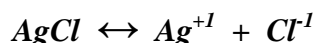1. Answer the following questions relating to the solubility of the chlorides of silver and lead.

   (a) At 10^0C, 8.9 x 10^{-5} g of AgCl (s) **will dissolve** in 100. mL of water.

   (i) Write the equation of the dissociation of AgCl (s) in water.

$$AgCl \leftrightarrow Ag^{+1} + Cl^{-1}$$

   (ii) Calculate the solubility, in mol L^{-1}, of AgCl (s) in water at 10^0C.

$$\frac{8.9 \times 10^{-5} \ g}{0.100 \ L} \times \frac{1 \ mol \ AgCl}{143.323 \ g} = 6.21 \times 10^{-6} \ M$$

   (iii) Calculate the value of the solubility-product constant, K_{sp}, for AgCl (s) at 10^0C.

$$K_{sp} = [Ag^{+1}][Cl^{-1}] = x^2 = (6.21 \ x \ 10^{-6})^2 = 3.86 \ x \ 10^{-11}$$

Some of you MAY want to build an ICE CHART to answer this question. Looking at the answer in part (i), the balanced equation shows a one to one to one ratio...

   (b) At 25^0C (a different temperature than in (a)), the value of K_{sp}, for $PbCl_2$ (s) is 1.6 x 10^{-5} and the value of K_{sp}, for AgCl (s) is 1.8 x 10^{-10}.

   (i) If 60.0 mL of 0.0400 M NaCl (aq) is added to 60.0 mL of 0.0300 M $Pb(NO_3)_2$ (aq), will a precipitate form? Assume that volumes are additive. Show calculations to support your answer. (K_{sp} vs. Q_{sp})

$$NaCl \ (aq) + Pb(NO_3)_2 \ (aq) \rightarrow NaNO_3 \ (aq) + PbCl_2 \ (s)$$
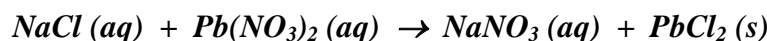
If the concentrations of the ions is sufficient enough to achieve an equilibrium value (Q_{sp}) greater than the K_{sp} value then a precipitate will form.

$$PbCl_2 \ (s) \leftrightarrow Pb^{+2} \ (aq) + 2 \ Cl^{-1} \ (aq)$$
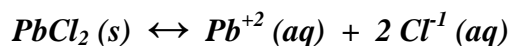
$$Cl^{-1} : \frac{60 \ mL}{} \times \frac{1 \ L}{1000 \ mL} \times \frac{0.0400 \ mol}{L} = 0.0024 \ mol \ ; \ [Cl^{-1}] = \frac{0.0024 \ mol}{0.120 \ L} = 0.020 \ M$$

$$Pb^{+2} : \frac{60 \ mL}{} \times \frac{1 \ L}{1000 \ mL} \times \frac{0.0300 \ mol}{L} = 0.0018 \ mol \ ; \ [Pb^{+2}] = \frac{0.0018 \ mol}{0.120 \ L} = 0.015 \ M$$

$$Q_{sp} = [Pb^{+2}][Cl^{-1}]^2 = (0.015 \ M)(0.020 \ M)^2 = 6.0 \ x \ 10^{-6}$$

$$Q_{sp} \ (6.0 \ x \ 10^{-6}) < K_{sp} \ (1.6 \ x \ 10^{-5})$$

A precipitate (ppt) WILL NOT form. The concentrations are NOT HIGH enough to reach the "solubility product" quotient. Until the concentrations are high enough, the solution will not become saturated (to form a ppt) A precipitate will NOT form until the concentrations are high enough for the Q_{sp} to exceed K_{sp} value.

(ii) Predict what SHOULD happen to the equilibrium value of $[Pb^{+2}]$ in 1.00 L of saturated $PbCl_2$ solution in which 0.250 mole of $NaCl\ (s)$ has been added. Assume no volume change occured.

$$PbCl_2 \leftrightarrow Pb^{+2} + 2\ Cl^{-1}$$

Since the system is AT equilibrium – adding more of a common ion (Cl^{-1}) to the solution will cause a shift away from the side of the addition. Since the chlorine ions will react with the lead ions to form the solid lead(II) chloride, the concentration of the lead ions will DECREASE.

(iii) If $0.100\ M$ $NaCl\ (aq)$ is added slowly to a beaker containing both $0.120\ M$ $AgNO_3\ (aq)$ and $0.150\ M$ $Pb(NO_3)_2\ (aq)$ at 25^0C, which will precipitate first, $AgCl\ (s)$ or $PbCl_2\ (s)$? EXPLAIN WHY

AgCl will precipitate out of solution before lead ions have a chance to precipitate out of solution because $PbCl_2$'s K_{sp} is much high than AgCl – the lower the K_{sp} the LESS the solution can hold as ions.

Lead(II) chloride will not precipitate out of solution UNTIL all of the silver ions have precipitated out (assuming the concentration of chlorine ions is much greater than the K_{sp} value will allow).

2. One prepares a solution by (attempting to) dissolving 22.50 grams manganese(II) iodate ($K_{sp} = 4.37 \times 10^{-7}$) in 2.88 L of water. Calculate the boiling point of this solution. Also, calculate how many grams of manganese(II) iodate crystals (solid) will remain (as a solid) at the bottom of the container once the system is at equilibrium.

$$K_{sp} = [\, Mn^{+2}\,]\,[\, IO_3^{-1}\,]^2$$

	$Mn(IO_3)_2$	$\leftrightarrow$	Mn^{+2}	$+$	$2\ IO_3^{-1}$
I	-		0		0
C	-		$+ x$		$+ 2x$
E	-		x		$2x$

$$4.37 \times 10^{-7} = [\, Mn^{+2}\,]\,[\, IO_3^{-1}\,]^2$$

$$4.37 \times 10^{-7} = [\, x\,]\,[\, 2x\,]^2 = 4x^3$$

$$x^3 = \frac{4.37 \times 10^{-7}}{4} = 1.09 \times 10^{-7}$$

$$x = \sqrt[3]{1.09 \times 10^{-7}} = 0.004781\ M = [\, Mn^{+2}\,]$$

Number of grams of Mn(IO$_3$)$_2$ Dissociated:

$$\frac{2.88\ L}{} \times \frac{4.781 \times 10^{-3}\ mol\ Mn}{1\ L} \times \frac{1\ mol\ Mn(IO_3)_2}{1\ mol\ Mn} \times \frac{404.758\ g}{1\ mol\ Mn(IO_3)_2} = 5.57\ g$$

Number of grams that remain in the solid state (un-dissociated):

22.50 g – 5.57 g = 16.93 g

$$\Delta T_b = i\ k_b\ molality$$

for water, molarity $=$ molality (be sure to show the calculations to "prove" this)

1 mL (water) = 1 gram (water)

$$\Delta T_b = (\,3\,)(\,0.51^0C \cdot kg \cdot mol^{-1}\,)(4.781 \times 10^{-3}\ mol \cdot kg^{-1}\,) = 0.00731^0C$$

B.P. for water = 100 ^0C + 0.00731 ^0C = 100.00731 ^0C (insignificant change)