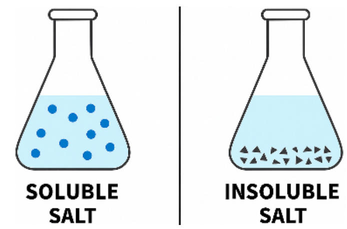
Solution Equilibria **Name**________________________________

The Solubility Product Constant (K_{sp})

Read from **Lesson 1: Solubility of Salts** in the **Chemistry Tutorial** Section, **Chapter 16** of **The Physics Classroom**: Part a: The Solubility Product Constant, K_{sp} Part b: Solubility and K_{sp}

Soluble vs. Insoluble Salts



- **Salts** are ionic compounds formed from cations and anions.
- **Soluble salts** dissolve readily in water; they dissociate completely.
- **Insoluble salts** dissolve only slightly: they reach a **saturation point** where no more ions can dissolve.
- Even "insoluble" salts dissolve a little—these are called **sparingly soluble** or **slightly soluble**.
- Use **solubility rules** to predict whether a salt will dissolve.

Dissociation of Salts

- Dissociation: the process of a salt breaking into ions when it dissolves.
- Na_3PO_4 (s) $\rightarrow$ 3 Na^+(aq) + PO_4^{3-} (aq) *(This is a **soluble** salt: 100% dissociated)*
- $PbCl_2$(s) $\rightleftarrows$ Pb^{2+}(aq) + 2 Cl^-(aq) *(This is an **insoluble** salt: not 100% dissociated)*
    - Adding excess $PbCl_2$ to water results in a saturated solution with undissolved solid remaining.

The Solubility Product Constant (K_{sp})

- **K_{sp}** quantifies the extent to which an insoluble salt dissolves.
- Larger **K_{sp}** → greater solubility.
- **K_{sp}** for lead (II) chloride: $PbCl_2$(s) $\rightleftarrows$ Pb^{2+}(aq) + 2 Cl^-(aq) is **$K_{sp} = [Pb^{2+}] * [Cl^-]^2$** (***remember, solids are omitted from equilibrium expressions)

Calculating Ion Concentrations from K_{sp}

1. Write the balanced dissociation equation.
2. Write the K_{sp} expression.
3. Set up an ICE table (Initial, Change, Equilibrium).
4. Substitute values and solve for x (amount dissolved).
5. Use algebra to find ion concentrations.

Example: The K_{sp} for $PbCl_2$ is 1.7×10^{-5} at 25°C. Determine the ion concentrations at equilibrium.

	$PbCl_2$ (s)	$\rightleftarrows$	**Pb^{2+}(aq)**	**+**	**2 Cl^-(aq)**
Initial	Does not matter		**0**		**0**
Change	**-x**		**+ x**		**+ 2x**
Equilibrium	Does not matter		**x**		**2x**

$K_{sp} = 1.7 \times 10^{-5} = [Pb^{2+}] * [Cl^-]^2 = (x) * (2x)^2$ $K_{sp} = 1.7 \times 10^{-5} = 4x^3$

x = 0.016, so **$[Pb^{2+}]$ = 0.016 M,** 2x = 0.032 so **$[Cl^-]$ = 0.032 M**

Molar Solubility

- **Definition**: Moles of salt that dissolve per liter of water. (Can be found by solving for **x** in the ICE table.)
- Reverse Calculations: find K_{sp} from molar solubility by using the same steps as above, but in reverse.

Solubility (g/L)

- Convert molar solubility to grams/L using molar mass.

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Questions

1. Use the solubility rules to determine whether each of the following salts is **soluble** or **insoluble** in water. For each compound: write its correct **chemical formula** and then label it as **soluble** or **insoluble**.

Compound Name	Chemical Formula	Solubility (Soluble / Insoluble)
a. Magnesium bromide		
b. Iron(II) carbonate		
c. Chromium(III) hydroxide		
d. Ammonium sulfate		
e. Silver iodide		
f. Tin(IV) sulfite		
g. Barium acetate		
h. Copper(II) sulfide		
i. Calcium phosphate		
j. Potassium hydroxide		

2. From your list in Question 1, choose the compounds you identified as **insoluble** in water. For each one, complete the table below:

Compound Formula	Dissociation Equation	K_{sp} Expression
Example: $BaSO_4$	$BaSO_4\ (s) \rightleftharpoons Ba^{2+}\ (aq) + SO_4^{2-}\ (aq)$	$K_{sp} = [Ba^{2+}][SO_4^{2-}]$

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3. Sophie Soluble and Aaron Agin are debating the solubility of lead(II) compounds. Sophie claims lead(II) carbonate is less soluble than lead(II) chloride due to its smaller K_{sp}. Aaron disagrees, arguing that lead(II) carbonate is more soluble because it dissociates into fewer ions. How can Sophie clarify that K_{sp} —not ion count—determines solubility?

4. Use tools such as a dissociation equation, an ICE table, a K_{sp} expression, and good algebra skills to calculate ion concentrations in the following solutions.

   a. 0.50 M potassium carbonate

   b. 0.33 M aluminum bromide

   c. a saturated solution of silver acetate at 25°C. ($K_{sp} = 1.9 \times 10^{-3}$)

   d. a saturated solution of calcium hydroxide at 25°C. ($K_{sp} = 5.0 \times 10^{-6}$)

   e. a saturated solution of magnesium phosphate at 25°C. ($K_{sp} = 1.0 \times 10^{-24}$)

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5. Use tools such as a dissociation equation, an ICE table, a K_{sp} expression, and good algebra skills to calculate the K_{sp} for the following compounds:
   a. A saturated solution of calcium fluoride has the following ion concentrations at equilibrium:
      $[Ca^{2+}] = 2.1 \times 10^{-4}$ M and $[F^{-}] = 4.2 \times 10^{-4}$ M

   b. In a saturated solution of lead(II) hydroxide, the $[OH^{-}]$ is measured to be 3.0×10^{-7} M.

   c. The molar solubility of magnesium hydroxide is 1.12×10^{-4} mol/L at 25°C.

6. Calculate the solubility of the following solutions.
   a. Copper(I) chloride in mol/L ($K_{sp} = 1.7 \times 10^{-7}$)

   b. Magnesium fluoride in g/L ($K_{sp} = 5.2 \times 10^{-11}$)

   c. Calcium iodate in g/L ($K_{sp} = 6.5 \times 10^{-6}$)