Student Name: ___ ________
Student Number:_____________
Partner's Name :_ __
Teaching Assistant's Name: ___Tom Lacelle ______

Lab Day: Wednesday night
Lab Week: 2
Date: September.21, 2016

Laboratory Report Form
Experiment 1.
An introduction to periodic trends or Is solubility periodic?

Checklist:

- **Raw Data Sheet written in pen, signed by TA and photocopy is attached**
- **Report Form typed and attached**

Student's Initials _____

Experimental Objectives:

In this experiment, the solubility of ionic hydroxide salts of Group two metal cations were investigated in order to discover its relationship with periodic trends.

Procedure:

All glassware used in the experiment was washed first, the burette was rinsed and then filled with about 50ml of standard HCL solution (0.002M). A few drops of Standard solution of Mg(OH)2 was used to rinse a volumetric pipette, then 50ml of saturated Mg(OH)2 was transferred to an Erlenmeyer flask. It is important to move the index finger quickly to control the volume level of volumetric pipette. A few drops of phenolphthalein indicator was added to the flask, then titrated with about 40ml of dilute HCL. Then 1.0ml of Ca(OH)2 was transferred to an Erlenmeyer flask and diluted with 50ml of distilled water, the above procedure was repeated. The last one started with 0.5ml of Sr(OH)2 diluted with 50ml of distilled water, same procedure also applied. Each of the titration with the hydroxide salts was repeated three times and a total of 9 sets of data was collected.

Results/Observation:

Table1. A summary of the moles and volume of HCl(0.002M) used in titration and concentration of OH-, K_{sp}, solubility of the ionic hydroxide salt $Mg(OH)_2$.

	Trial1	Trial2	Trial3
Initial Volume of HCL(ml)	1.95	1.80	1.92
Final Volume of HCL(ml)	43.08	43.19	43.15

	Trial1	Trial2	Trial3
Volume used(ml)	41.13	41.39	41.23
Average Volume(ml)	41.25		
Moles of Acid required to reach the endpoint(mol)	8.25×10^{-5}		
[OH-] for each trial (mol/L)	1.6452×10^{-3}	1.6548×10^{-3}	1.6492×10^{-3}
Average [OH-] (mol/L)	1.6497×10^{-3}		
K_{sp} for each trial	2.2265×10^{-9}	2.2657×10^{-9}	2.2428×10^{-9}
Average K_{sp}	2.245×10^{-9}		
Solubility (mg/L) for each trial	47.96	48.24	48.07
Average Solubility (mg/L)	48.09		

Table2. A summary of the moles and volume of HCl(0.002M) used in titration and concentration of OH-, K_{sp}, solubility of the ionic hydroxide salt $Ca(OH)_2$.

	Trial1	Trial2	Trial3
Initial Volume of HCL(ml)	1.98	13.59	28.01
Final Volume of HCL(ml)	13.59	28.01	41.72
Volume used(ml)	11.97	14.42	13.71
Average Volume(ml)	13.37		
Moles of Acid required to reach the endpoint(mol)	2.674×10^{-5}		
[OH-] for each trial (mol/L)	0.02394	0.02884	0.02742
Average [OH-] (mol/L)	0.02673		
K_{sp} for each trial	6.8603×10^{-6}	1.1994×10^{-5}	1.0308×10^{-5}
Average K_{sp}	9.7×10^{-6}		
Solubility (mg/L) for each trial	885.78	1067.08	1014.54
Average Solubility (mg/L)	989.38		

Table3. A summary of the moles and volume of HCl(0.002M) used in titration and concentration of OH-, K_{sp} and solubility of the ionic hydroxide salt $Sr(OH)_2$.

	Trial1	Trial2	Trial3
Initial Volume of HCL(ml)	2.58	0.91	25.11
Final Volume of HCL(ml)	25.11	21.71	48.50
Volume used(ml)	22.53	20.8	23.39
Average Volume(ml)	22.24		
Moles of Acid required to reach the endpoint(mol)	$4,448\times10^{-5}$		
[OH-] for each trial (mol/L)	0.09012	0.0832	0.09356
Average [OH-] (mol/L)	0.08896		
K_{sp} for each trial	3.66×10^{-4}	2.88×10^{-4}	4.0949×10^{-4}
Average K_{sp}	3.545×10^{-4}		
Solubility (mg/L) for each trial	3334.44	3078.4	3461.72
Average Solubility (mg/L)	3291.52		

- Saturated Mg(OH)2, Ca(OH)2 and Sr(OH)2 solutions are clear and colorless, HCl solution is also colorless. After the addition of phenolphthalein, the saturated solutions turned to dark pink. As the titration proceeds, they turned to pale pink at the end point.

- Sample Calculation (moles of acid required to reach the endpoint)
    - $[HCl]=2.0\times10^{-3}M$
    - 41.25ml = 0.04125L
    - $n = 0.04125L \times 2.0\times10^{-3}M = 8.25\times10^{-5}$ mol

- Sample calculation ($[OH^{-}]$, K_{sp} and solubility in mg/L)
    - $Ca(OH)_2 \rightleftharpoons Ca^{2+} + 2OH^{-}$
    - $Ca(OH)_2 + 2HCl \rightleftharpoons CaCl_2 + 2H_2O$
    - We used 1ml of saturated $Ca(OH)_2$ to dilute 50 ml of unsaturated solution and an average of 13.37 ml HCl.
    - moles of OH- in the sample = 0.01337L * 0.002mol/L $=2.674\times10^{-5}$ mol (because each mole of H+ can neutralize one OH-)

- **[OH−]** = (2.674×10^{-5} mol)/(0.001L)= 0.02674 (mol/L)
- [Ca^{2+}] = 0.02674 (mol/L)/2 = 0.01337 (mol/L)
- $K_{sp} = [Ca^{2+}][OH^{-}]^2$
- so $\mathbf{K_{sp}}$ = 0.01337(mol/L) * $[0.02674 \text{ (mol/L)}]^2 = 9.560\times10^{-6}$
- **Solubility** = 0.01337 (mol of $Ca(OH)_2$ /L) * 74g $Ca(OH)_2$/mol = 0.099g/L= 989.38mg/L

Discussion:

This experiment is trying to prove the theory that elements in the periodic table are arraigned and grouped in a manner in which the elements share similar chemical properties. In this experiment, the iconic hydroxide salts (Mg, Ca, Sr) of Group 2 metal cations are used.

The theoretical value of Ksp for Mg(OH)2, Ca(OH)2, Sr(OH)2 are, 5.61×10^{-12}, 5.5×10^{-6} , 6.4×10^{-3} , respectively, whereas my average experimental data for Mg(OH)2, Ca(OH)2, Sr(OH)2 are 2.245×10^{-9}, 9.7×10^{-6} , 3.545×10^{-4} respectively. Based on two sets of data, we can find a trend that , the Ksp value increases when moving down the periodic table on group2 metals. Two sets of data are close but still differ to some extent. The average experimental value of solubility(mg/L) for Mg(OH)2, Ca(OH)2, Sr(OH)2 are 48.09, 989.38, 3291.52 respectively. A trend that the solubility increases when going down the periodic table is found. And this actually makes sense.

Because the solubility of solids depend on lattice energy and the energy of hydration of ions. Lattice energy is required to separate the solids into ions and energy released when ions are hydrated is called energy of hydration. If the hydration energy is greater than lattice, the energy will help break apart and accelerate dissolution. For the alkaline earth metal hydroxides, as you go down the group, energy of hydration increases a lot, thus the hydroxide of metals are more soluble down the group.

The general trend between two sets of data are the same, but the experimental value are slightly differ from the theoretical value. There are a few sources a errors in this experiment. Firstly, volumetric petit was not rinsed properly, water may have remained on the wall. Secondly, when trying to get 1.0ml of saturated Ca(OH)2, in stead of using a volumetric pipet, a graduated cylinder was used, and it is not as accurate as a volumetric pipet. Thirdly, when transfer solution using a volumetric pipet from an Erlenmeyer flask, sometimes there is not enough solution remained in the flask, this created air bubbles in the volumetric pipet, and could affect experimental results.

For the subsequent labs, always make sure that there are enough solution remained in an Erlenmeyer flask and to rinse the volumetric pipet carefully with the relevant solution.

Questions:

1. Solubility would be greater in 0.01M hydrochloric acid than that in pure distilled water. The ionic equation is $Mg(OH)^2(s)$ <-> Mg^{2+} (aq)+ $2OH^-$ (aq). When adding HCl, H+ ions react with OH- to make water. Because the equilibrium is broke, and forced to move forward.

2.

- **K_{sp}** of Ba(OH)2 = 5×10^{-3}
- Pretend that we are gonna use 0.04L of HCl
- Then mol of OH- in the sample = 0.04L * 0.002mol/L = 0.00008mol (because one mole of OH- neutralize one mol of H+)
- Assume we are gonna use xL of saturated Ba(OH)2
- then $[OH^-]$ = 0.00008mol/ xL = (0.00008/x) (mol/L)
- $[Ba^{2+}]$ = (0.00008/x)/2 (mol/L) = (0.00008/2x) (mol/L)
- $K_{sp} = [Ba^{2+}] [OH^-]^2$
- $(0.00008/2x)* (0.00008/x)^2 = 5\times10^{-3}$
- Solve for x , $x = 3.7\times10^{-4}$ L
- x = 0.37 ml

3. a) $OH^-(aq) + H^+(aq) \longrightarrow H2O(l)$

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b) The analyses are Mg(OH)2, Ca(OH)2, Sr(OH)2.

c) The titrant solution is 0.002M HCl.

d) Indicator such as phenolphthalein is required.

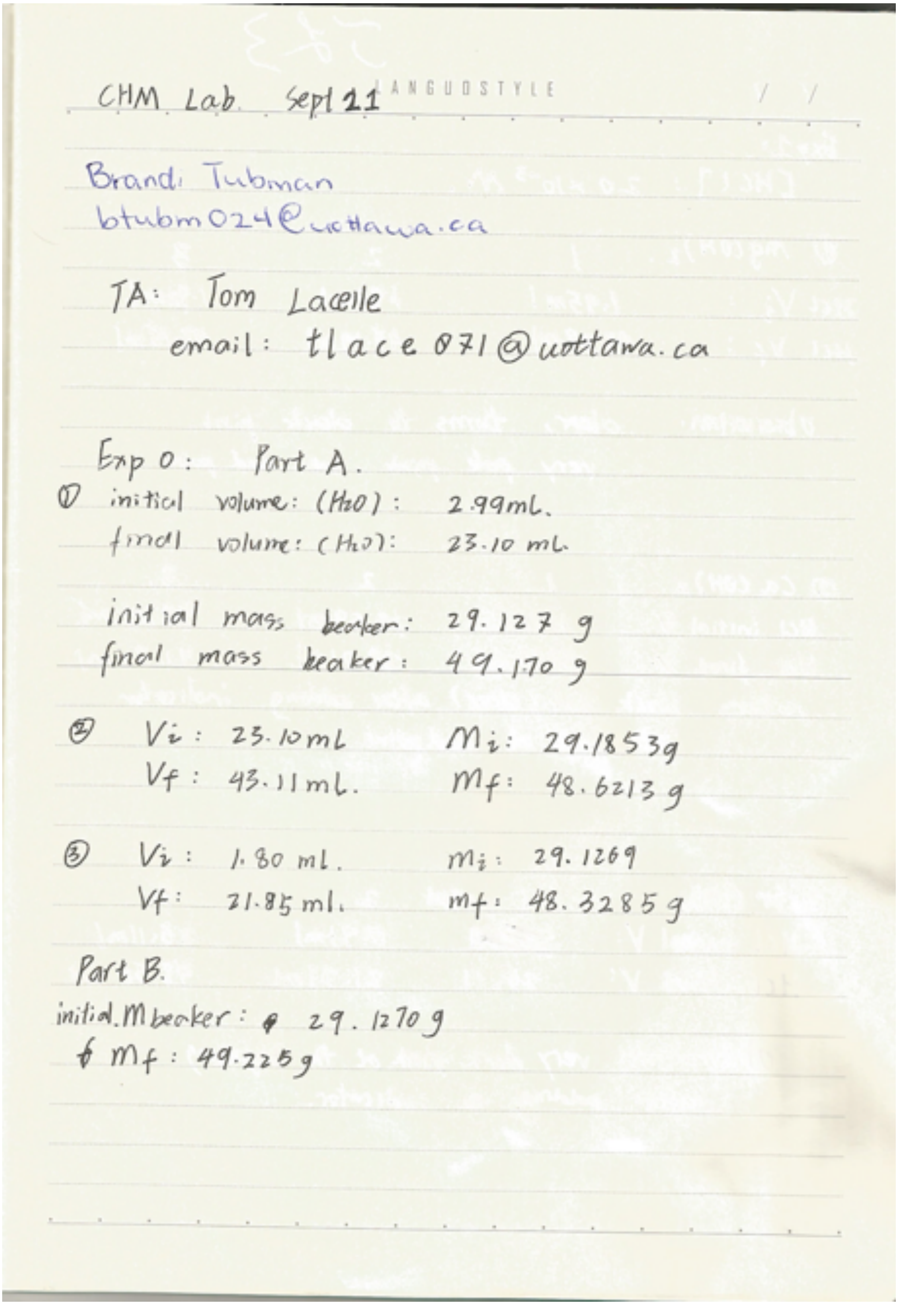
e) When a clear color change is observed, the titration is over.

f) An endpoint is indicated by clear color change of indicator at the end of a titration. An equivalence point is when the moles of a standard solution (titrant) equal the moles of a solution of unknown concentration (Analyte).

Reference:

- CHM2353 Lab manual Fall2016, University of Ottawa, Department of Chemistry,
- Calculate the K_{sp} of a Saturated Solution When Given Titration Data, http://www.chemteam.info/Equilibrium/calc-Ksp-from-titration-data.html (accessed Sept25, 2016)
- Chang, R. (2010). Chemistry: 10th Edition, Solubilities Within a Family Lab Report, http://lawaspect.com/solubilities-within-family-lab-report-3/ (accessed Sept25, 2016)
- Lange's Handbook, pps. 8-6 to 8-11; L.G. Sillen and A.E. Martell, "Stability Constants of Metal-Ion Complexes," The Chemical Society, London, 1964 (Special Publ. No. 17).

Raw Data



CHM Lab. Sept 21

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Exp 0: Part A.

① initial volume: (H_2O): 2.99mL.
final volume: (H_2O): 23.10 mL.

initial mass beaker: 29.127 g
final mass beaker: 49.170 g

② Vi: 23.10mL Mi: 29.1853g
Vf: 43.11mL. Mf: 48.6213 g

③ Vi: 1.80 ml. mi: 29.1269
Vf: 21.85 ml. mf: 48.3285 g

Part B.
initial.Mbeaker: 29.1270 g
mf: 49.225g

JL3

/ /

Exp 1:

[HCl] : 2.0×10^{-3} M.

① $Mg(OH)_2$.	1	2	3
HCl V_i :	1.95ml	1.80ml	1.92ml
HCl V_f :	43.08ml	43.19ml	43.15ml

observation. clear, turns to dark pink.
very pale pink at end point

② $Ca(OH)_2$.	1.	2.	3.
HCl initial V:	1.98ml	13.59ml	28.01.ml
HCl final V:	13.59ml	28.01ml.	41.72ml

color: dark pink (clear) after adding indicator
pale pink at endpoint.

③ $Sr(OH)_2$.	1	2	3
HCl initial V:	2.58ml [illegible]	0.91ml.	25.11ml
HCl final V:	26.11	21.71ml	48.50 ml.

observation: very dark pink at the beginning
after adding the indicator.