

Equilibrium

7.10 Reaction Quotient and Le Chatelier's Principle

7.11 Introduction to Solubility Equilibria

7.12 Common-Ion Effect

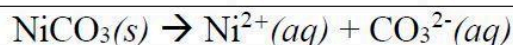
Worksheet Key

- 1) Which of the following is more soluble in water at 25°C?



Ba(OH)₂ is more soluble, as it has a larger K_{sp} value.

- 2) The solubility product constant, K_{sp} , for nickel (II) carbonate is 1.3×10^{-7} at 25°C.
a. Write the balance chemical equation for this dissolving process.



- b. Write the equilibrium expression.

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

- c. Find the maximum molar concentration of all ions in solution and the molar solubility of the solution.

$$K_{\text{sp}} = [\text{Ni}^{2+}][\text{CO}_3^{2-}] = (s)(s) \quad s = \text{molar solubility}$$

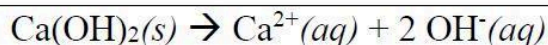
$$K_{\text{sp}} = s^2$$

$$1.3 \times 10^{-7} = s^2$$

$$s = \sqrt{1.3 \times 10^{-7}}$$

$$s = [\text{Ni}^{2+}] = [\text{CO}_3^{2-}] = 3.6 \times 10^{-4} M$$

- 3) The solubility product constant, K_{sp} , for calcium hydroxide is 6.5×10^{-6} at 25°C.
a. Write the balance chemical equation for this dissociation in water.



- b. Write the equilibrium expression.

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

- c. Find the molar concentration of all ions in this solution and the molar solubility of the solution.

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

$$K_{sp} = (s)(2s)^2 \quad s = \text{molar solubility}$$

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

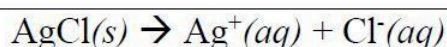
$$6.5 \times 10^{-6} = 4s^3$$

$$s = \sqrt[3]{\frac{6.5 \times 10^{-6}}{4}} = [\text{Ca}^{2+}] = 1.2 \times 10^{-2} M; [\text{OH}^-] = 2s = 2(1.2 \times 10^{-2} M) = 2.4 \times 10^{-2} M$$

- d. Calculate the maximum mass of calcium hydroxide that will dissolve in 100 mL of distilled water at 25°C.

$$100 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.012 \text{ mol Ca(OH)}_2}{1 \text{ L}} \times \frac{74.10 \text{ g Ca(OH)}_2}{1 \text{ mol Ca(OH)}_2} = 8.9 \times 10^{-2} \text{ g Ca(OH)}_2$$

- 4) A 250 mL saturated solution of silver chloride is prepared at 25°C. The solubility product constant, K_{sp} , for silver chloride is 1.8×10^{-10} at 25°C.
- a. Write the balance chemical equation for this dissolving process.



- b. Write the equilibrium expression.

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

- c. Find the molar concentrations of all ions in this solution and the molar solubility of the solution.

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

$$K_{sp} = (s)(s) \quad s = \text{molar solubility}$$

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

$$1.8 \times 10^{-10} = s^2$$

$$s = \sqrt{1.8 \times 10^{-10}}$$

$$s = [\text{Ag}^+] = [\text{Cl}^-] = 1.3 \times 10^{-5} M$$

- d. What are the molar concentrations of all ions in a 500 mL saturated solution of silver chloride at 25°C?

$$[\text{Ag}^+] = [\text{Cl}^-] = 1.3 \times 10^{-5} M$$

Concentration is independent of volume. It is always measured in moles / 1 L.

- 5) What is minimum volume of distilled water that is needed to completely dissolve a 3.0 g sample of CuCO_3 at 25°C? (K_{sp} for CuCO_3 is 2.5×10^{-10} at 25°C.)

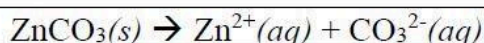
$$K_{\text{sp}} = 2.5 \times 10^{-10} = s^2, s = 1.6 \times 10^{-5} M$$

$$3.0 \text{ g CuCO}_3 \times \frac{1 \text{ mol CuCO}_3}{123.56 \text{ g CuCO}_3} = 0.024 \text{ mol CuCO}_3$$

$$\frac{0.024 \text{ mol CuCO}_3}{V} = 1.6 \times 10^{-5} M$$

$$V = 1500 \text{ L}$$

- 6) At a certain temperature, 9.5×10^{-5} g of zinc carbonate will dissolve in 150 mL of water.
- Write the balanced chemical equation for the dissociation of zinc carbonate in water.



- Calculate the molar solubility of zinc carbonate in water at this temperature.

$$9.5 \times 10^{-5} \text{ g ZnCO}_3 \times \frac{1 \text{ mol ZnCO}_3}{125.40 \text{ g ZnCO}_3} = 7.6 \times 10^{-7} \text{ mol ZnCO}_3$$

$$s = \frac{7.6 \times 10^{-7} \text{ mol ZnCO}_3}{0.15 \text{ L}} = 5.1 \times 10^{-6} M$$

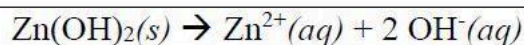
- Calculate the value of the solubility product constant, K_{sp} , for zinc carbonate at this temperature.

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

$$K_{\text{sp}} = (s)(s) = s^2 \quad s = \text{molar solubility}$$

$$K_{\text{sp}} = (5.1 \times 10^{-6})^2 = 2.6 \times 10^{-11}$$

- 7) A 1.0 L saturated solution of zinc hydroxide is prepared, and the concentration of Zn^{2+} is measured to be $4.22 \times 10^{-6} \text{ M}$ at 25°C .
- Write the balanced chemical equation for the dissociation of zinc hydroxide in water.



- Calculate the molar concentration of OH^- in solution.

$$[\text{OH}^-] = 2[\text{Zn}^{2+}] = 2(4.22 \times 10^{-6} \text{ M}) = 8.44 \times 10^{-6} \text{ M}$$

- Calculate the value of the solubility product constant, K_{sp} .

$$K_{\text{sp}} = [\text{Zn}^{2+}][\text{OH}^-]^2 = (4.22 \times 10^{-6})(8.44 \times 10^{-6})^2 = 3.01 \times 10^{-16} \text{ M}$$

- Calculate the maximum mass of zinc hydroxide that will dissolve in 150 mL of distilled water at 25°C

$$150 \text{ mL} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{4.22 \times 10^{-6} \text{ mol Zn(OH)}_2}{1 \text{ L}} \times \frac{99.41 \text{ g Zn(OH)}_2}{1 \text{ mol Zn(OH)}_2} = 6.3 \times 10^{-5} \text{ g Zn(OH)}_2$$

- What is the molar concentration of Zn^{2+} if 200 mL of water evaporates from the solution?

$$[\text{Zn}^{2+}] = 4.22 \times 10^{-6} \text{ M}; \text{ concentration is not affected by changes in volume.}$$

- 8) The solubility product constant, K_{sp} , for barium sulfate is 1.1×10^{-10} at 25°C . Will a precipitate of $\text{BaSO}_4(s)$ form when 210 mL of $4.75 \times 10^{-2} \text{ M}$ $\text{Ba(NO}_3)_2$ is mixed with 315 mL of 0.450 M Li_2SO_4 .

$$\text{Find } [\text{Ba}^{2+}] \quad 0.21 \text{ L solution} \times \frac{4.75 \times 10^{-2} \text{ mol Ba}^{2+}}{1 \text{ L solution}} = 1.0 \times 10^{-2} \text{ mol Ba}^{2+}$$

$$[\text{Ba}^{2+}] = \frac{1.0 \times 10^{-2} \text{ mol Ba}^{2+}}{0.21 \text{ L} + 0.315 \text{ L}} = 2.0 \times 10^{-2} \text{ M}$$

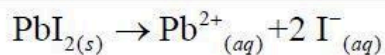
$$\text{Find } [\text{SO}_4^{2-}] \quad 0.315 \text{ L solution} \times \frac{0.450 \text{ mol SO}_4^{2-}}{1 \text{ L solution}} = 0.142 \text{ mol SO}_4^{2-}$$

$$[\text{SO}_4^{2-}] = \frac{0.142 \text{ mol SO}_4^{2-}}{0.21 \text{ L} + 0.315 \text{ L}} = 0.27 \text{ M}$$

$$Q = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = (2.0 \times 10^{-2})(0.27) = 5.4 \times 10^{-3}$$

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

- 9) Find the solubility product constant for lead (II) iodide if the concentration of I^- is found to be $1.25 \times 10^{-3} M$ when the solution is saturated.



$$K_{sp} = [Pb^{2+}][I^{-}]^2 = (0.5 \times [I^{-}])[I^{-}]^2 = (0.5 \times 1.25 \times 10^{-3})(1.25 \times 10^{-3})^2 = 9.77 \times 10^{-10} M$$

- 10) Use the K_{sp} value obtained in the previous question to determine if a precipitate will form when 350 mL of $5.5 \times 10^{-2} M$ lead (II) nitrate is mixed with 250 mL of $4.8 \times 10^{-2} M$ sodium iodide.

$$\text{Find } [Pb^{2+}] \quad 0.35 \text{ L solution} \times \frac{5.5 \times 10^{-2} \text{ mol } Pb^{2+}}{1 \text{ L solution}} = 1.9 \times 10^{-2} \text{ mol } Pb^{2+}$$

$$[Pb^{2+}] = \frac{1.9 \times 10^{-2} \text{ mol } Pb^{2+}}{0.35 \text{ L} + 0.25 \text{ L}} = 3.2 \times 10^{-2} M$$

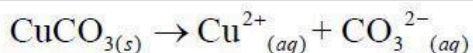
$$\text{Find } [I^{-}] \quad 0.25 \text{ L solution} \times \frac{4.8 \times 10^{-2} \text{ mol } I^{-}}{1 \text{ L solution}} = 1.2 \times 10^{-2} \text{ mol } I^{-}$$

$$[I^{-}] = \frac{1.2 \times 10^{-2} \text{ mol } I^{-}}{0.35 \text{ L} + 0.25 \text{ L}} = 0.020 M$$

$$Q = [Pb^{2+}][I^{-}]^2 = (3.2 \times 10^{-2})(0.020)^2 = 1.3 \times 10^{-5} M$$

$Q > K_{sp}$, so a precipitate will form

- 11) Find the solubility product constant for copper (II) carbonate if the concentration of Cu^{2+} is found to be $1.5 \times 10^{-6} M$ when the solution is saturated.



$$K_{sp} = [Cu^{2+}][CO_3^{2-}] = (1.5 \times 10^{-6})(1.5 \times 10^{-6}) = 2.3 \times 10^{-12} M$$

- 12) Use the K_{sp} value obtained in the previous question to determine if a precipitate will form when 325 mL of $0.50 \times 10^{-6} M$ copper (II) nitrate is mixed with 325 mL of $0.50 \times 10^{-6} M$ potassium carbonate.

$$\text{Find } [\text{Cu}^{2+}] \quad 0.325 \text{ L solution} \times \frac{0.50 \times 10^{-6} \text{ mol Cu}^{2+}}{1 \text{ L solution}} = 1.6 \times 10^{-7} \text{ mol Cu}^{2+}$$

$$[\text{Cu}^{2+}] = \frac{1.6 \times 10^{-7} \text{ mol Cu}^{2+}}{0.325 \text{ L} + 0.325 \text{ L}} = 2.5 \times 10^{-7} M$$

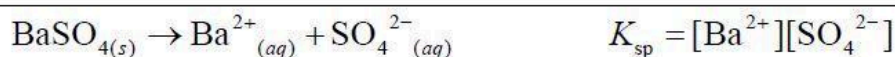
$$\text{Find } [\text{CO}_3^{2-}] \quad 0.325 \text{ L solution} \times \frac{0.50 \times 10^{-6} \text{ mol CO}_3^{2-}}{1 \text{ L solution}} = 1.6 \times 10^{-7} \text{ mol CO}_3^{2-}$$

$$[\text{CO}_3^{2-}] = \frac{1.6 \times 10^{-7} \text{ mol CO}_3^{2-}}{0.325 \text{ L} + 0.325 \text{ L}} = 2.5 \times 10^{-7} M$$

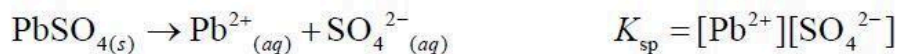
$$Q = [\text{Cu}^{2+}][\text{CO}_3^{2-}] = (2.5 \times 10^{-7})(2.5 \times 10^{-7}) = 6.3 \times 10^{-14} M$$

$Q < K_{sp}$, no precipitate will form

- 13) If a $0.50 M$ solution of K_2SO_4 is slowly poured into a beaker containing $0.25 M$ barium nitrate and $0.30 M$ lead (II) nitrate at 25°C , what will be the first precipitate that forms? K_{sp} for barium sulfate is 1.1×10^{-10} and K_{sp} for lead (II) sulfate is 1.6×10^{-8} .



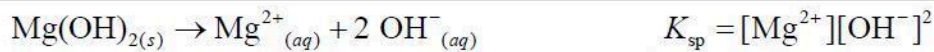
$$[\text{SO}_4^{2-}] = \frac{K_{sp}}{[\text{Ba}^{2+}]} = \frac{1.1 \times 10^{-10}}{0.25 M} = 4.4 \times 10^{-10} M$$



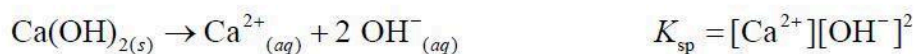
$$[\text{SO}_4^{2-}] = \frac{K_{sp}}{[\text{Pb}^{2+}]} = \frac{1.6 \times 10^{-8}}{0.30 M} = 5.3 \times 10^{-8} M$$

$\text{BaSO}_{4(s)}$ will be the first precipitate that forms, as $[\text{SO}_4^{2-}]$ will reach $4.4 \times 10^{-10} M$ before it reaches $5.3 \times 10^{-8} M$.

- 14) If a 0.50 M solution of KOH is slowly poured into a beaker containing 0.35 M magnesium nitrate and 0.032 M calcium nitrate at 25°C, what will be the first precipitate that forms? K_{sp} for calcium hydroxide is 6.5×10^{-6} and K_{sp} for magnesium hydroxide is 6.3×10^{-10} .



$$[\text{OH}^{-}] = \sqrt{\frac{K_{sp}}{[\text{Mg}^{2+}]}} = \frac{6.3 \times 10^{-10}}{0.35 M} = 4.2 \times 10^{-5} M$$



$$[\text{OH}^{-}] = \sqrt{\frac{K_{sp}}{[\text{Ca}^{2+}]}} = \frac{6.5 \times 10^{-6}}{0.032 M} = 1.4 \times 10^{-2} M$$

$\text{Mg(OH)}_{2(s)}$ will be the first precipitate that forms, as $[\text{OH}^{-}]$ will reach $4.2 \times 10^{-5} M$ before it reaches $1.4 \times 10^{-2} M$.

- 15) Explain why the solubility of AgBr decreases when NaBr is added to the system.

When additional Br^{-} is added to the system, the equilibrium in the following reaction ($\text{AgBr}(s) \rightleftharpoons \text{Ag}^{+}(aq) + \text{Br}^{-}(aq)$) shifts to the left to relieve the stress (Le Chatelier's principle). LO 6.23

- 16) The solubility product constant, K_{sp} , for Ag_2SO_4 is 1.2×10^{-5} at 25°C. Find $[\text{Ag}^{+}]$ and $[\text{SO}_4^{2-}]$ after 0.755 g of AgNO_3 are added to a 500.0 mL saturated solution of Ag_2SO_4 and equilibrium is established. Assume that the total volume of the solution remains the same and the final temperature is 25°C.

Step 1) Find moles of Ag^{+} in original solution

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

$$K_{sp} = 4s^2s = 4s^3$$

$$s = \sqrt[3]{\frac{1.2 \times 10^{-5}}{4}}$$

$$s = [\text{SO}_4^{2-}] = 1.4 \times 10^{-2} M$$

$$n_{\text{Ag}^{+}} = 0.5000 \text{ L} \times \frac{2.8 \times 10^{-2} \text{ mol Ag}^{+}}{1 \text{ L}}$$

$$n_{\text{Ag}^{+}} = 1.4 \times 10^{-2} \text{ mol Ag}^{+}$$

$$[\text{Ag}^{+}] = 2s = 2(1.4 \times 10^{-2} M) = 2.8 \times 10^{-2} M$$

Step 2) Find moles of Ag^{+} that were added

$$0.755 \text{ g AgNO}_3 \times \frac{1 \text{ mol AgNO}_3}{169.88 \text{ g AgNO}_3} \times \frac{1 \text{ mol Ag}^{+}}{1 \text{ mol AgNO}_3} = 4.44 \times 10^{-3} \text{ mol Ag}^{+}$$

Step 3) Find $[\text{Ag}^+]$ in the final solution

$$[\text{Ag}^+] = \frac{1.4 \times 10^{-2} \text{ mol Ag}^+ + 0.444 \times 10^{-2} \text{ mol Ag}^+}{0.5000 \text{ L}}$$

$$[\text{Ag}^+] = \frac{1.8 \times 10^{-2} \text{ mol Ag}^+}{0.5000 \text{ L}}$$

$$[\text{Ag}^+] = 0.036M$$

Step 4) Find $[\text{SO}_4^{2-}]$ in the final solution

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

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

$$[\text{SO}_4^{2-}] = \frac{1.2 \times 10^{-5}}{0.036^2}$$

$$[\text{SO}_4^{2-}] = 9.3 \times 10^{-3}M$$