

## CHEMISTRY 206

### Experiment 5: EXPERIMENTS IN CHEMICAL EQUILIBRIUM

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#### Instructor's Informal Preamble

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In these experiments, you will investigate several systems where reactions do not proceed to completion. In reality, this is true of all chemical processes. However, in many cases, the reaction proceeds so nearly to completion that for stoichiometric calculations it can safely be assumed that the limiting reactant is entirely consumed. In reactions that seem to stop *before* all the limiting reactant is consumed, what happens is that the products begin to react as they are formed, to give back the original reactants. Remember that reaction rates depend on concentration: as more product forms, the backward reaction becomes faster, and at the same time, the forward reaction slows down since reactant has been consumed. Eventually, an *equilibrium* situation is reached where the forward and backward reaction rates are equal. Equilibrium is reached quickly for reactions with large rate constants (with small activation energies,  $E_a$ , or at high temperature,  $T$ ), but slowly for reactions with small rate constants (with large  $E_a$ , or at low  $T$ ).

When a chemical system reaches equilibrium, the system will appear from the outside to be unchanging, because if substances are being consumed but also produced at the same rate, then their concentrations will not actually change. Students often misinterpret this phenomenon as the reaction “stopping”, but in fact, at the molecular level the situation is still *dynamic*, with molecules colliding and reacting non-stop. However, the unchanging nature of the *concentrations* at equilibrium is quite useful, because it allows us to determine whether or not a reaction has reached equilibrium. If we see that the concentrations in a chemical system are unchanging, we can make two hypotheses:

- 1) the activation energy for the reaction is so large that the reaction cannot proceed in one or both directions under these conditions (*i.e.*, temperature is too low...);
- 2) the system is at equilibrium (*this one is more likely!*).

To test these hypotheses, we can add more of one of the substances and see if this causes the reaction to “start up” again, which can only happen if hypothesis #1 is *not* correct. If the system *is* at equilibrium, then if we add more reactant, the forward reaction will suddenly be faster than the reverse reaction. Therefore, we will notice a net forward reaction for a while, until the forward and reverse reaction rates balance out again and the reaction reaches a new equilibrium situation and appears to “stop”. We can predict this kind of outcome using Le Châtelier's Principle, the handy rule that states that if a system at equilibrium is disturbed (by adding or removing a substance, or changing the pressure or temperature), it will respond in a way that counteracts the disturbance. To understand why this rule works, you must think about how each type of disturbance affects the rate of both the forward and reverse reactions, like we just did.

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#### Introduction

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If a system is at equilibrium, the concentrations of the reactants and products will be unchanging, and they will be related to each other by a constant called the *equilibrium constant*,  $K_{\text{eq}}$ . For a reaction:



we express  $K_{\text{eq}}$  as:

$$K_{\text{eq}} = [C]^c[D]^d / [A]^a[B]^b$$

The equilibrium constant for a particular reaction has a constant value under a particular set of conditions. The value of  $K_{\text{eq}}$  for a reaction will vary, however, if the temperature is changed, because the rate constants ( $k$ ) for the forward and reverse reactions will each change, and not by the same amount (recall Arrhenius' equation:  $k = Ae^{-E_a/RT}$ ). Therefore, when studying reactions at equilibrium and/or their response to disturbances, it is necessary to maintain a constant (and known) temperature; we will perform all of our reactions at room temperature.

It is important to remember that even though the equilibrium constant for a particular reaction has a constant value, this does not mean that every reaction mixture looks the same when it has reached equilibrium. Countless different combinations of reactant and product concentrations can lead to a balance of forward and reverse rates and therefore yield the same value of  $K_{\text{eq}}$ . When a system reaches equilibrium, we describe that particular situation as one of system's *equilibrium positions*. If this reaction mixture is disturbed and then re-equilibrates, its new equilibrium position will be either more enriched in reactants (if the equilibrium "shifted left") or more enriched in products (if the equilibrium "shifted right").

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#### Experiment Summary

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*Experiment 5 involves two sets of equilibrium-related experiments. The following instructions cover a two-week period. Please work in pairs.*

- **PART 1:** Determination of the effect of various influences on the position of equilibrium. *These experiments will be performed in the first week.*
- **PART 2:** Determination of an equilibrium constant. *In the first week you will only mix the solutions. The actual measurements will be made in the second week.*

**PART 1: Determination of the Effect of Various Influences on the Position of Equilibrium**

In these experiments you will be studying three different kinds of chemical system that come to equilibrium very quickly. You will be investigating what happens when the equilibrium is disturbed after it has been established. Recall the rule called *Le Châtelier's Principle*, which states that, if a system already at equilibrium is disturbed, it will react in a way that will minimize the effect of the disturbance. The disturbance can take the form of a change in the overall temperature or pressure or concentration of one of the reagents or products. The last of these effects is easiest to demonstrate and is used in the following three experiments:

***a) Solution Equilibria - Ionic Compounds***

In a saturated solution in contact with excess solid solute, such as sodium chloride, the equilibrium can be written:



and the equilibrium constant is written:

$$K_{\text{eq}} = [\text{Na}^+(\text{aq})][\text{Cl}^-(\text{aq})] / [\text{NaCl(s)}]$$

The solid settles out of the solution, and the concentration of the solid can be treated as constant. We typically do not include it in the expression; instead, its value becomes part of a new equilibrium constant:

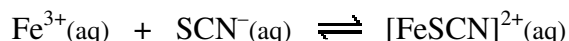
$$K_{\text{sp}} = [\text{Na}^+(\text{aq})][\text{Cl}^-(\text{aq})]$$

where  $K_{\text{sp}}$  is a specific type of equilibrium constant called the *solubility product*.

If we take a solution that has been saturated with sodium chloride and then add an additional source of one of the ions (such as aqueous HCl), the solution will initially be supersaturated. The reaction quotient,  $Q$ , will be greater than the equilibrium constant,  $K_{\text{sp}}$ , so the reverse process (*precipitation*) will initially be faster than the forward process (*dissolution*). As the system gradually restores equilibrium, you will notice that some of the ions will be precipitated from the solution, as we would predict using Le Châtelier's Principle.

***b) Complex Ion Equilibria***

The bright colours of many compounds of the transition metals are associated with compounds called complexes, in which anions or small molecules form *coordinate* bonds with a metal ion. One example is the reaction of aqueous ferric ( $\text{Fe}^{3+}$ ) ions with thiocyanate anion ( $\text{SCN}^-$ ) to produce a blood-red complex:



By noting the intensity of the solution's colour, you will be examine the effect of changing the  $\text{Fe}^{3+}$  concentration, or of decreasing the concentration of  $\text{SCN}^-$  by adding  $\text{Ag}^+$ , which forms a very stable colourless (white) complex with thiocyanate.

**c) Acid/Base Equilibria**

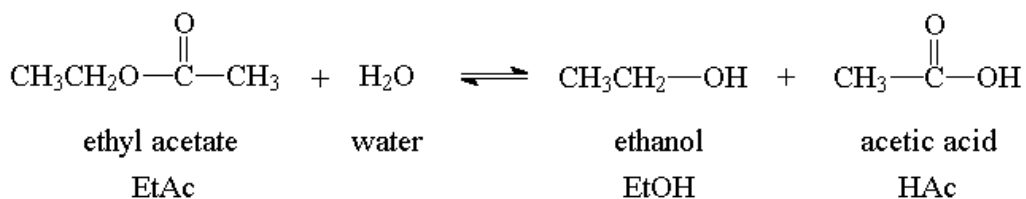
When weak acids or bases are dissolved in water, there is an equilibrium established between the ionized form of the compound and the non-ionized form. An example is a solution of ammonia in water. The equilibrium can be written in several ways, one of which is:



Adding ammonium ion or changing the pH will cause the equilibrium to shift, but the result is not visible because all the species are colourless. The addition of a pH indicator such as phenolphthalein renders the effects visible: phenolphthalein is colourless in acidic solution but bright pink in basic solution. Remember that a solution's pH is a measure of its  $\text{H}_3\text{O}^+$  concentration (which varies inversely with  $\text{OH}^-$  concentration), which means that the indicator's colour allows you to determine whether your equilibrium has shifted towards more or less  $\text{OH}^-$ .

**PART 2: Determination of an Equilibrium Constant**

In this part, you will be investigating the *hydrolysis* (cleavage by water) of ethyl acetate:



The equilibrium expression for this reaction is therefore:

$$K_{\text{eq}} = [\text{EtOH}][\text{HAc}] / [\text{EtAc}][\text{H}_2\text{O}]$$

$$= (\text{moles of EtOH/vol})(\text{moles of HAc/vol}) / (\text{moles of EtAc/vol})(\text{moles of H}_2\text{O/vol})$$

We will perform this reaction in mixtures containing relatively similar quantities of water and ethyl acetate, so it would not be accurate to assume that the concentration of water in the system remains constant during the reaction. For this reason, water's concentration must be included in the equilibrium expression. *Only substances whose concentrations do not change significantly during a reaction, such as solids and pure liquids (or "nearly pure" liquids, like the solvent in a dilute solution), are left out of equilibrium expressions, because they can be treated as constant and incorporated into the value of  $K_{\text{eq}}$ .*

The hydrolysis of ethyl acetate is normally a very very slow reaction, but it is catalyzed significantly by hydrogen ions; in acidic solution, the forward and reverse reactions' activation energies are small enough that equilibrium can be established in about a day. Because the hydrogen ions act as a catalyst and are therefore not consumed overall, their concentration does not appear in the equation for the equilibrium constant. In this experiment, the hydrogen ions are added in the form of hydrochloric acid.

Since you will be making up the reaction mixtures, you will know the *initial* concentrations of the reactants and acid catalyst. The concentration of acetic acid *at equilibrium* will be determined by performing a titration with sodium hydroxide, which will neutralize the both the HCl catalyst and the acetic acid. (*The hydrolysis of ethyl acetate is slow enough, even when acid catalyzed, that the position of the equilibrium will not be significantly affected by the titration.*) This result will allow you to calculate the concentrations of all the substances at equilibrium and therefore also the value of the equilibrium constant for the reaction.

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### Prelaboratory Assignment

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Read the Introduction and Procedure sections carefully and answer the prelaboratory questions. Your demonstrator will inspect and collect in your prelab before you are permitted to begin the experiment – remember to keep a copy of it for yourself, and have the TA sign your receipt record.

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### Materials

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#### **PART 1: Determination of the Effect of Various Influences on the Position of Equilibrium**

- 9 - test tubes
- 2 - 100 mL Erlenmeyer flasks
- 1 - 100 mL graduated cylinder
- Saturated sodium chloride solution
- Concentrated (12 M) hydrochloric acid – *use carefully!*
- 1 M hydrochloric acid
- 0.1 M iron(III) chloride, FeCl<sub>3</sub>
- 0.1 M potassium thiocyanate, KSCN
- Concentrated ammonia solution
- Ammonium chloride (crystalline)
- 0.1 M silver nitrate, AgNO<sub>3</sub>
- Phenolphthalein indicator

#### **PART 2: Determination of an Equilibrium Constant**

- 6 - sample bottles with caps
- 3 - 50 mL burets or 3 - 5 mL graduated pipettes (needed in week 1)
- 1 - 50 mL buret (needed in week 2)
- 1 - 250 mL Erlenmeyer flask (needed in week 2)
- Ethyl acetate
- 3 M hydrochloric acid
- 1 M standard sodium hydroxide solution
- Phenolphthalein indicator

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## Procedures

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### PART 1: Determination of the Effect of Various Influences on the Position of Equilibrium

#### All in the First Week.

##### *a) Solubility Equilibria*

Take careful note of your observations – remember to also observe *before* mixing. In your report, you will be asked to explain your observations using Le Châtelier's Principle.

1. Place 0.5 mL (10 drops) of saturated sodium chloride solution in each of two dry (if possible) test tubes. *[This solution has been prepared for you by adding solid sodium chloride to water at room temperature until no more will dissolve, and then filtering off the small excess of solid remaining.]*
2. Add 10 drops of 12 M hydrochloric acid (**caution!**) to one of the test tubes, being careful not to spill any. Over about half an hour, observe and note the results.
3. To the other test tube, add 10 drops of 1 M hydrochloric acid and note what happens in this case, initially and after about half an hour.

##### *b) Complex Ion Equilibria*

At each step, carefully note the changes that occur. In your report, you will be asked to explain what you saw in terms of Le Châtelier's Principle.

1. *[This step may have been done for you.]* Make up a stock solution containing the complex ion  $[\text{FeSCN}]^{2+}$  by mixing 2 mL of 0.1 M iron(III) chloride solution with 2 mL of 0.1 M potassium thiocyanate solution. Dilute the solution to 100 mL in a 100 mL graduated cylinder.
2. Pour about 5 mL of the stock solution into each of four test tubes labeled 1,2,3 and 4.
3. Leave test tube 1 unchanged for comparison of colour intensity with the others.
4. To test tube 2, add about 1 mL of the 0.1 M  $\text{FeCl}_3$  solution & observe.
5. To test tube 3, add about 1 mL of the 0.1 M KSCN solution & observe.
6. To test tube 4, add 0.1 M  $\text{AgNO}_3$  dropwise until no further change is observed.

##### *c) Acid/Base Equilibria*

Observe and note carefully the results. In your report, you will be asked to account for the changes in terms of Le Châtelier's Principle.

1. In a fume hood, make up a dilute solution of ammonia: add 4 drops of concentrated ammonia solution to 100 mL of deionized water, then add 3 drops of phenolphthalein indicator. *[This solution may already be prepared for you.]* Place about 5 mL of the dilute ammonia solution into each of four test tubes.

2. Leave one of the test tubes unchanged for comparison.
3. To the second tube, add several small crystals of ammonium chloride (enough to cover the tip of a spatula) and stir.
4. To the third tube, add a few drops of concentrated hydrochloric acid (*use care*).
5. To the fourth tube, add 1 drop of concentrated ammonia solution (*use care*).

## **PART 2: Determination of an Equilibrium Constant**

### **First Week - Solution Preparation**

1. Clean and dry six (6) sample bottles.
2. Label them "A1", "A2", "B1", "B2", "C1" and "C2". [*Note: You should prepare one set of solutions A-C, and your partner should prepare the other set.*]
3. With the burets or graduated pipettes provided, add the following volumes of reagents as accurately as possible to the bottles (cap the bottles immediately after adding the ethyl acetate, because it is quite volatile):

<b>Reagent</b>	<b>Bottles A1 &amp; A2</b>	<b>Bottles B1 &amp; B2</b>	<b>Bottles C1 &amp; C2</b>
~3 M HCl solution (mL)	5.00	5.00	5.00
Distilled water (mL)	---	1.00	2.00
Ethyl acetate (mL)	5.00	4.00	3.00

Note that the concentration of the "3 M" hydrochloric acid is not exact. You will determine its exact concentration (*i.e.*, you will *standardize* it) in the second week.

4. Cap each bottle tightly and swirl the contents to mix the reagents; two immiscible layers will remain. Store the sealed bottles in your locker until the second week.

### **Second Week - Standardization of the ~3 M Hydrochloric Acid**

In order to determine the exact concentration of (*i.e.*, "standardize") the 3 M HCl solution used as the catalyst, you will titrate the HCl solution with a sodium hydroxide titrant solution whose concentration is accurately known (a *standard solution*). Be sure to use **the same** hydrochloric acid stock solution as you used to make up the solutions A, B and C.

*[Standardizing the HCl solution also allows you to perfect your titration technique before moving on to the samples you prepared last week.]*

1. Clamp a clean, dry 50 mL buret in a vertical position on a retort stand using a buret clamp. You can use pieces of rubber tubing to pad the glass so it fits snugly into the clamp. Test the stopcock – it should be closed tightly but still movable.

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- Carefully fill the buret with 1.00 M sodium hydroxide solution, pouring gently to avoid forming lots of bubbles. *[Bubbles cause error in volume readings; to dissipate them, wait a few minutes and/or very gently tap on the top of the buret.]* Ensure that the buret tip is filled by overfilling the buret and then draining it to the 0.00 mark (or slightly below it). Carefully read the level at the bottom of the meniscus to a precision of 0.05 mL; this is your initial volume measurement. After a few moments, check the volume again to be sure it is not leaking.
- Obtain a clean 250 mL Erlenmeyer flask and add to it exactly 5.00 mL of the 3 M HCl solution. Then add approximately 20 mL of distilled water and 3 drops of phenolphthalein solution. Place the flask under the buret and adjust the buret so the tip sits down 1-2 cm inside the mouth of the flask to minimize splashing.
- You are now ready to titrate. *[A technical tip: If you are right-handed, operate the buret's stopcock with your left hand (wrap your hand around the buret so the body of the stopcock is cradled in your palm; this gives your fingers finer control of the tap) and swirl the Erlenmeyer flask continually with your right hand.]* Slowly add the titrant solution from the buret to the HCl solution until you see the first hint of permanent pink coloration – this marks the endpoint. The endpoint will come suddenly after about 15 mL of titrant have been added – take care to not overshoot it; even one extra drop means you've added about 0.10 mL too much titrant. *[Hint: start off by adding titrant in a slow continuous stream, swirling constantly, until you notice the pink colour taking a little longer to dissipate with each swirl of the flask. Then start adding dropwise until you reach the endpoint.]* At the endpoint, read the level of titrant solution in the buret again. The volume added (called the *titre*) is obtained by the taking the difference between the endpoint volume reading and the initial volume reading.
- Each partner should repeat this titration procedure at least twice, until they have each obtained two successive runs with a difference in titre of less than 1%.

### Second Week - Determination of Total Moles of Acid Present at Equilibrium

These measurements are to be performed in the second week, *i.e.*, after the ethyl acetate hydrolysis reaction has reached equilibrium. Each partner should perform the titrations for one set of bottles A-C.

- Empty the contents of one sample bottle, *e.g.* A1, into a clean Erlenmeyer flask. Twice, rinse the bottle with 10 mL of distilled water, and carefully transfer the rinsings to the Erlenmeyer flask. *[The added water will not have a rapid effect on the reaction equilibrium, and is not to be included in subsequent calculations.]*
- Titrate the sample in the flask with 1.00 M sodium hydroxide solution using 3 drops of phenolphthalein as indicator, just as described above for the HCl solution alone. *[Remember: each sample here contains 5 mL of the ~3 M hydrochloric acid plus an unknown quantity of acetic acid; be sure to start with a full buret.]*
- Repeat this procedure (steps 1-2) for the other five sample bottles.

## Calculations

The Tables in the lab report section indicate the quantities to be recorded and calculated step by step. **It would be wise to construct an ICE table, even though one has not been provided.** The following paragraphs are intended to provide additional explanations.

To calculate equilibrium constants in general situations, it is necessary to know the **concentrations** (*i.e.*, moles/volume) of the reactants and products. In calculating the equilibrium constant for the reaction we are studying here, the volume terms happen to cancel out, so only the numbers of moles of each reactant and product are actually needed. (Remember also that the catalyst  $H^+$  would appear on both sides of the chemical equation, so it would therefore cancel out in the expression for the equilibrium constant.)

You must first calculate the number of moles of the reagents ethyl acetate (EtAc) and  $H_2O$  initially placed in each bottle. The initial number of moles of EtAc (*Table 3*) is easily calculated from the volume added, its density (0.893 g/mL) and its molar mass.

The water initially present in each bottle (*Table 3*) comes from two sources:

- 1) the pure water added (for bottles B and C); and
- 2) the 5.00 mL of hydrochloric acid (HCl gas dissolved in water) added to each bottle.

Determining the number of grams of pure water added is straightforward, because the density of pure water is 1.00 g/mL. However, in order to calculate the amount of water added via each 5.00 mL *aliquot* (a fancy word for portion) of hydrochloric acid, it is necessary to accurately know the concentration of the HCl solution, and more specifically, the number of moles of HCl present in each 5.00 mL aliquot – that is why you ran the first set of titrations. The titration data (*Table 1*) allow you to obtain the number of moles of HCl in each aliquot, based on the stoichiometry of the reaction between HCl and NaOH; thus, you can calculate the mass of HCl in each aliquot. The rest of each aliquot's mass was water. The total mass of each aliquot of HCl solution can be calculated using its density (measured for you:  $d = 1.05$  g/mL), and you can then calculate the mass of water in each aliquot by difference. Knowing this, the total amount of water (in moles) initially present in each bottle is calculated by taking the sum of the amount of pure water added and the amount of water added via the hydrochloric acid.

The next step is to determine the equilibrium concentrations of the species in the mixture. The second set of titration data (*Table 2*) enables you to calculate the total number of moles of acid species (HCl and HAc, which both react 1:1 with NaOH) present in the equilibrium mixtures. Because you have already figured out how many moles of HCl were added to the flasks in the first place (and *catalysts are not consumed*), your titration data lets you calculate how much HAc must also have been neutralized – this represents the quantity of HAc present in the sample when the EtAc hydrolysis reaction was at equilibrium (*Table 3*). At this point, you can use the reaction's stoichiometry to calculate how many moles of EtOH must also have been present at equilibrium. Similarly, you can calculate the number of moles of both EtAc and water that must have been consumed and, therefore, calculate how many moles of each remained at equilibrium. Once you have calculated the equilibrium quantities of all four substances (EtAc,  $H_2O$ , HAc, EtOH), you can calculate three values for the equilibrium constant,  $K_{eq}$ , for the hydrolysis of EtAc at room temperature, each obtained using a different set of initial reaction conditions (A-C). Within experimental error, they should yield the same value for  $K_{eq}$ .

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### Prelaboratory Questions

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**Note:** This prelab covers 2 experiments, therefore 10 marks have been assigned to it (instead of the usual 5). No prelab is required for the second week, except, of course, for those of you who for some reason might have missed the first week.

1. (4 marks) Write a brief summary, in your own words, of the purpose of the experiments (Parts 1 and 2) and the steps you will follow. One page - maximum.
2. (1.5 marks) Write the expression for the equilibrium constants for the following reactions:



3. (3 marks – 1 here, 2 on next page) For reaction 2(a) above, suppose equilibrium has been established. Explain how and why the position of equilibrium will shift in response to the following changes:
  - a. Additional nitrogen is added to the system.



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## Experiment 5: EXPERIMENTS IN CHEMICAL EQUILIBRIUM

### PART 1: Determination of the Effect of Various Influences on the Position of Equilibrium

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#### Laboratory Report

**Please note:** You must hand in your report for PART 1 when you return to the lab to complete PART 2 *Determination of an Equilibrium Constant*.

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#### Solubility Equilibria

1. (0.5 mark) What did you observe upon adding the concentrated hydrochloric acid to the sodium chloride solution?
2. (1 mark) Explain your observations.
3. (0.5 mark) What did you observe upon adding the dilute hydrochloric acid to the sodium chloride solution?
4. (1 mark) Explain your observations.

#### Complex Ion Equilibria

1. (0.5 mark) What did you observe upon adding  $\text{Fe}^{3+}$  ion to the  $[\text{FeSCN}]^{2+}$  solution?
2. (1 mark) Explain your observations.

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3. (0.5 mark) What did you observe upon adding  $\text{SCN}^-$  to the  $[\text{FeSCN}]^{2+}$  solution?
4. (1 mark) Explain your observations.
5. (0.5 mark) What did you observe upon adding  $\text{Ag}^+$  to the  $[\text{FeSCN}]^{2+}$  solution?
6. (1 mark) Explain your observations.

### Acid/Base Equilibria

1. (0.5 mark) What did you observe upon adding ammonium ion to the dilute ammonia solution containing phenolphthalein?
2. (1 mark) Explain your observations.
3. (0.5 mark) What did you observe upon adding hydrochloric acid to the dilute ammonia solution containing phenolphthalein?
4. (1 mark) Explain your observations.
5. (0.5 mark) What did you observe upon adding concentrated ammonia solution to the dilute ammonia solution containing phenolphthalein?
6. (1 mark) Explain your observations.

### Supplementary Questions

1. (1.5 marks) Explain in your own words why a saturated solution in contact with some undissolved solid is considered to be an equilibrium situation.

2. (1.5 marks) Rationalize how the addition of the silver nitrate to the  $\text{Fe}^{3+} / \text{SCN}^-$  equilibrium system had an effect even though neither  $\text{Ag}^+$  nor  $\text{NO}_3^-$  appears in the equation shown in the text.

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### PART 2: Determination of an Equilibrium Constant

#### Laboratory Report

(4 marks) <b>Table 1 – DATA: Titration of Stock Hydrochloric Acid</b> (each partner use his/her own data here)				
	Trial 1	Trial 2	Trial 3	Trial 4
Concentration of standard NaOH (M)				
Initial volume reading (mL)				
Final volume reading (mL)				
Volume of standard NaOH used (mL)				
Moles of NaOH used = moles HCl in 5.00 mL HCl solution				
Average of last 2 trials ( <i>see text</i> ) <b>To Table 3 (2)</b>				
Mass of HCl in 5.00 mL HCl solution				
Mass of 5.00 mL HCl solution (d = 1.05 g/mL)				
Mass of H <sub>2</sub> O in 5.00 mL HCl solution				
Moles H <sub>2</sub> O in 5.00 mL HCl solution <b>To Table 3 (1)</b>				

**(3 marks) Table 2 – DATA: Titration of Equilibrium Mixtures**  
(use both partners' data here)

	<b>Bottle A1</b>	<b>Bottle A2</b>	<b>Bottle B1</b>	<b>Bottle B2</b>	<b>Bottle C1</b>	<b>Bottle C2</b>
Initial volume reading (mL)						
Final volume reading (mL)						
Volume of standard NaOH used (mL)						
Moles of NaOH used = total moles HCl plus HAc						
Average total moles of HCl plus HAc. <i>To Table 3 (3)</i>						

**(4 marks) Table 3 – CALCULATIONS: To Find Equilibrium Constant**

	<b>Flasks A</b>	<b>Flasks B</b>	<b>Flasks C</b>
Initial moles of EtAc			
Initial moles of pure H <sub>2</sub> O added			
Initial moles of H <sub>2</sub> O <i>added with HCl (1)</i>			
Initial <i>total</i> moles of H <sub>2</sub> O			
Equilibrium moles of HAc plus HCl (3)			
Moles of HCl (2)			
Equilibrium moles of HAc			
Equilibrium moles of EtOH			
Equilibrium moles of EtAc			
Equilibrium moles of H <sub>2</sub> O			
Equilibrium constant *			
<b>Average equilibrium constant, <math>K_{eq}</math></b>			

*\*hint: make an ice table...*

### Supplementary Questions

1. (1.5 mark) What is the average % deviation of your  $K_{eq}$  values (see Chem 206 Exp.1A for a reminder of % deviation)? List three sources of experimental error that are likely responsible for this variability.
2. (0.5 mark) Based on your average value of the equilibrium constant, does the equilibrium favour the formation of the products ethanol and acetic acid, the starting reagents, or neither?
3. (1 mark) Suppose you stoppered the flasks after your titrations to determine the equilibrium constant and kept the solutions for another day so that equilibrium was re-established. What would happen to the amount of ethanol present? What would be the relationship of the original equilibrium constant you calculated to the new one?
4. (1 mark) Would the results of this experiment have been affected if you had forgotten to add the hydrochloric acid to the reaction bottles? Why or why not?