the left-hand species ("L") and the right hand species ("R") in the solution, at each step of the procedure. For example, for the cobalt system, "L" is $[Co(H_2O)_6]^{2+}$ and "R" is $[CoCl_4]^2$. As an example of the recording system, if L is yellow and R is blue, and you observe a bluish-green solution, you might estimate — based on the logic that a green solution would contain rougly equal amounts of L and R — that there is a little more of the blue R-species in solution, and you would mark this "some L, but more R" mixture on the scale like this:



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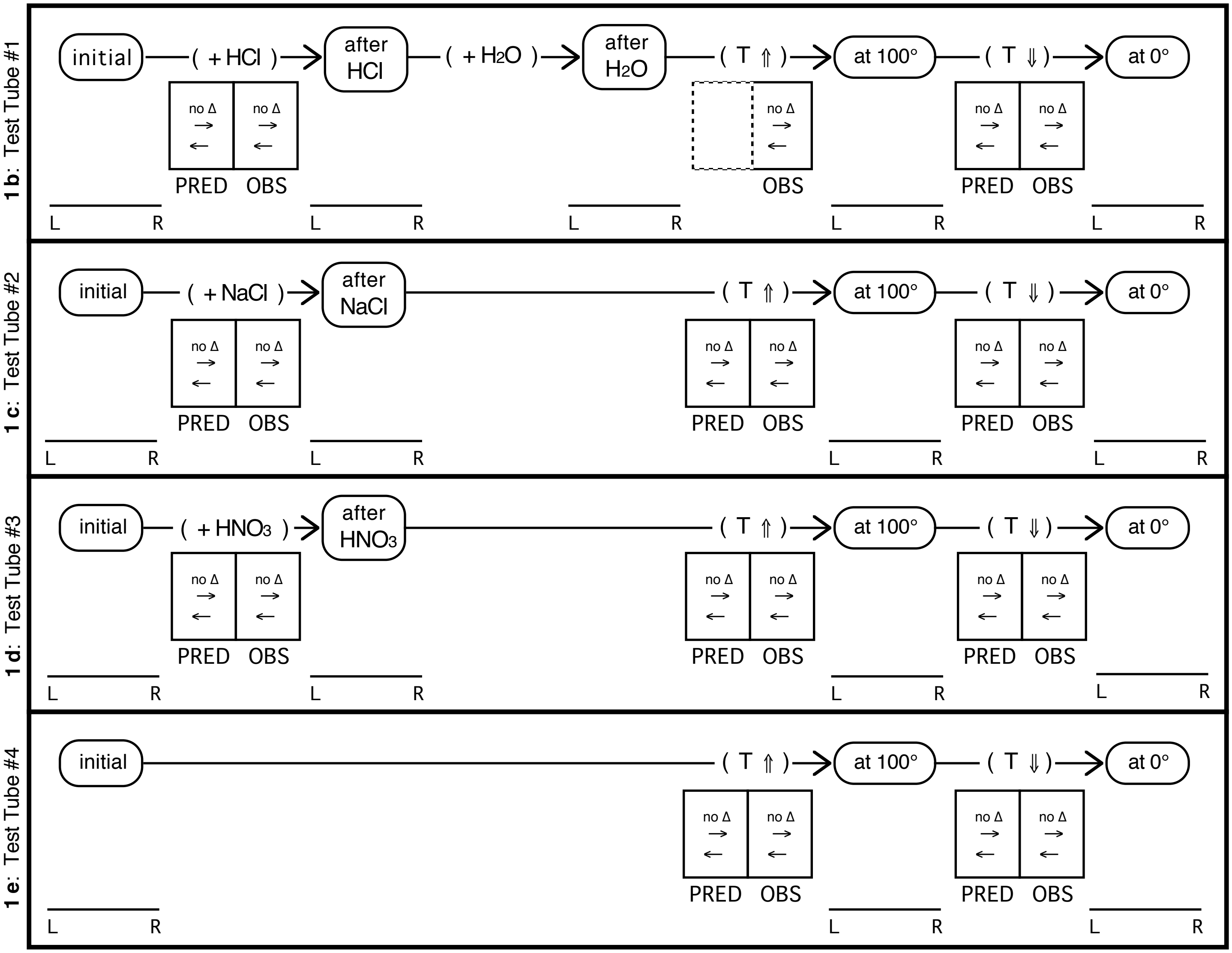
- Your prediction, based on Le Chatelier's Principle, about changes in the equilibrium mixture after adding a "stress" (such as adding a chemical, heating etc.) — that there is no change, a shift to the right, or a shift to the left — is indicated by circling the appropriate item in the "PRED" column.
- Your conclusion based on your observation of the colors before and after the "stress" of whether the equilibrium shifted to the left, right, or not at all - circle the appropriate item in the "OBS" column.

Steps 1a - 1e : Equilibrium Systems Involving Cobalt(II) Complexes

Write your observations about the color of each species in the cobalt(II) equilibrium system:

COLOR OBS:

REACTION: $[Co(H_2O)_6]^{2+}$ + 4 Cl $\rightleftarrows$ $[CoCl_4]^{2-}$ + 6 H_2O

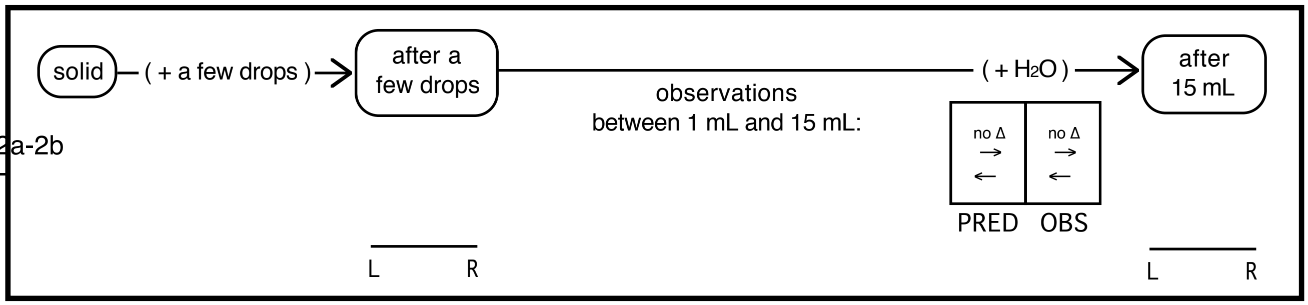


Steps 2a - 2e: Equilibrium Systems Involving Copper(II) Complexes

Write the reaction-equation for the appropriate copper(II) equilbrium system in Steps 2a-2b, and your observations about the color of each species:

COLOR OBS:

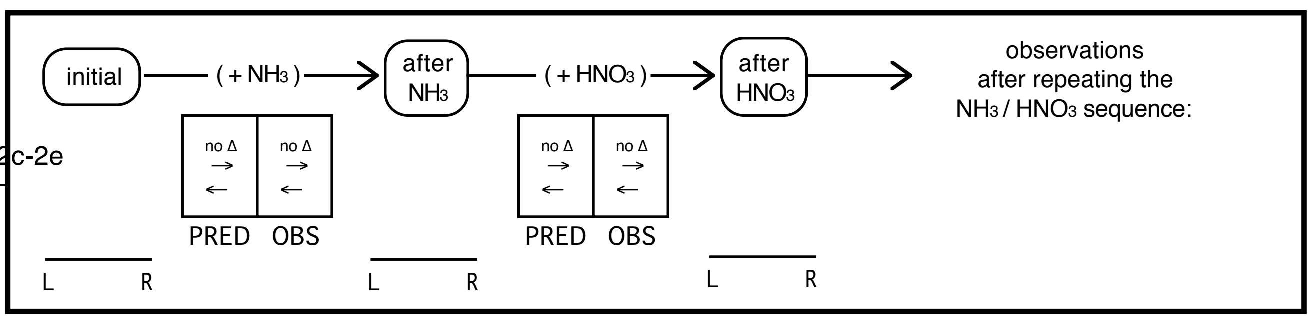
REACTION: + $\rightleftarrows$ +



Write the reaction-equation for the appropriate copper(II) equilbrium system in Steps 2c-2e, and your observations about the color of each species:

COLOR OBS:

REACTION: + $\rightleftarrows$ +



QUESTIONS:

1) Was there always agreement between your theory-based predictions and experimental observations? If not, when did disagreements occur? Can you construct hypotheses to explain each discrepancy?
2) Based on your experimental observations, is the forward reaction of the $[Co(H_2O)_6]^{2+}$ / $[CoCl_4]^{2-}$ equilibrium system endothermic or exothermic? What evidence supports your conclusion?
3) If a student observes a "violet" colored solution and interprets the color as "an equal mix of pink and blue," what can he/she conclude about the concentration ratio of $[Co(H_2O)_6]^{2+}$ and $[CoCl_4]^{2-}$ in the solution? { Hint: Read page 17, "You should be aware ... [re: the intensity] of their two colors." }
4) Propose two theories that would explain why you observed the "in-between colors" in Step 2b. Which of these alternative theories do you think is more likely, and why?
5) In Step 2d of Part B, explain why you observed an equilibrium shift even though neither ionic component of HNO_3 (H_3O^+ and NO_3^-) appears in the chemical equation for the equilibrium system.

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Answers for Questions, from Grading Criteria for Spring 1996, written by Jacquie Scott:

1. Addition of Cl^- ions from NaCl should push the equilibrium to the right, ie. towards the production of the blue $[CoCl_4]^{2-}$ species. This is counter to the student's observation. The most likely cause is that NaCl is not very soluble in water and so the concentration of Cl^- is too low to push the equilibrium far enough to the right to be observed.
2. Students should observe in test tube #1 (Part B - 1b) that the $CoCl_2$/ HCl/ H_2O solution becomes more blue as it is heated and more pink as it is cooled - thus the equibrium is endothermic in the forward direction (as written).
3. The blue color of the $[CoCl_4]^{2-}$ is 100x as intense than the pink color of the $[Co(H_2O)_6]^{2+}$ species. Thus if the color appears roughly an equal mixture of blue and pink there is actually much more $[Co(H_2O)_6]^{2+}$ in the solution (approximately 100 times less $[CoCl_4]^{2-}$ than $[Co(H_2O)_6]^{2+}$).
4. The green color may arise from an intermediate species between $[CuBr_4]^{2-}$ and $[Cu(H_2O)_6]^{2+}$, for example $[CuBr_3(H_2O)_6]^{1-}$, or it may simply be the combination of the brown of the $[CuBr_4]^{2-}$ and blue of the $[Cu(H_2O)_6]^{2+}$.
5. The HNO_3(aq) reacts with NH_3(aq) via an acid-base reaction. This reduces the concentration of NH_3(aq) in the system causing the equilibrium to shift towards the left, and as a consequence, creates more $[Cu(H_2O)_6]^{2+}$

Overview of Experimental Results, from Staff Notes for Spring 1995, written by Jacquie Scott:

Part A: Colors of Complex Ions/Coordination Compounds

1. Complex Ions of Cobalt

   $[Co(H_2O)_6]^{2+}$ — pink solution
   "few crystals $CoCl_2$ + 10 mL water"

   $[CoCl_4]^{2-}$ — medium blue solution
   "few crystals $CoCl_2$ + 3 mL 12 M HCl"

2. Complex Ions and Coordination Complexes of Copper

   $[Cu(H_2O)_6]^{2+}$ — light blue or aqua solution
   "few crystals $CuSO_4$ + **3** mL water"

   $[Cu(NH_3)_4(H_2O)_2]^{2+}$ — dark (navy) blue solution
   Add 6M NH_4OH to water/$CuSO_4$ solution*

   $Cu(OH)_2$ — aqua-blue solid
   "few crystals $CuSO_4$+ 10 mL water
   + 1 mL 6M NaOH"

   $[CuBr_4]^{2-}$ — brown solution (note: this species may also sometimes appear green)
   "few crystals $CuSO_4$ + 2-3 mL KBr sat'd water"

*$Cu(OH)_2$ may initially precipitate out of solution. However, it will dissolve with the continued addition of 6 M NH_4OH. Students may need to add more of the 6 M NH_4OH than is actually called for in the manual (1-2 drops). The amount of base needed will vary with the amount of $CuSO_4$ used to make up the original solution.

Part B: Factors Affecting Chemical Equilibrium

1. Complex Ions of Cobalt

a. Color of initial 0.1M $CoCl_2(aq)$ solution: pink

	Observations	Predominant Species
b. Test Tube 1		
Effect of HCl	blue-violet solution	$[CoCl_4]^{2-}$
Effect of H_2O	light blue/ violet solution	$[CoCl_4]^{2-}$
At 100 °C	deep blue solution	$[CoCl_4]^{2-}$
At 0 °C	light violet solution	$[CoCl_4]^{2-}$
c. Test Tube 2		
Effect of NaCl	solution remains pink	$[Co(H_2O)_6]^{2+}$
At 100 °C	violet solution	$[CoCl_4]^{2-}$
At 0 °C	pink solution	$[Co(H_2O)_6]^{2+}$
d. Test Tube 3		
Effect of HNO_3	solution remains pink	$[Co(H_2O)_6]^{2+}$
At 100 °C	solution remains pink	$[Co(H_2O)_6]^{2+}$
At 0 °C	solution remains pink	$[Co(H_2O)_6]^{2+}$
e. Test Tube 4		
At room temp.	pink solution	$[Co(H_2O)_6]^{2+}$
At 100 °C	solution remains pink	$[Co(H_2O)_6]^{2+}$
At 0 °C	solution remains pink	$[Co(H_2O)_6]^{2+}$

Students should skip Part B - step 1f.

2. Equilibrium Systems with Complexes of Copper(II)

	Observations	Predominant Species
a. & b $CuBr_2$	Color of $CuBr_2$ solid	metallic gray solid
$CuBr_2$ in few drops H_2O	dark brown solution	$[CuBr_4]^{2-}$
$CuBr_2$ in 2-3 mL H_2O	dark green solution	? maybe $[CuBr_3(H_2O)]^{1-}$
$CuBr_2$ in 15 mL H_2O	aqua blue solution	$[Cu(H_2O)_6]^{2+}$
c. Addition of $NH_3(aq)$	aqua-blue ppt. → dark (navy) blue solution	$Cu(OH)_2(s)$ → $[Cu(NH_3)_4(H_2O)_2]^{2+}$
d. Addition of HNO_3	aqua blue solution	$[Cu(H_2O)_6]^{2+}$

e. Reversibility sequence Yes, the reaction is reversible; however, very little (if any) precipitate was formed the second and third times.

Students should skip Part B - Steps 2f to 2i.

SUMMARY OF RESULTS

1. $[Co(H_2O)_6]^{2+}(aq) + 4\ Cl^{-}(aq) \rightarrow [CoCl_4]^{2-}(aq) + 6\ H_2O(l)$
   pink blue

   Addition of Reagent
   a. HCl - equilibrium shifts to the right
   b. NaCl - equilibrium lies to the left
   c. HNO_3 - equilibrium lies to the left

   Addition of Heat
   a. $CoCl_2$ solution - equilibrium lies to the left
   b. HCl + $CoCl_2$ solution - equilibrium lies to the right
   c. NaCl + $CoCl_2$ solution - equilibrium shifts to the right
   d. HNO_3 + $CoCl_2$ solution - equilibrium lies to the left

   Cooling the Solution after Heating to 100°C
   a. $CoCl_2$ solution - equilibrium lies to the left
   b. HCl + $CoCl_2$ solution - equilibrium shifts slightly to the left
   c. NaCl + $CoCl_2$ solution - equilibrium shifts to the left
   d. HNO_3 + $CoCl_2$ solution - equilibrium lies to the left

2. $[Cu(H_2O)_6]^{2+} + 4\ NH_3 \rightarrow [Cu(NH_3)_4(H_2O)_2]^{2+} + 4\ H_2O$
   aqua dark blue

   a. Addition of NH_3 - equilibrium shifts to the right
   b. Addition of HNO_3 - equilibrium shifts to the left
   (Sequences a & b are reversible.)

3. $[Cu(H_2O)_6]^{2+}(aq) + 4\ Br^{-}(aq) \rightarrow [CuBr_4]^{2-}(aq) + 6H_2O(l)$
   lt. blue dark brown

   a. Solution with $[CuBr_4]^{2-}$ as predominant species shifts to the left as H_2O is added.