

# Solubility Constant Hmwk.

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I	$1.3 \times 10^{-5}$	0	0
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C		+x	+x
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E		+x	+x
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$$x = 1.3 \times 10^{-5}$$

$$K_{sp} = ?$$

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

$$K_{sp} = (1.3 \times 10^{-5})^2$$

$$K_{sp} = \underline{\underline{1.7 \times 10^{-10}}}$$



$$I \quad 2.7 \times 10^{-3}$$

C

$$+x$$

$$+2x$$

E

$$x$$

$$2x$$

$$K_{sp} = [\text{Mg}] [\text{F}]^2$$

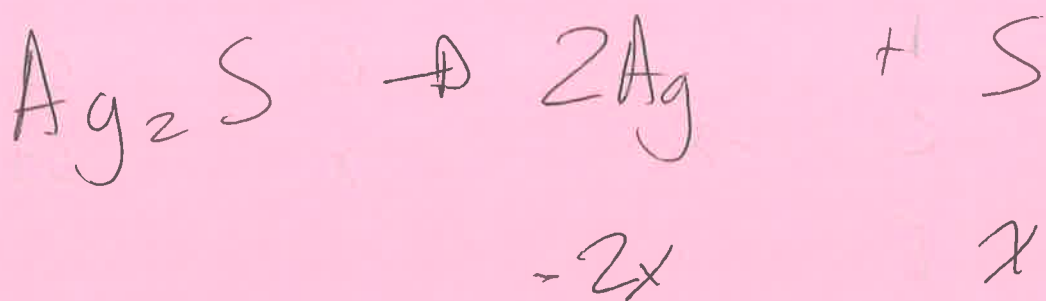
$$= (x) (2x)^2$$

$$= (x) (4x^2)$$

$$= (2.7 \times 10^{-3}) [(4)(2.7 \times 10^{-3})]$$

$$= \underline{\underline{7.94 \times 10^{-6}}}$$

123)



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

$$K_{sp} = (4x^2)(x)$$

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

$$5.6 \times 10^{-49} = 4x^3$$

$$1.4 \times 10^{-49} = x^3$$

$$x = \sqrt[3]{1.4 \times 10^{-49}}$$

$$x = \underline{\underline{5.19 \times 10^{-13}}}$$

124)



$$K_{sp} = [\text{Pb}][\text{Br}]^2$$

$$K_{sp} = (x)(2x)^2 = 4x^3$$

$$6.6 \times 10^{-6} = 4x^3$$

$$x^3 = 1.65 \times 10^{-6}$$

$$x = 1.18 \times 10^{-2} \text{ mol/L}$$

$$m = M \times n$$

$$= 367.0 \text{ g/mol} \times 1.18 \times 10^{-2} \text{ mol/L}$$

$$= \underline{4.34 \text{ g/L}}$$

125)

$$K_{sp} \text{ AgCl} = 1.77 \times 10^{-10} \quad \left. \vphantom{K_{sp} \text{ AgCl}} \right\} \text{ from chart}$$

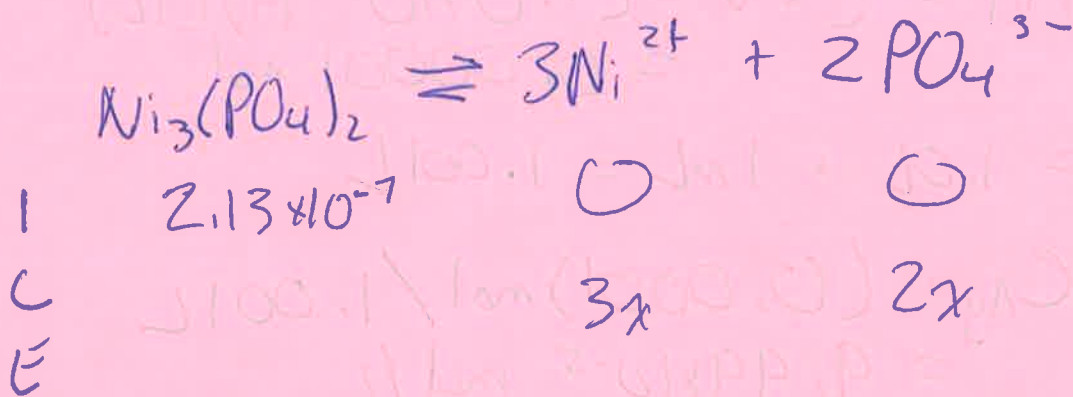
$$K_{sp} \text{ CuCl} = 1.72 \times 10^{-9}$$

Since CuCl has a higher solubility, it would ionize more.



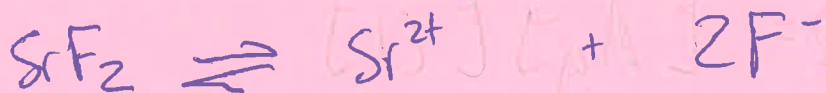
Solubility =  $7.8 \times 10^{-5} \text{ g/L}$  → convert to mol/L

$$\text{solubility} = \frac{7.8 \times 10^{-5} \text{ g}}{366.02 \text{ g/mol}} = 2.13 \times 10^{-7}$$



$$\begin{aligned} K_{sp} &= (3x)^3 (2x)^2 \\ &= 108 x^5 \\ &= 108 (2.13 \times 10^{-7}) \\ &= \underline{\underline{4.75 \times 10^{-92}}} \end{aligned}$$

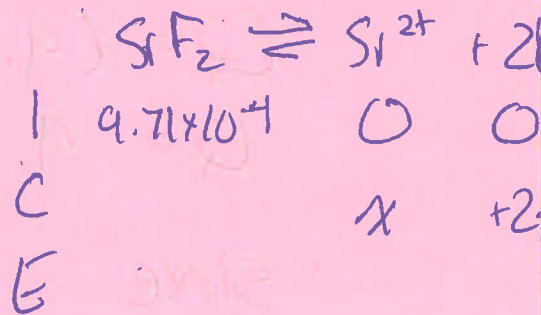
127)



$$x = 12.2 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} \times 10 \text{ (mL} \rightarrow \text{L)}$$

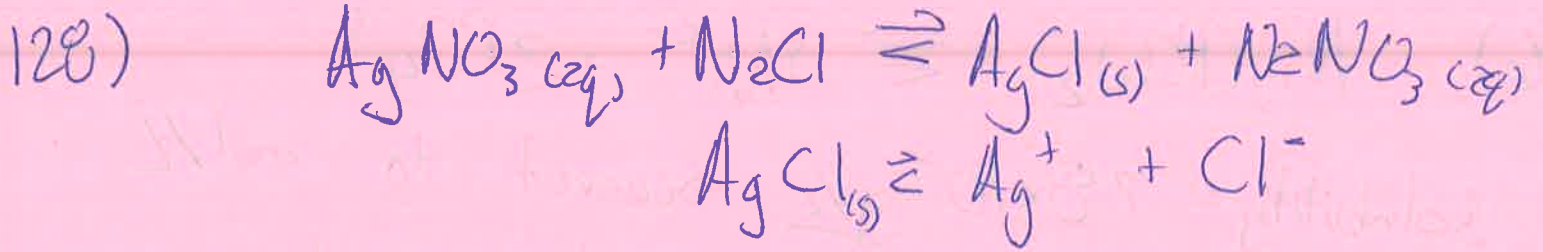
$$\text{solubility} = 0.122 \text{ g/L} \div 125.62 \text{ g/mol}$$

$$\text{solubility} = 9.71 \times 10^{-4} \text{ mol/L}$$



$$\begin{aligned} K_{sp} &= (x)(2x)^2 \\ &= 4x^3 \\ &= \underline{\underline{4.0 \times 10^{-9}}} \end{aligned}$$





$$n_{\text{Ag}} = C \times V = (0.1 \text{ M}) (0.001 \text{ L})$$
$$= 0.0001 \text{ mol}$$

$$n_{\text{Cl}} = C \times V = (1.0 \times 10^{-5} \text{ M}) (1 \text{ L})$$
$$= 0.00001 \text{ mol}$$

$$V_T = 1.0 \text{ L} + 1 \text{ mL} = 1.001 \text{ L}$$

$$C_{\text{Ag}} = (0.0001) \text{ mol} / 1.001 \text{ L}$$
$$= 9.99 \times 10^{-5} \text{ mol/L}$$

$$C_{\text{Cl}} = (0.00001) \text{ mol} / 1.001 \text{ L}$$
$$= 9.99 \times 10^{-6} \text{ mol/L}$$

$$\text{AgCl } K_{\text{sp}} = 1.77 \times 10^{-10}$$

$$Q = [\text{Ag}^+] [\text{Cl}^-]$$

$$Q = (9.99 \times 10^{-5}) (9.99 \times 10^{-6})$$

$$Q = 9.98 \times 10^{-10}$$

• since  $Q > K_{\text{sp}}$ , a precipitate will form