Today we will practice the skill of THINKING LIKE A CHEMIST while considering the concept of solubility. Platinum stars will be on the line.
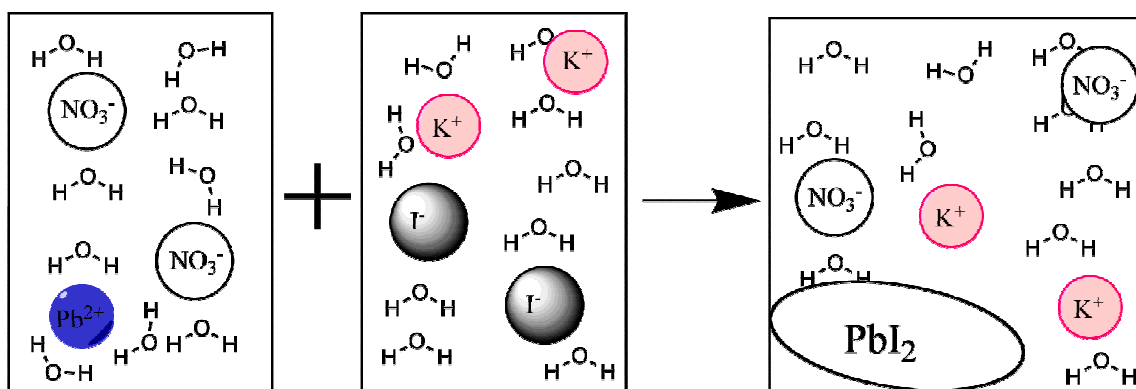
Consider the following demonstration, and describe macroscopically.

1. A solution of lead II nitrate is mixed with a solution of potassium iodide.
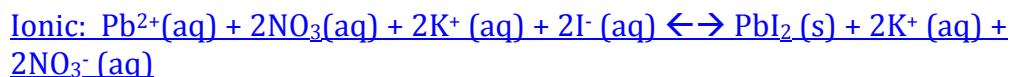
   Two clear solutions are mixed together and a <u>yellow</u> precipitate forms.

2. Describe the demonstration from a microscopic perspective using words and a picture.

   Lead ions and nitrate ions are mixed with potassium ions and iodide ions. As they mix, the concentration of lead ions and iodide ions decreases and these two form a precipitate, while the remaining spectator ions remain in solution.


3. Use chemical equations to model the changes.

   <u>Formula:</u> $Pb(NO_3)_2$ (aq) + 2KI (aq) $\leftarrow\rightarrow$ PbI_2 (s) + 2KNO_3^-(aq)

   <u>Ionic:</u> Pb^{2+}(aq) + 2NO_3(aq) + 2K^+ (aq) + 2I^- (aq) $\leftarrow\rightarrow$ PbI_2 (s) + 2K^+ (aq) + 2NO_3^- (aq)

   <u>Net:</u> Pb^{2+}(aq) + 2I^- (aq) $\leftarrow\rightarrow$ PbI_2 (s)

Be prepared to explain if called upon in class.

THINK ABOUT THIS INSOLUBLE SALTS FROM DIFFERENT PERSPECTIVES!

4. Consider that the Ksp of PbI$_2$ is 1.4 x 10^{-8}. Predict the concentration of Pb^{2+} ions in solution in a saturated solution of PbI$_2$. Predict the concentration of the I^{-1} ions in the saturated solution.

PbI$_2$ (s) ←→ Pb^{2+} (aq) + 2I$^-$ (aq)
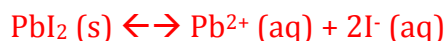
For every one Pb^{2+}, there are two I$^-$ ions, so when we replace Pb^{2+} with "x", the concentration of I$^-$ will be "2x"

K$_{sp}$ = 1.4 x 10^{-8} = [Pb^{2+}][I$^-$]^2 = 1.4 x 10^{-8}

[x][2x]^2 = 1.4 x 10^{-8}
4x^3 = 1.4 x 10^{-8}
x^3 = 3.5 x 10^{-9}
x = 1.52 x 10^{-3}

x = [Pb^{2+}] = 1.52 x 10^{-3} M
2x = [I$^-$] = 3.04 x 10^{-3} M

The solubility product is the value of the ion product of the molar concentrations of all the ions in the formula unit at saturation (or equilibrium). Consider adding more iodide ion to this saturated solution of PbI$_2$ in the form of adding some solid, soluble KI salt to the solution.

5. Will the concentration of I$^-$ in the solution change? Will the concentration of the Pb^{2+} ion change? Why or why not?

[I$^-$] will increase due to this added iodide from KI
Since K$_{sp}$ must remain constant and the concentration of I$^-$ increases, [Pb^{2+}] will decrease, because not as much of the original solid needs to be dissolved in order to reach equilibrium

You must consider that the solubility product for a given salt is the product of the molar concentrations of the ions raised to the power of the coefficient. If you add more of one ion, then the concentration of that added ion would increase. Because the solubility product value is a constant, if one factor in that expression increases then the other factor will need to decrease in order to keep the product (the value of K$_{sp}$) constant! This effect is called the **COMMON ION EFFECT**.

6. Assume you add solid PbI$_2$ to a 0.5 molar aqueous solution of KI. What will be the concentration of Pb^{2+} in the solution? What will be the molar solubility of the salt?

R PbI$_2$ (s) ←→ Pb^{2+} (aq) + 2I$^-$ (aq)

I   0   0   0.5 M
C   -x   +x   +2x
E   -x   +x   0.5 + 2x

$K_{sp} = 1.4 \times 10^{-8} = [Pb^{2+}] [I^-]^2$
$1.4 \times 10^{-8} = [x] [0.5 + 2x]^2$
Since the concentration of x will be significantly less (more than 100x less) than 0.5 M, we can ignore the "2x" factor
However, we do not get rid of the "x" factor, since the concentration of lead is NOT zero, this factor must be included
$1.4 \times 10^{-8} = [x] [0.5]^2$
$1.4 \times 10^{-8} = 0.25x$
$x = 5.6 \times 10^{-8}$ M

The concentration of Pb^{2+} will be 5.6×10^{-8} M
Since Pb^{2+} is in a 1:1 ratio with the solid, PbI_2, the molar solubility (or the total moles of PbI_2 that will dissolve) will also be 5.6×10^{-8} M

7. What will be the concentration of the I^- anion?
    From the RICE table above,
    $[I^-] = 0.5 \pm 2x$
    $[I^-] = 0.5 \pm 2(5.6 \times 10^{-8})$
    $[I^-] = 0.5 \pm 1.12 \times 10^{-7} = \sim 0.5$ M

          Be prepared to explain your result if called upon.

When we demonstrated the precipitation reactions in class, we mixed together solutions containing high concentrations of the soluble salts to insure we could have a concentration of the insoluble salt product high enough to see a nice precipitate. However, if the concentrations of the ions are low enough, you might not see a precipitate forming! One can predict the formation of a precipitate based on calculating the ION product. The ION product is of the similar form as the K_{sp} expression, except the ION product is the product of the concentrations at any point in time, not just at the equilibrium condition. The ION product is also referred to as the **REACTION QUOTIENT** and for PbI_2 will be expressed as: $Q_{sp} = [Pb^{2+}][I^-]^2$. If the value of Q_{sp} is greater than K, then a precipitate will form. If the value is less that K, then a precipitate will NOT form!

8. Suppose you mix 1 liter each of 0.2 M solutions of $Pb(NO_3)_2$(aq) and KI(aq). Follow these steps to determine if a precipitate will form! First determine the molar concentrations of the two salts in the mixture.

    (0.2 mol/L $Pb(NO_3)_2$)(1L $Pb(NO_3)_2$) = 0.2 moles $Pb(NO_3)_2$
    (0.2 mol/L KI)(1L KI) = 0.2 moles KI
    Total Volume = 1L + 1L = 2L

New Concentrations:
0.2 mol $Pb(NO_3)_2 \div 2L = 0.1$ M $Pb(NO_3)_2$
0.2 mol $KI \div 2L = 0.1$ M KI

9. Calculate the value of Q_{sp} for the PbI_2.
Net Ionic Equation:
Pb^{2+} (aq) + $2I^-$(aq) → PbI_2 (s)
$Q_{sp}= [Pb^{2+}][I^-]^2 = [.1M][.1M]^2 = 0.001$

Because Q is based on the concentration given, that we have right now, we do not consider the x to 2x factors that show up in equilibrium situations (K_{sp}).

The concentration of I^- is the same as the concentration of KI from #8 (1:1 ratio)
The concentration of Pb^{2+} is the same as the concentration of $Pb(NO_3)_2$ from #8 (1:1 ratio).

10. Determine if the value of Q_{sp} is greater or less than K_{sp}. Predict if a precipitate will form under these conditions.
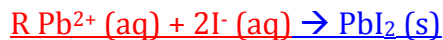
Yes, a precipitate will form:
$Q_{sp} > K_{sp}$ because $0.001 > 1.4 \times 10^{-8}$

11. If a precipitate did form, predict the mass of the precipitate that will fall out of solution.

We will first assume that this reaction goes to completion, which creates a limiting reagent problem.

All units in this RICE table are worked in MOLES

(0.1 mol/L) x 2L = 0.2 moles

R Pb^{2+} (aq) + $2I^-$ (aq) → PbI_2 (s)
I   0.2   0.2   0
C   -0.1   -0.2   +0.1
E   0.1 moles   0   0.1 moles

0.1 moles PbI_2 x (461 grams/mol) = 46.1 grams

There are 0.1 moles of Pb^{2+} left in the 2 L of solution, giving $[Pb^{2+}] = 0.05$ M

To determine how much I^- would redissolve at equilibrium:
$K_{sp} = 1.4 \times 10^{-8} = [Pb^{2+}][I^-]^2$
$1.4 \times 10^{-8} = (0.05)(x)^2$
$x = 5.3 \times 10^{-4}$ M

12. Assume that you cannot change the concentration of the I^- ion in this solution. However, you can change the Pb^{2+} ion concentration in your solution. To what concentration would you have to lower the Pb^{2+} ion concentration to prevent the precipitation from occurring?

$[I^-]$ = 0.1M and $[Pb^{2+}]$ = x
K_{sp} = 1.4 x 10^{-8} = [x] $[0.1]^2$
[x] = $[Pb^{2+}]$ =1.4 x 10^{-6} M
So there would have to be an enormous decrease in the lead concentration to avoid the formation of a precipitate.

Be prepared to share your answers with the class if called upon.