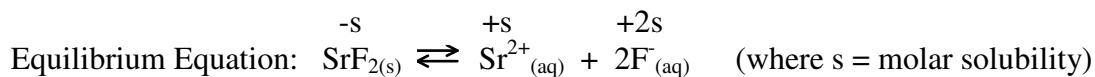


**Chemistry 12**  
**Tutorial 10—Ksp Calculations**  
**Solutions**

**1. Calculate the solubility of SrF<sub>2</sub> in moles/Litre in water.**



$$K_{sp} = [\text{Sr}^{2+}] [\text{F}^{-}]^2$$

$$K_{sp} = s (2s)^2$$

$$K_{sp} = s \times 4s^2$$

$$K_{sp} = 4s^3$$

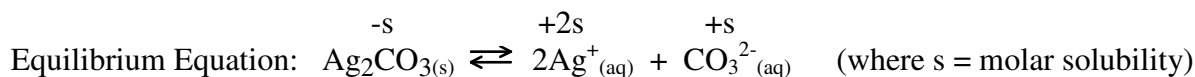
$$s^3 = \frac{K_{sp}}{4}$$

$$s = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4.3 \times 10^{-9}}{4}} = 1.0244 \times 10^{-3} \text{ M}$$

Because the K<sub>sp</sub> was 2 SD's, the answer will be: Molar Solubility = 1.0 x 10<sup>-3</sup> M

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**2. Calculate the solubility of Ag<sub>2</sub>CO<sub>3</sub> in grams/Litre.**



$$K_{sp} = [\text{Ag}^{+}]^2 [\text{CO}_3^{2-}]$$

$$K_{sp} = (2s)^2 \times s$$

$$K_{sp} = 4s^2 \times s$$

$$K_{sp} = 4s^3$$

$$s^3 = \frac{K_{sp}}{4}$$

$$s = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{8.5 \times 10^{-12}}{4}} = 1.286 \times 10^{-4} \text{ M}$$

$$\text{Molar Solubility} = 1.286 \times 10^{-4} \text{ M}$$

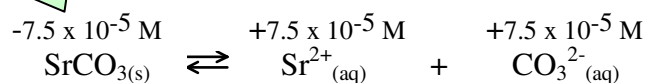
To get solubility in g/L:

$$\frac{1.286 \times 10^{-4} \text{ mol}}{\text{L}} \times \frac{275.8 \text{ g Ag}_2\text{CO}_3}{1 \text{ mol Ag}_2\text{CO}_3} = 3.5 \times 10^{-2} \text{ g/L (or 0.035 g/L)}$$

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3. At a certain temperature, the *solubility* of  $\text{SrCO}_3$  is  $7.5 \times 10^{-5} \text{ M}$ . Calculate the *K<sub>sp</sub>* for  $\text{SrCO}_3$ .

Molar Solubility



$$\begin{aligned} K_{sp} &= [\text{Sr}^{2+}] [\text{CO}_3^{2-}] \\ &= (7.5 \times 10^{-5})^2 \end{aligned}$$

$$\mathbf{K_{sp} = 5.6 \times 10^{-9}}$$

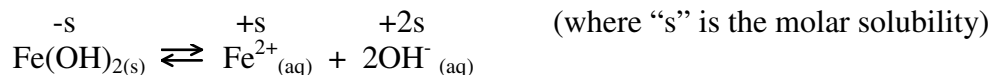
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### Answers to Self Test on Tutorial 10

- The  $K_{sp}$  is just a  $K_{eq}$  for an ionic substance dissolving in water.
- Give the *Net-Ionic Equation* and the *K<sub>sp</sub> expression* for each of the following dissolving in water. (*The first one is done as an example.*)

Substance	Net-Ionic Equation	K <sub>sp</sub> Expression
$\text{Ag}_2\text{SO}_4(s)$	$\text{Ag}_2\text{SO}_4(s) \rightleftharpoons 2\text{Ag}^+_{(aq)} + \text{SO}_4^{2-}_{(aq)}$	$K_{sp} = [\text{Ag}^+]^2 [\text{SO}_4^{2-}]$
$\text{CaCO}_3(s)$	$\text{CaCO}_3(s) \rightleftharpoons \text{Ca}^{2+}_{(aq)} + \text{CO}_3^{2-}_{(aq)}$	$K_{sp} = [\text{Ca}^{2+}] [\text{CO}_3^{2-}]$
$\text{Ca}_3(\text{PO}_4)_2(s)$	$\text{Ca}_3(\text{PO}_4)_2(s) \rightleftharpoons 3\text{Ca}^{2+}_{(aq)} + 2\text{PO}_4^{3-}_{(aq)}$	$K_{sp} = [\text{Ca}^{2+}]^3 [\text{PO}_4^{3-}]^2$
$\text{AgClO}_3(s)$	$\text{AgClO}_3(s) \rightleftharpoons \text{Ag}^+_{(aq)} + \text{ClO}_3^-_{(aq)}$	$K_{sp} = [\text{Ag}^+] [\text{ClO}_3^-]$

3. a) Calculate the **molar solubility** (solubility in moles/Litre) of  $\text{Fe}(\text{OH})_2$  in water.



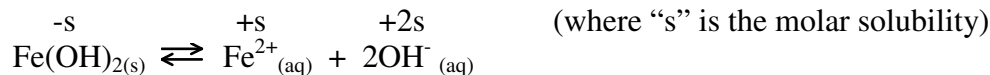
$$K_{sp} = [\text{Fe}^{2+}] [\text{OH}^{-}]^2$$

$$K_{sp} = s (2s)^2$$

$$K_{sp} = 4s^3$$

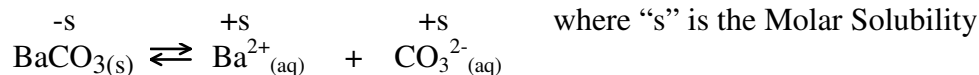
$$s = \sqrt[3]{\frac{K_{sp}}{4}} = \sqrt[3]{\frac{4.9 \times 10^{-17}}{4}} = 2.3 \times 10^{-6} \text{ M}$$

- b) What is the  $[\text{OH}^{-}]$  in a saturated solution of  $\text{Fe}(\text{OH})_2$  ?



$$\text{Since } [\text{OH}^{-}] = 2s ; \text{ then } [\text{OH}^{-}] = 2(2.3052 \times 10^{-6}) = 4.6 \times 10^{-6} \text{ M}$$

4. Calculate the **solubility** of  $\text{BaCO}_3$  in grams per Litre.



$$K_{sp} = [\text{Ba}^{2+}] [\text{CO}_3^{2-}]$$

$$K_{sp} = s^2$$

$$s = \sqrt{K_{sp}} = \sqrt{2.6 \times 10^{-9}} = 5.099 \times 10^{-5} \text{ M}$$

Solubility in g/L:

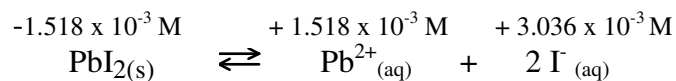
$$5.099 \times 10^{-5} \text{ mol/L} \times 197.3 \text{ g/mol} = 1.0 \times 10^{-2} \text{ g/L}$$

5. The solubility of  $\text{PbI}_2$  at a certain temperature is 0.70 grams per Litre.

- a) Calculate the solubility in moles/Litre

$$\frac{0.70 \text{ g}}{\text{L}} \times \frac{1 \text{ mol}}{461.0 \text{ g}} = 1.5 \times 10^{-3} \text{ mol/L (unrounded = } 1.518 \times 10^{-3} \text{ M)}$$

b) Calculate the *value of K<sub>sp</sub>* for PbI<sub>2</sub> at this temperature



$$K_{sp} = [\text{Pb}^{2+}] \times [\text{I}^{-}]^2$$

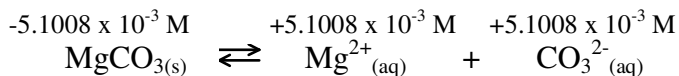
$$K_{sp} = 1.518 \times 10^{-3} (3.036 \times 10^{-3})^2 = \mathbf{1.4 \times 10^{-8}}$$

- 6 It is found that 0.043 grams of MgCO<sub>3</sub> is all that can dissolve in 100.0 mL of solution at a certain temperature. From this information, calculate the *K<sub>sp</sub>* for MgCO<sub>3</sub> at this temperature.

**Plan:** g → mol → M (molar solubility) → K<sub>sp</sub>

$$0.043 \text{ g MgCO}_3 \times \frac{1 \text{ mol MgCO}_3}{84.3 \text{ g MgCO}_3} = 5.1008 \times 10^{-4} \text{ mol MgCO}_3$$

$$[\text{MgCO}_3] = \frac{5.1008 \times 10^{-4} \text{ mol}}{0.1000 \text{ L}} = 5.1008 \times 10^{-3} \text{ M (This is the molar solubility)}$$



$$K_{sp} = [\text{Mg}^{2+}] [\text{CO}_3^{2-}]$$

$$= (5.1008 \times 10^{-3})^2$$

$$\mathbf{K_{sp} = 2.6 \times 10^{-5}}$$

7. Two separate experiments were done with combinations of  $\text{Cu}^{2+}$  and  $\text{IO}_3^-$  ions. Use the information given to fill in the missing value.

The Net-Ionic Equation for equilibrium is:  $\text{Cu}(\text{IO}_3)_2(\text{s}) \rightleftharpoons \text{Cu}^{2+}(\text{aq}) + 2\text{IO}_3^-(\text{aq})$

Experiment #	$[\text{Cu}^{2+}]$	$[\text{IO}_3^-]$
1	0.00327 M	0.00654 M
2	0.00240 M	?

Use the results from Experiment #1 to calculate the  $K_{\text{sp}}$  for  $\text{Cu}(\text{IO}_3)_2$ :

$$\begin{aligned} K_{\text{sp}} &= [\text{Cu}^{2+}] [\text{IO}_3^-]^2 \\ &= 0.00327 \times (0.00654)^2 \\ &= 1.399 \times 10^{-7} \end{aligned}$$

Use the  $K_{\text{sp}}$  and  $[\text{Cu}^{2+}]$  to calculate  $[\text{IO}_3^-]$  in Experiment #2:

$$\begin{aligned} K_{\text{sp}} &= [\text{Cu}^{2+}] [\text{IO}_3^-]^2 \\ [\text{IO}_3^-]^2 &= \frac{K_{\text{sp}}}{[\text{Cu}^{2+}]} \\ [\text{IO}_3^-] &= \sqrt{\frac{K_{\text{sp}}}{[\text{Cu}^{2+}]}} = \sqrt{\frac{1.399 \times 10^{-7}}{0.00240}} \end{aligned}$$

$$[\text{IO}_3^-] = 7.63 \times 10^{-3} \text{ M}$$