Ksp Problem Set

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1. Calculate the molar solubility and Ksp for each of the following.

  a. $6.11 \cdot 10^{-4}$ g $BaSO_4$ dissolves in 250.0 mL of water at 25 °C

$$\text{mass} := 6.11 \cdot 10^{-4} \cdot \text{gm}$$

$$V := 250.0 \cdot \text{mL}$$

$$MW_{BaSO4} := (137.329 + 32.066 + 15.9994 \cdot 4) \cdot \text{gm} \cdot \text{mole}^{-1}$$

$$\text{mole}_{BaSO4} := \frac{\text{mass}}{MW_{BaSO4}} \qquad \text{mole}_{BaSO4} = 2.618 \times 10^{-6} \, \text{mol}$$

$$M := \frac{\text{mole}_{BaSO4}}{V} \qquad M = 1.047 \times 10^{-5} \, \text{mole} \cdot \text{liter}^{-1}$$

The molar solubility is simply the concentration of the solution at equlibrium:

$$M = 1.047 \times 10^{-5} \, \text{mole} \cdot \text{liter}^{-1}$$

Since $BaSO_4 \rightarrow Ba^{2+} + SO_4^{2-}$

$$K_{sp\_BaSO4} := M^2$$

$$K_{sp\_BaSO4} = 1.097 \times 10^{-10} \left(\frac{\text{mole}}{\text{liter}} \right)^2$$

b.0.3295 g BaF_2 dissolves in 250.0 mL of water at 25 °C

$$mass := 0.3295 \cdot gm$$

$$V := 250.0 \cdot mL$$

$$MW_{BaF2} := (137.329 + 18.998 \cdot 2) \cdot gm \cdot mole^{-1}$$

$$mole_{BaF2} := \frac{mass}{MW_{BaF2}} \qquad mole_{BaF2} = 1.879 \times 10^{-3} mol$$

$$M := \frac{mole_{BaF2}}{V} \qquad M = 7.517 \times 10^{-3} mole \cdot liter^{-1}$$

The molar solubility is simply the concentration of the solution at equlibrium:

$$M = 7.517 \times 10^{-3} mole \cdot liter^{-1}$$

Since BaF_2 --> Ba^{2+} + 2 F^-

$$K_{sp\_BaF2} := 4 \cdot M^3$$

$$K_{sp\_BaF2} = 1.699 \times 10^{-6} \left(\frac{mole}{liter} \right)^3$$

c. 4.807*10-4 g AgCl dissolves in 250.0 mL of water at 25 °C

$$\text{mass} := 4.807 \cdot 10^{-4} \cdot \text{gm}$$

$$V := 250.0 \cdot \text{mL}$$

$$MW_{AgCl} := (107.8682 + 35.4527) \cdot \text{gm} \cdot \text{mole}^{-1}$$

$$\text{mole}_{AgCl} := \frac{\text{mass}}{MW_{AgCl}} \qquad \text{mole}_{AgCl} = 3.354 \times 10^{-6} \, \text{mol}$$

$$M := \frac{\text{mole}_{AgCl}}{V} \qquad M = 1.342 \times 10^{-5} \, \text{mole} \cdot \text{liter}^{-1}$$

The molar solubility is simply the concentration of the solution at equlibrium:

$$M = 1.342 \times 10^{-5} \, \text{mole} \cdot \text{liter}^{-1}$$

Since AgCl --> Ag+ + Cl-

$$K_{sp\_AgCl} := M^2$$

$$K_{sp\_AgCl} = 1.8 \times 10^{-10} \left(\frac{\text{mole}}{\text{liter}} \right)^2$$

d. 1.667 g CaF_2 dissolves in 100.0 L of water at 25 °C
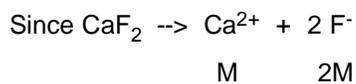
$$mass := 1.667 \cdot gm$$

$$V := 100 \cdot L$$

$$MW_{CaF2} := (40.078 + 2 \cdot 18.998) \cdot gm \cdot mole^{-1}$$

$$mole_{CaF2} := \frac{mass}{MW_{CaF2}} \qquad mole_{CaF2} = 0.021 \, mol$$

$$M := \frac{mole_{CaF2}}{V} \qquad M = 2.135 \times 10^{-4} \, mole \cdot liter^{-1}$$

The molar solubility is simply the concentration of the solution at equilibrium:

$$M = 2.135 \times 10^{-4} \, mole \cdot liter^{-1}$$

Since $CaF_2 \longrightarrow Ca^{2+} + 2 F^-$

$$\qquad\qquad M \qquad 2M$$

$$K_{sp} = M \cdot (2 \cdot M)^2$$

$$K_{sp\_CaF2} := 4M^3$$

$$K_{sp\_CaF2} = 3.894 \times 10^{-11} \left(\frac{mole}{liter}\right)^3$$

2.Use the K_{sp} values calculated above to determine the mass of each that will dissolve in 2.0 L of deionized water.

a.BaSO$_4$

$$K_{sp\_BaSO4} = X^2$$

$$X := \sqrt{K_{sp\_BaSO4}}$$

$$X = 1.047 \times 10^{-5} \, \text{mole} \cdot \text{liter}^{-1}$$

$$\text{mole\_disolved} := X \cdot 2 \cdot \text{liter}$$

$$\text{mole\_disolved} = 2.094 \times 10^{-5} \, \text{mol}$$

$$\text{mass\_disolved} := \text{mole\_disolved} \cdot MW_{BaSO4}$$

$$\text{mass\_disolved} = 4.888 \times 10^{-3} \, \text{gm}$$

b.BaF$_2$

$$K_{sp\_BaF2} = 4 \cdot X^3$$

$$X := \left(\frac{K_{sp\_BaF2}}{4} \right)^{\left(\frac{1}{3} \right)}$$

$$X = 7.517 \times 10^{-3} \, \text{mole} \cdot \text{liter}^{-1}$$

$$\text{mole\_disolved} := X \cdot 2 \cdot \text{liter}$$

$$\text{mole\_disolved} = 0.015 \, \text{mol}$$

$$\text{mass\_disolved} := \text{mole\_disolved} \cdot MW_{BaF2}$$

$$\text{mass\_disolved} = 2.636 \, \text{gm}$$

c.AgCl

$$K_{sp\_AgCl} = X^2$$

$$X := \sqrt{K_{sp\_AgCl}}$$

$$X = 1.342 \times 10^{-5} \, mole \cdot liter^{-1}$$

$$mole\_disolved := X \cdot 2 \cdot liter$$

$$mole\_disolved = 2.683 \times 10^{-5} \, mol$$

$$mass\_disolved := mole\_disolved \cdot MW_{AgCl}$$

$$mass\_disolved = 3.846 \times 10^{-3} \, gm$$

d.CaF$_2$

$$K_{sp\_CaF2} = 4 \cdot X^3$$

$$X := \left(\frac{K_{sp\_CaF2}}{4} \right)^{\left(\frac{1}{3} \right)}$$

$$X = 2.135 \times 10^{-4} \, mole \cdot liter^{-1}$$

$$mole\_disolved := X \cdot 2 \cdot liter$$

$$mole\_disolved = 4.27 \times 10^{-4} \, mol$$

$$mass\_disolved := mole\_disolved \cdot MW_{CaF2}$$

$$mass\_disolved = 0.033 \, gm$$

3.Use the K_{sp} values calculated above to determine.

   a.The mass of $BaSO_4$ that will dissolve in 500 mL of 0.10 M sodium sulfate.

   Sodium sulfate is a salt, it dissociates completely according to

   Na2SO4 --> 2Na1+ + SO42-

   $$C_{Na2SO4} := 0.1 \cdot mole \cdot liter^{-1}$$

   $$C_{SO4\_initial} := C_{Na2SO4}$$

   Next the equlibrium for the solubility of $BaSO_4$

   BaSO4 <--> Ba2+ + SO42-

   $$K_{sp\_BaSO4} = X \cdot (C_{SO4\_initial} + X)$$

   If you assume that X is small this rearanges to:

   $$K_{sp\_BaSO4} = X \cdot C_{SO4\_initial}$$

   $$X := \frac{K_{sp\_BaSO4}}{C_{SO4\_initial}}$$

   $$X = 1.097 \times 10^{-9} mole \cdot liter^{-1}$$

   Which shows that the assumption was reasonable

   From this we can calculate the mass of $BaSO_4$ that dissolves

   $$V := 500 \cdot mL$$

   $$mole_{BaSO4} := V \cdot X \qquad\qquad mole_{BaSO4} = 5.483 \times 10^{-10} mol$$

   $$mass_{BaSO4} := mole_{BaSO4} \cdot MW_{BaSO4}$$

   $$mass_{BaSO4} = 1.28 \times 10^{-7} gm$$

b.The mass of BaF_2 that will dissolve in 250 mL of 0.10 M potassium fluoride.

Potassium fluoride is a salt, it dissociates completely according to
$$KF \rightarrow K1+ \; + \; F1-$$

$$C_{KF} := 0.1 \cdot mole \cdot liter^{-1}$$

$$C_{F\_initial} := C_{KF}$$

Next the equilibrium for the solubility of BaF_2
$$BaF2 \longleftrightarrow Ba2+ \; + \; 2 \; F1-$$

$$K_{sp\_BaF2} = X \cdot \left(C_{F\_initial} + 2X\right)^2$$

If you assume that X is small this rearanges to:

$$K_{sp\_BaF2} = X \cdot C_{F\_initial}^2$$

$$X := \frac{K_{sp\_BaF2}}{C_{F\_initial}^2}$$

$$X = 1.699 \times 10^{-4} \; mole \cdot liter^{-1}$$

Which shows that the assumption was reasonable

From this we can calculate the mass of BaF_2 that dissolves

$$V := 250 \cdot mL$$

$$mole_{BaF2} := V \cdot X \qquad\qquad mole_{BaF2} = 4.248 \times 10^{-5} \; mol$$

$$mass_{BaF2} := mole_{BaF2} \cdot MW_{BaF2}$$

$$mass_{BaF2} = 7.448 \times 10^{-3} \; gm$$

c. The mass of AgCl that will dissolve in 5.0 L of 10^{-4} M lithium chloride.

Lithium is a salt, it dissociates completely according to
LiCl --> Li1+ + Cl1-

$$C_{LiCl} := 10^{-4} \cdot mole \cdot liter^{-1}$$

$$C_{Cl\_initial} := C_{LiCl}$$

Next the equlibrium for the solubility of AgCl
AgCl <--> Ag1+ + Cl1-

$$K_{sp\_AgCl} = X \cdot (C_{Cl\_initial} + X)$$

If you assume that X is small this rearanges to:

$$K_{sp\_AgCl} = X \cdot C_{Cl\_initial}$$

$$X := \frac{K_{sp\_AgCl}}{C_{Cl\_initial}}$$

$$X = 1.8 \times 10^{-6} \, mole \cdot liter^{-1}$$

Which shows that the assumption was reasonable

From this we can calculate the mass of $BaSO_4$ that dissolves

$$V := 5 \cdot L$$

$$mole_{AgCl} := V \cdot X \qquad\qquad mole_{AgCl} = 9 \times 10^{-6} \, mol$$

$$mass_{AgCl} := mole_{AgCl} \cdot MW_{AgCl}$$

$$mass_{AgCl} = 1.29 \times 10^{-3} \, gm$$

d.The mass of CaF_2 that will dissolve in 5000.0 L of 0.10 M potassium fluoride.

Potassium fluoride is a salt, it dissociates completely according to
KF --> K1+ + F1-

$$C_{KF} := 0.1 \cdot mole \cdot liter^{-1}$$

$$C_{F\_initial} := C_{KF}$$

Next the equilibrium for the solubility of CaF_2
CaF2 <--> Ca2+ + 2 F1-

$$K_{sp\_CaF2} = X \cdot (C_{F\_initial} + 2X)^2$$

If you assume that X is small this rearanges to:

$$K_{sp\_CaF2} = X \cdot C_{F\_initial}^2$$

$$X := \frac{K_{sp\_CaF2}}{C_{F\_initial}^2}$$

$$X = 3.894 \times 10^{-9} \, mole \cdot liter^{-1}$$

Which shows that the assumption was reasonable

From this we can calculate the mass of CaF_2 that dissolves

$$V := 5000 \cdot L$$

$$mole_{CaF2} := V \cdot X \qquad\qquad mole_{CaF2} = 1.947 \times 10^{-5} \, mol$$

$$mass_{CaF2} := mole_{CaF2} \cdot MW_{CaF2}$$

$$mass_{CaF2} = 1.52 \times 10^{-3} \, gm$$

4. Use the K_{sp} values calculated in the previious problem set to determine the mass of precipitate formed and the concentration all ions remaining in soultion when:

a. 50.0 mL of 0.10 M barium acetate is mixed with 100.0 mL of 0.10 M sodium sulfate.

Sodium sulfate is a salt, it dissociates completely according to

$$Na_2SO_4 \longrightarrow 2Na^{1+} + SO_4^{2-}$$

$$C_{Na2SO4} := 0.1 \cdot mole \cdot liter^{-1}$$

$$C_{SO4\_initial} := C_{Na2SO4} \cdot \frac{50 \cdot mL}{100 \cdot mL} \qquad\qquad C_{SO4\_initial} = 0.05\, mole \cdot liter^{-1}$$

Barium Acetate is a salt, it dissociates completely according to

$$Ba(CH3COO)2 \longrightarrow Ba2+ + 2\, CH3COO-$$

$$C_{BaCH3COO} := 0.1 \cdot mole \cdot liter^{-1}$$

$$C_{Ba\_initial} := C_{BaCH3COO} \cdot \frac{50 \cdot mL}{100 \cdot mL} \qquad\qquad C_{Ba\_initial} = 0.05\, mole \cdot liter^{-1}$$

Since Barium Sulfate is insoluble, it will for a precipitate. In this direction the reaction has a large value for K, so we'll assume it goes to completion. Since the moles of barium and sulfate are equal there is no limiting reagent.

$$V := 100 \cdot mL$$

$$mole_{BaSO4} := C_{Ba\_initial} \cdot V \qquad\qquad mole_{BaSO4} = 5 \times 10^{-3}\, mol$$

$$mass_{BaSO4} := mole_{BaSO4} \cdot MW_{BaSO4} \qquad\qquad mass_{BaSO4} = 1.167\, gm$$

Next the equlibrium for the solubility of $BaSO_4$

$$BaSO4 \longleftrightarrow Ba2+ + SO42-$$

$$K_{sp\_BaSO4} = X^2$$

$$X := \left(K_{sp\_BaSO4}\right)^{\frac{1}{2}}$$

$$X = 1.047 \times 10^{-5}\, mole \cdot liter^{-1}$$

From this we can give the concentration of all the ions in solution:

$$C_{Ba} := X \qquad\qquad\qquad C_{Ba} = 1.047 \times 10^{-5}\, mol\, liter^{-1}$$

$$C_{SO4} := X \qquad\qquad\qquad C_{SO4} = 1.047 \times 10^{-5}\, mol\, liter^{-1}$$

$$C_{CH3COO} := C_{BaCH3COO} \cdot \frac{50 \cdot mL}{100 \cdot mL} \qquad\qquad C_{CH3COO} = 0.05\, mol\, liter^{-1}$$

$$C_{Na} := C_{Na2SO4} \cdot \frac{50 \cdot mL}{100 \cdot mL} \qquad\qquad C_{Na} = 0.05\, mol\, liter^{-1}$$

b.100.0 mL of 0.15 M barium chlorate is mixed with 250.0 mL of 0.10 M potassium fluoride.

potassium fluorie is a salt, it dissociates completely according to

$$KF \longrightarrow K^{1+} + F^{1-}$$

$$C_{KF} := 0.1 \cdot mole \cdot liter^{-1}$$

$$C_{F\_initial} := C_{KF} \cdot \frac{250 \cdot mL}{350 \cdot mL} \qquad\qquad C_{F\_initial} = 0.071 \, mole \cdot liter^{-1}$$

Barium Chlorate is a salt, it dissociates completely according to

$$Ba(ClO_3)_2 \longrightarrow Ba^{2+} + 2\,ClO_3^{1-}$$

$$C_{BaClO3} := 0.15 \cdot mole \cdot liter^{-1}$$

$$C_{Ba\_initial} := C_{BaClO3} \cdot \frac{100 \cdot mL}{350 \cdot mL} \qquad\qquad C_{Ba\_initial} = 0.043 \, mole \cdot liter^{-1}$$

Since Barium Fluoride is insoluble, it will form a precipitate. In this direction the reaction has a large value for K, so we'll assume it goes to completion. Since the concentration of the barium is smaller, it is the limiting reagent.

$$V := 350 \cdot mL$$

$$mole_{BaF2} := C_{Ba\_initial} \cdot V \qquad\qquad mole_{BaF2} = 0.015 \, mol$$

$$mass_{BaF2} := mole_{BaF2} \cdot MW_{BaF2} \qquad\qquad mass_{BaF2} = 2.63 \, gm$$

Next the equlibrium for the solubility of BaF_2

$$BaF_2 \longleftrightarrow Ba^{2+} + 2\,F^{1-}$$

$$C_F := C_{F\_initial} - C_{Ba\_initial}$$

$$K_{sp\_BaF2} = X \cdot (C_F + 2 \cdot X)^2$$

Assuming X is small, this simplifies to:

$$K_{sp\_BaF2} = X \cdot (C_F)^2$$

$$X := \frac{K_{sp\_BaF2}}{C_F{}^2}$$

$$X = 2.082 \times 10^{-3} \, mole \cdot liter^{-1}$$

From this we can give the concentration of all the ions in solution:

$$C_{Ba} := X \qquad\qquad C_{Ba} = 2.082 \times 10^{-3} \, mol \, liter^{-1}$$

$$C_F = 0.029 \, mol \, liter^{-1}$$

$$C_{ClO3} := 2 C_{BaClO3} \cdot \frac{100 \cdot mL}{350 \cdot mL} \qquad\qquad C_{ClO3} = 0.086 \, mol \, liter^{-1}$$

$$C_K := C_{KF} \cdot \frac{250 \cdot mL}{350 \cdot mL} \qquad\qquad C_K = 0.071 \, mol \, liter^{-1}$$

c.75.0 mL of 0.015 M silver nitrate is mixed with 1.0 L of 10^{-4} M lithium chloride.

Lithium Chloride is a salt, it dissociates completely according to

$LiCl \longrightarrow Li^{1+} + Cl^{1-}$
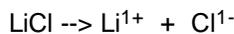
$C_{LiCl} := 10^{-4} \cdot mole \cdot liter^{-1}$

$C_{Cl\_initial} := C_{LiCl} \cdot \dfrac{1000 \cdot mL}{1075 \cdot mL}$
$\qquad\qquad\qquad$
$C_{Cl\_initial} = 9.302 \times 10^{-5} \, mole \cdot liter^{-1}$

Silver Nitrate is a salt, it dissociates completely according to

$AgNO3 \longrightarrow Ag+ + NO31-$

$C_{AgNO3} := 0.015 \cdot mole \cdot liter^{-1}$

$C_{Ag\_initial} := C_{AgNO3} \dfrac{75 \cdot mL}{1075 \cdot mL}$
$\qquad\qquad\qquad$
$C_{Ag\_initial} = 1.047 \times 10^{-3} \, mole \cdot liter^{-1}$

Since silver chloride is insoluble, it will form a precipitate. In this direction the reaction has a large value for K, so we'll assume it goes to completion. However, the concentrations are pretty small here. We better check Q to see if a precipitate will form:

$Q := C_{Ag\_initial} \cdot C_{Cl\_initial}$

$Q = 9.735 \times 10^{-8} \left(\dfrac{mole}{liter} \right)^2$

$K_{sp\_AgCl} = 1.8 \times 10^{-10} \left(\dfrac{mole}{liter} \right)^2$

Since Q is greater than K, a precipitate will form and chloride will be the limiting reagent.

$V := 1075 \cdot mL$

$mole_{AgCl} := C_{Cl\_initial} \cdot V$
$\qquad\qquad\qquad$
$mole_{AgCl} = 1 \times 10^{-4} \, mol$

$mass_{AgCl} := mole_{AgCl} \cdot MW_{AgCl}$
$\qquad\qquad\qquad$
$mass_{AgCl} = 0.014 \, gm$

Next the equilibrium for the solubility of AgCl

$AgCl \longleftrightarrow Ag1+ + Cl1-$

$C_{Cl} := C_{Cl\_initial} - C_{Ag\_initial}$

$K_{sp\_AgCl} = X \cdot (C_{Cl} + X)$

Assuming X is small this simplifies to:

$$K_{sp\_AgCl} = X \cdot C_{Cl}$$

$$X := \frac{\left(K_{sp\_AgCl}\right)}{C_{Cl}}$$

$$X = -1.888 \times 10^{-7}\, mole \cdot liter^{-1} \qquad\qquad C_{Cl} = -9.535 \times 10^{-4}\, mol\, liter^{-1}$$

So the assumption was fine and now we can go on.

From this we can give the concentration of all the ions in solution:

$$C_{Ag} := X \qquad\qquad\qquad\qquad\qquad\qquad C_{Ag} = -1.888 \times 10^{-7}\, mol\, liter^{-1}$$

$$C_{Cl} = -9.535 \times 10^{-4}\, mol\, liter^{-1}$$
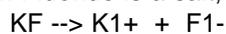
$$C_{NO3} := C_{AgNO3} \cdot \frac{75 \cdot mL}{1075 \cdot mL} \qquad\qquad C_{NO3} = 1.047 \times 10^{-3}\, mol\, liter^{-1}$$

$$C_{Li} := C_{LiCl} \cdot \frac{1000 \cdot mL}{1075 \cdot mL}$$

$$C_{Li} = 9.302 \times 10^{-5}\, mol\, liter^{-1}$$

d.100 L of 0.50 M calcium chloride is mixed with 5000.0 L of 0.10 M potassium fluoride.

Potassium Fluoride is a salt, it dissociates completely according to
KF --> K1+ + F1-

$$C_{KF} := 0.1 \cdot mole \cdot liter^{-1}$$

$$C_{F\_initial} := C_{KF} \cdot \frac{5000 \cdot liter}{5100 \cdot liter} \qquad\qquad C_{F\_initial} = 0.098\, mole \cdot liter^{-1}$$

Calcium Chloride is a salt, it dissociates completely according to
CaCl2 --> Ca2+ + 2 Cl1-

$$C_{CaCl2} := 0.015 \cdot mole \cdot liter^{-1}$$

$$C_{Ca\_initial} := C_{CaCl2} \frac{100 \cdot liter}{5100 \cdot liter} \qquad\qquad C_{Ca\_initial} = 2.941 \times 10^{-4}\, mole \cdot liter^{-1}$$

Since Calcium Fluoride is insoluble, it will form a precipitate. In this direction the reaction has a large value for K, so we'll assume it goes to completion. However, the concentrations are pretty small here. We better check Q to see if a precipitate will form:

$$Q := C_{Ca\_initial} \cdot \left(C_{F\_initial}\right)^2$$

$$Q = 2.827 \times 10^{-6} \left(\frac{mole}{liter}\right)^3$$

$$K_{sp\_CaF2} = 3.894 \times 10^{-11} \left(\frac{mole}{liter}\right)^3$$

Since Q is greater than K, a precipitate will form and calcium will be the limiting reagent.

$$V := 5100 \cdot liter$$

$$mole_{CaF2} := C_{Ca\_initial} \cdot V \qquad\qquad mole_{CaF2} = 1.5 \, mol$$

$$mass_{CaF2} := mole_{CaF2} \cdot MW_{CaF2} \qquad\qquad mass_{CaF2} = 117.111 \, gm$$

Next the equlibrium for the solubility of CaF2
        CaF2 <--> Ca2+ + 2 F1-

$$C_F := C_{F\_initial} - C_{Ca\_initial}$$

$$K_{sp\_CaF2} = X \cdot \left(C_F + 2 \cdot X\right)^2$$

Assuming X is small this simplifies to:

$$K_{sp\_CaF2} = X \cdot C_F^2$$

$$X := \frac{\left(K_{sp\_CaF2}\right)}{C_F^2}$$

$$X = 4.075 \times 10^{-9} \, mole \cdot liter^{-1} \qquad\qquad C_F = 0.098 \, mol \, liter^{-1}$$

So the assumption was fine and now we can go on.
From this we can give the concentration of all the ions in solution:

$$C_{Ca} := X \qquad\qquad\qquad\qquad\qquad C_{Ca} = 4.075 \times 10^{-9} \, mol \, liter^{-1}$$

$$C_F = 0.098 \, mol \, liter^{-1}$$

$$C_{Cl} := C_{CaCl2} \frac{100 \cdot liter}{5100 \cdot liter} \qquad\qquad C_{Cl} = 2.941 \times 10^{-4} \, mol \, liter^{-1}$$

$$C_K := C_{KF} \cdot \frac{5000 \cdot liter}{5100 \cdot liter} \qquad\qquad C_K = 0.098 \, mol \, liter^{-1}$$