FLINN SCIENTIFIC

Name________________________________

Solubility Matters Data Sheet

Preparation Table

Mass of sodium thiosulfate pentahydrate, g	
Molecular weight of sodium thiosulfate pentahydrate, g/mole	
Moles of sodium thiosulfate pentahydrate, moles	
Mass of aqueous solution (sodium thiosulfate plus water), g	
Concentration of standard sodium thiosulfate solution (moles of thiosulfate per gram of solution), moles/g	

Titration Table

	Trial 1	Trial 2	Trial 3
Mass of weight buret (initial), g			
Mass of weight buret (final), g			
Mass of thiosulfate solution added, g			
Moles of thiosulfate added			
Moles of iodate present			
Volume of *saturated* calcium iodate solution tested, mL			
Concentration of iodate ion in saturated solution, moles/L			
Concentration of calcium ion, moles/L			
Solubility product constant (K_{sp})			

Calculations and Post-Lab Questions

Calculations

(Show all work on a separate sheet of paper and record your answers in the appropriate spaces on the Data Sheet.)

Preparation

1. Calculate the *molecular weight* of sodium thiosulfate pentahydrate and use this value to convert the mass of sodium thiosulfate present to *moles*.
2. Calculate the *concentration* of the standard thiosulfate solution in *moles of thiosulfate per gram of solution*. *Note:* The units of concentration were chosen to be moles per gram of solution because the amount of solution added was measured in grams.

Titration Table

3. Find the *mass of sodium thiosulfate solution* used for each titration trial.
4. Multiply the mass of thiosulfate solution used by the concentration of solution to obtain the *number of moles* of thiosulfate solution added for each titration trial.
5. Based on the stoichiometry of the balanced chemical equation for the overall reaction of thiosulfate ion with iodate (Equation 6), calculate the *number of moles of iodate* present in each titration sample.
6. For each titration sample, divide the number of moles of iodate by the volume of the saturated calcium iodate solution tested to calculate the *concentration of iodate ion* in moles/L.
7. Use the ratio of calcium to iodate ions in the molecular formula of calcium iodate as a conversion factor to calculate the *concentration of calcium ion* present in the saturated solution of calcium iodate.
8. Use Equation 2 to calculate the value of the *solubility product constant* (K_{sp}) for each titration trial.

Post-Lab Questions *(Answer on a separate sheet of paper.)*

1. Determine the *mean value* of the solubility product for calcium iodate and the *average deviation* for your three trials.
2. The literature value of the K_{sp} for calcium iodate is 6.47×10^{-6} at 25 °C. Use the following equation to calculate the percent error in your determination of the solubility product for calcium iodate. *Note:* Do not be alarmed if your average deviation is very large. This is not uncommon in K_{sp} experiments, even under the most careful conditions.

$$\text{percent error} = \frac{|\text{ experimental value } - \text{ literature value }|}{\text{literature value}} \times 100\%$$

3. A student did not notice that the test tube in which she obtained her saturated calcium iodate solution had some water in it. Discuss the effect of this error on the calculated value of the solubility product—was her K_{sp} value likely to be higher or lower as a result?
4. Calcium oxalate is a sparingly soluble ionic compound composed of calcium ions (Ca^{2+}) and oxalate ions ($C_2O_4^{2-}$). It is a main culprit in kidney stone disease, due to the formation of calcium oxalate crystals in the urine. A typical value for the concentration of calcium in urine is 0.3 grams per liter. Given a K_{sp} value for calcium oxalate of 1.3×10^{-8}, calculate the minimum concentration of oxalate ion in urine that could lead to precipitation of calcium oxalate.