EXPERIMENT 13
**SOLUBILITY PRODUCT OF POTASSIUM ACID TARTRATE**

PURPOSE:
1. To determine experimentally the molar solubility of potassium acid tartrate in water and in a solution of
    potassium nitrate.
2. To examine the effect of a common ion on the solubility of a slightly soluble salt.

PRINCIPLES:
When a large amount of a slightly soluble ionic compound , $[M^+] [A^-]$, is mixed with water, the compound partially dissolves in water and a saturated solution is formed. Some of the solid remains undissolved.
An equilibrium occurs between the undissolved solid and the ions in the solution.

$$MA(s) \rightleftharpoons M^+(aq) + A^-(aq)$$
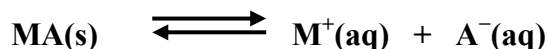
The equilibrium **constant** for the solubility process is called the Solubility Product Constant (Ksp)

$$Ksp = [M^+] [A^-]$$

The Ksp for a slightly soluble salt is determined by measuring the concentrations of the M^+ and A^- ions in a saturated solution. I this experiment, we can measure the concentration of the anion because it is a weak acid which can be titrated with a strong base.
Tartaric acid ($H_2C_4H_4O_6$, or H_2Tar) is a weak diprotic acid. If enough KOH is added to a solution of tartaric acid to neutralize only **half** the acidic hydrogens, the salt potassium acid tartrate ($KHC_4H_4O_6$ or KHTar) is obtained as a slightly soluble salt.
When KHTar is dissolved in water, the following equilibrium exists:

$$KHTar(s) \rightleftharpoons K^+(aq) + HTar^-(aq)$$

$HTar^-(aq)$ is a weak acid and as such it can be titrated with a strong base like sodium hydroxide and this neutralization reaction proceeds to completion:

$$HTar^-(aq) + OH^-(aq) \longrightarrow Tar^{2-}(aq) + H_2O(l)$$

Thus we can determine experimentally the concentration of a $HTar^-$ solution by titrating it with a sodium hydroxide solution of known concentration.
In this experiment, you will measure the solubility of KHTar in two solvent systems: pure water and 0.10 M KNO_3. The solubility of KHTar will be determined by titration of the $HTar^-$ with standard NaoH.
The Ksp of KHTar in these solvents is given by:

$$Ksp = [K^+] [HTar^-]$$

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In pure water the only sources of potassium and acid tartrate ions are the dissolved KHTar, which in view of the negligible amount of dissociation of $HTar^-$, leads to the following relation:

$$[K^+] = [HTar^-]$$

The concentration of the $HTar^-$ is found by knowing how many moles of $HTar^-$ are contained in a known volume of solution, Vs.
Since $HTar^-$ reacts with the sodium hydroxide in a 1:1 mole ratio:

$$HTar^-(aq) \ + \ OH^-(aq) \ \longrightarrow \ Tar^{2-}(aq) \ + \ H_2O(l) \, ,$$

the moles of $HTar^-$ are equal to the moles of NaOH used to reach the endpoint .

$$\textbf{moles } HTar^- = \textbf{moles NaoH} = \ (M_{NaOH}) \ \textbf{(Volume of NaOH solution)}$$

$$\boxed{[HTar^-] \ = \ \frac{(M_{NaOH}) \text{ (Volume of NaOH solution)}}{\text{Volume of KHTar solution}}}$$

In the 0.10 M KNO_3 solution, the concentration of $HTar^-$ is found in the same fashion. However, there are now two sources of potassium ions: from the dissolved KHTar and from the strong electrolyte , KNO_3 , which is totally dissociated into potassium and nitrate ions in solution.
The concentration of potassium ions is given by:

$$[K^+] = [KNO_3] \ + \ [HTar^-]$$

It follows that:

$$Ksp = \Big[\ [KNO_3] \ + \ [HTar^-] \ \Big] \Big[\ HTar^- \ \Big]$$

Recall that the following equilibrium exists in solution:

$$KHTar(s) \ \rightleftharpoons \ K^+(aq) \ + \ HTar^-(aq)$$

The addition of K^+ ions from KNO_3 (potassium nitrate) will shift the solubility equilibrium to the left, thus lowering the solubility of KHTar.
This lowering of solubility by addition of a salt containing a common ion is called the **common ion effect.**

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PROCEDURE:

Use a centigram or milligram balance to weigh out about 2 g of potassium acid tartrate into each of two different 250 mL labeled Erlenmeyer flasks.

One flask should be labeled "**WATER**" and the other "**KNO₃**"

To one flask add about 150 mL of deionized water and to the other about 150 mL of 0.10 M KNO_3 solution.

Agitate the flasks by swirling them vigorously in turn for about 15 minutes. At the end of this time, let the solid settle out while measuring the temperature in each flask (they should be the same).

Then filter the supernatant liquid in each flask, collecting the clear, colorless filtrate solution. Make sure that you use a separate filter paper, a clean dry filter funnel and a clean, dry labeled beaker to receive the filtrate from each flask.

Rinse and fill your buret with the standardized NaOH solution. Rinse out your clean 100 mL graduated cylinder with a few mL of one of the filtered KHTar solutions.

Then carefully measure out 50.0 mL portions of this solution (KHTar in water) into two labeled 250 mL Erlenmeyer flasks and add two drops of phenolphthalein indicator solution. Titrate the two samples to the "pale pink endpoint" recording the volume used for the two titrations.

After completing these first two titrations, rinse the flasks and graduated cylinder with deionized water and treat the other filtered KHTar solution (KHTar in 0.10 M KNO_3 solution) in similar way.

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Name: _____

Partner: _____

REPORT FORM
Part I: Molar Solubility of Potassium Acid Tartrate (KHTar) in Water

Molarity of Standardized NaOH solution: _____ M

Temperature of Solution: _____ 0 C

	Titrations	
	First	Second
Final buret reading (mL):	_____	_____
Initial buret reading (mL):	_____	_____
Volume of titrant (mL):	_____	_____
Moles of NaOH used:	_____	_____
Moles of KHTar used to neutralize the above nimber of moles of NaOH:	_____	_____
Volume of saturated solution of KHTar (L)	_____	_____
Molarity of saturated solution of KHTar (M):	_____	_____

Molar Solubility of
KHTar (M), Average (Experimental)   _____ M

SOLUBILITY PRODUCT OF POTASSIUM ACID TARTRATE

Calculate Ksp of KHTar in water from your experimental data

1.   Write the Solubility Equilibrium Expression for **KHTar.**

2.   Write the Solubility Product expression for **KHTar**

3.   Calculate Ksp for KHTar , in water, based on the experimentally determined molar
        solubility of KHTar

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Part II: Molar Solubility of Potassium Acid Tartrate (KHTar) in 0.10 M KNO3

Molarity of Standardized NaOH solution: _____ M

Temperature of Solution: _____ 0 C

 Titrations

	First	Second
Final buret reading (mL):	_____	_____
Initial buret reading (mL):	_____	_____
Volume of titrant (mL):	_____	_____
Moles of NaOH used:	_____	_____
Moles of KHTar used to neutralize the above nimber of moles of NaOH:	_____	_____
Volume of saturated solution of KHC4H4O4 (L)	_____	_____
Molarity of saturated solution of KHTar (M):	_____	_____

Molar Solubility of
KHTar (M),in 0.10 M KNO3 Average (**Experimental**) [   ] M

SOLUBILITY PRODUCT OF POTASSIUM ACID TARTRATE

Calculate the theoretical Molar solubility of KHC4H4O4 in 0.10 M KNO3 using the Ksp of KHC4H4O4 in water determined in Part I (Show calculations below)

Solubility Expression:
Concentrations:
Starting

Change:

Equilibrium:

Show calculations below:

Molar Solubility of
KHTar (M), in 0.10 M KNO3 (**Calculated**)   M

Calculation of Percent Error
Molar solubility of KHTar in 0.10 M KNO3
(experimental)   **M**

Molar solubility of KHTarO4 in 0.10 M KNO3
(calculated)   **M**

Percent Error: **%**

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CONCLUSIONS:
1. (a) What is the <u>experimentally</u> determined **Molar solubility of KHTar in water ?**
    (Part I)

          _____ **M**

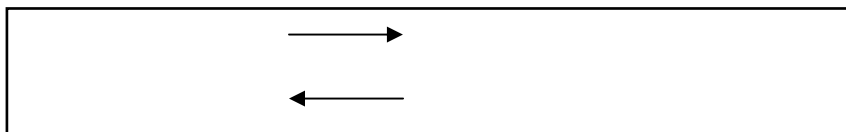 (b) What is the <u>experimentally</u> determined **Molar solubility of KHTar in 0.10M KNO3 ?**
    (Part II)

          _____ **M**

  (c) Explain why the two Molar Solubilities listed in 1(a) and 2(b) above are quite different

    2. Use your experimental results to predict the **Solubility of KHTar in 0.10 M Sodium Acid Tartrate (NaHTar)**

Solubility Equilibrium:

Concentrations:

Starting

Change:

Equilibrium: