

**K<sub>sp</sub> Problems Worksheet Solutions**

1. a)  $\text{CaSO}_4 \rightleftharpoons \text{Ca}^{+2} + \text{SO}_4^{-2}$  If the solubility is  $5.0 \times 10^{-3}$  mol/L then we will let 'x' be equal to this amount.

$$\begin{aligned} K_{sp} &= [\text{Ca}^{+2}][\text{SO}_4^{-2}] = [x][x] = 5.0 \times 10^{-3} \\ \text{Therefore } K_{sp} &= (5.0 \times 10^{-3})^2 \\ K_{sp} &= 2.5 \times 10^{-5} \end{aligned}$$

- b)  $\text{MgF}_2 \rightleftharpoons \text{Mg}^{+2} + 2 \text{F}^{-1}$  If the solubility is  $2.7 \times 10^{-3}$  mol/L then we will let 'x' be equal to this amount.

$$\begin{aligned} K_{sp} &= [\text{Mg}^{+2}][\text{F}^{-1}]^2 = [x][2x]^2 = 2.7 \times 10^{-3} \\ K_{sp} &= 4x^3 \\ \text{Therefore } x &= 4(2.7 \times 10^{-3})^3 \\ K_{sp} &= 7.8 \times 10^{-8} \end{aligned}$$

- c) There is 1.02 grams dissolved in 100 mL. this must first be converted to moles/L.  
 moles = grams/molecular mass

$$\text{moles} = 1.02 \text{ g} / 166.89 \text{ g/moles} = 0.006 \text{ moles}$$

$$[\text{AgC}_2\text{H}_3\text{O}_2] = \text{moles/L} = 0.006 \text{ moles} / 0.1 \text{ L} = 0.06 \text{ moles/L}$$

- $\text{AgC}_2\text{H}_3\text{O}_2 \rightleftharpoons \text{Ag}^{+1} + \text{C}_2\text{H}_3\text{O}_2^{-1}$  If the solubility is 0.06 mol/L then we will let 'x' be equal to this amount.

$$\begin{aligned} K_{sp} &= [\text{Ag}^{+1}][\text{C}_2\text{H}_3\text{O}_2^{-1}] = [x][x] = 0.06 \\ \text{Therefore } K_{sp} &= (0.06)^2 \\ K_{sp} &= 3.6 \times 10^{-3} \end{aligned}$$

- d) Convert 12.2 mg in 100 mL of water into moles/L

$$12.2 \text{ mg} / 1000 \text{ mg/gram} = 0.0122 \text{ grams}$$

$$\begin{aligned} \text{moles} &= \text{grams/molecules mass} = 0.0122 \text{ grams} / 125.62 \text{ grams/mole} \\ &= 9.7 \times 10^{-5} \text{ moles} \end{aligned}$$

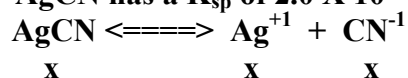
$$[\text{SrF}_2] = \text{moles/L} = 9.7 \times 10^{-5} / 0.1 \text{ L} = 9.7 \times 10^{-4} \text{ moles/L}$$

- $\text{SrF}_2 \rightleftharpoons \text{Sr}^{+2} + 2 \text{F}^{-1}$  If the solubility is  $9.7 \times 10^{-4}$  mol/L then we will let 'x' be equal to this amount.

$$\begin{aligned} K_{sp} &= [\text{Mg}^{+2}][\text{F}^{-1}]^2 = [x][2x]^2 = 9.7 \times 10^{-4} \\ K_{sp} &= 4x^3 \\ \text{Therefore } x &= 4(9.7 \times 10^{-4})^3 \end{aligned}$$

$$K_{sp} = 3.0 \times 10^{-9}$$

2. i) AgCN has a  $K_{sp}$  of  $2.0 \times 10^{-12}$



$$K_{sp} = [\text{Ag}^{+1}][\text{CN}^{-1}]$$

$$2.0 \times 10^{-12} = (x)(x)$$

$$x = (2.0 \times 10^{-12})^{1/2}$$

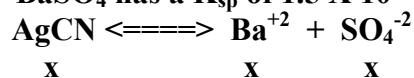
$$x = 1.4 \times 10^{-9} \text{ moles/L (This is the solubility of the silver cyanide)}$$

$$\text{molecular mass of Ag}^{+1} = 107.87 \text{ g/mole}$$

$$\text{Therefore } 107.87 \text{ g/mole} \times 1.4 \times 10^{-9} \text{ mole/L} = 1.51 \times 10^{-7} \text{ g/L}$$

$$= 1.51 \times 10^{-7} \text{ mg/mL}$$

- ii) BaSO<sub>4</sub> has a  $K_{sp}$  of  $1.5 \times 10^{-9}$



$$K_{sp} = [\text{Ba}^{+2}][\text{SO}_4^{-2}]$$

$$1.5 \times 10^{-9} = (x)(x)$$

$$x = (1.5 \times 10^{-9})^{1/2}$$

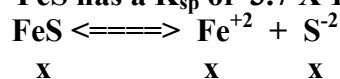
$$x = 3.9 \times 10^{-5} \text{ moles/L (This is the solubility of the barium sulphate)}$$

$$\text{molecular mass of Ba}^{+2} = 137.33 \text{ g/mole}$$

$$\text{Therefore } 137.33 \text{ g/mole} \times 3.9 \times 10^{-5} \text{ mole/L} = 5.36 \times 10^{-3} \text{ g/L}$$

$$= 5.36 \times 10^{-3} \text{ mg/mL}$$

- iii) FeS has a  $K_{sp}$  of  $3.7 \times 10^{-19}$



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

$$3.7 \times 10^{-19} = (x)(x)$$

$$x = (3.7 \times 10^{-19})^{1/2}$$

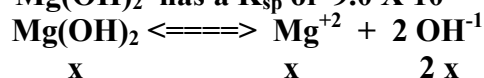
$$x = 6.1 \times 10^{-10} \text{ moles/L (This is the solubility of the iron(II) sulphate)}$$

$$\text{molecular mass of Fe}^{+2} = 55.85 \text{ g/mole}$$

$$\text{Therefore } 55.85 \text{ g/mole} \times 6.1 \times 10^{-10} \text{ mole/L} = 3.41 \times 10^{-8} \text{ g/L}$$

$$= 3.41 \times 10^{-8} \text{ mg/mL}$$

- iv) Mg(OH)<sub>2</sub> has a  $K_{sp}$  of  $9.0 \times 10^{-12}$



$$K_{sp} = [\text{Mg}^{+2}][\text{OH}^{-1}]^2$$

$$9.0 \times 10^{-12} = (x)(2x)^2$$

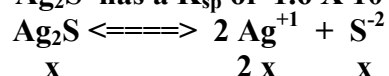
$$x = \left( \frac{9.0 \times 10^{-12}}{4} \right)^{1/3}$$

$$x = 1.3 \times 10^{-4} \text{ moles/L (This is the solubility of the magnesium hydroxide)}$$

$$\text{molecular mass of Mg}^{+2} = 24.31 \text{ g/mole}$$

$$\begin{aligned} \text{Therefore } 24.31 \text{ g/mole} \times 1.3 \times 10^{-4} \text{ mole/L} &= 3.18 \times 10^{-3} \text{ g/L} \\ &= 3.18 \times 10^{-3} \text{ mg/mL} \end{aligned}$$

v)  $\text{Ag}_2\text{S}$  has a  $K_{sp}$  of  $1.6 \times 10^{-49}$



$$K_{sp} = [\text{Ag}^{+1}]^2[\text{S}^{-2}]$$

$$1.6 \times 10^{-49} = (2x)^2(x)$$

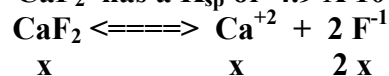
$$x = \left( \frac{1.6 \times 10^{-49}}{4} \right)^{1/3}$$

$$x = 3.4 \times 10^{-17} \text{ moles/L (This is the solubility of the silver sulphide)}$$

$$\text{molecular mass of 2 moles of Ag}^{+1} = 215.74 \text{ g/mole}$$

$$\begin{aligned} \text{Therefore } 215.74 \text{ g/mole} \times 3.4 \times 10^{-17} \text{ mole/L} &= 7.33 \times 10^{-15} \text{ g/L} \\ &= 7.33 \times 10^{-15} \text{ mg/mL} \end{aligned}$$

vi)  $\text{CaF}_2$  has a  $K_{sp}$  of  $4.9 \times 10^{-11}$



$$K_{sp} = [\text{Ca}^{+2}][\text{F}^{-1}]^2$$

$$4.9 \times 10^{-11} = (x)(2x)^2$$

$$x = \left( \frac{4.9 \times 10^{-11}}{4} \right)^{1/3}$$

$$x = 2.31 \times 10^{-4} \text{ moles/L (This is the solubility of the calcium fluoride)}$$

$$\text{molecular mass of Ca}^{+2} = 40.08 \text{ g/mole}$$

$$\begin{aligned} \text{Therefore } 40.08 \text{ g/mole} \times 2.31 \times 10^{-4} \text{ mole/L} &= 9.26 \times 10^{-3} \text{ g/L} \\ &= 9.26 \times 10^{-3} \text{ mg/mL} \end{aligned}$$

3.

	$K_{sp}$		Solubility
PbS	$8.4 \times 10^{-28}$	$K_{sp} = x^2$	$2.9 \times 10^{-14} \text{ mole/L}$
PbSO <sub>4</sub>	$1.8 \times 10^{-8}$	$K_{sp} = x^2$	$1.34 \times 10^{-4} \text{ mole/L}$
Pb(IO <sub>3</sub> ) <sub>2</sub>	$2.6 \times 10^{-13}$	$K_{sp} = 4x^3$	$4.02 \times 10^{-5} \text{ mole/L}$

a) The lead(II) sulphate is the most soluble.

b) The solubility of the lead(II) sulphate is  $1.34 \times 10^{-4} \text{ moles/L}$

c) Molecular mass of PbSO<sub>4</sub> is 303.27 g/mole

$$\begin{aligned} m &= n \times M \\ &= 1.34 \times 10^{-4} \text{ moles/L} \times 303.27 \text{ grams/mole} \\ &= 0.041 \text{ grams/L} \end{aligned}$$

d) Use the common ion effect and add something with sulphate ion in it that dissociates easily and more readily. i.e., H<sub>2</sub>SO<sub>4</sub>

e) The PbS concentration is  $2.9 \times 10^{-14}$  moles/L

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4. a)  $\text{Cu(OH)}_2 \rightleftharpoons \text{Cu}^{+2} + 2 \text{OH}^{-1}$

$$\begin{aligned} K_{sp} &= [\text{Cu}^{+2}][\text{OH}^{-1}]^2 \\ \text{therefore } [\text{Cu}^{+2}] &= \frac{K_{sp}}{[\text{OH}^{-1}]^2} = \frac{1.6 \times 10^{-9}}{(1.0 \times 10^{-4})^2} = \frac{1.6 \times 10^{-9}}{1.0 \times 10^{-8}} = 0.16 \text{ mole/L} \end{aligned}$$

$$\begin{aligned} \text{grams} &= \text{moles} * \text{molecular mass} \\ &= 0.16 \text{ moles / L} * 63.55 \text{ g/mole} \\ &= 10.168 \text{ g/L} \\ &= 10168 \text{ mg/L} \end{aligned}$$


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b)  $\text{Fe(OH)}_3 \rightleftharpoons \text{Fe}^{+1} + 3 \text{OH}^{-1}$

$$\begin{aligned} K_{sp} &= [\text{Fe}^{+1}][\text{OH}^{-1}]^3 \\ \text{therefore } [\text{Fe}^{+2}] &= \frac{K_{sp}}{[\text{OH}^{-1}]^3} = \frac{6.0 \times 10^{-38}}{(1.0 \times 10^{-4})^3} = \frac{6.0 \times 10^{-38}}{1.0 \times 10^{-12}} \\ &= 6.0 \times 10^{-26} \text{ mole/L} \end{aligned}$$

$$\begin{aligned} \text{grams} &= \text{moles} * \text{molecular mass} \\ &= 6.0 \times 10^{-26} \text{ mole/L} * 55.85 \text{ g/mole} \\ &= 3.35 \times 10^{-24} \text{ g/L} \\ &= 3.3 \times 10^{-21} \text{ mg/L} \end{aligned}$$

c)  $\text{Mg(OH)}_2 \rightleftharpoons \text{Mg}^{+2} + 2 \text{OH}^{-1}$

$$\begin{aligned} K_{sp} &= [\text{Mg}^{+2}][\text{OH}^{-1}]^2 \\ \text{therefore } [\text{Mg}^{+2}] &= \frac{K_{sp}}{[\text{OH}^{-1}]^2} = \frac{6.0 \times 10^{-12}}{(1.0 \times 10^{-4})^2} = \frac{6.0 \times 10^{-12}}{1.0 \times 10^{-8}} \end{aligned}$$

$$\begin{aligned} &= 6.0 \times 10^{-4} \text{ mole/L} \\ \text{grams} &= \text{moles} * \text{molecular mass} \\ &= 6.0 \times 10^{-4} \text{ mole/L} * 24.31 \text{ g/mole} \\ &= 1.4586 \times 10^{-2} \text{ g/L} = 0.014586 \text{ g/L} \\ &= 14.586 \text{ mg/L} \end{aligned}$$