

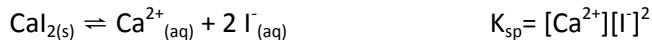
Solubility Problems SOLUTIONS

1. Write balanced chemical equations and the K_{sp} expressions for the dissolving of the following compounds in water.

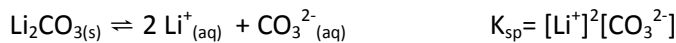
a) sodium sulfide



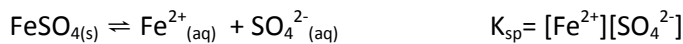
b) calcium iodide



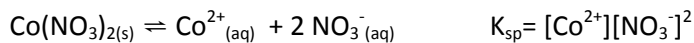
c) lithium carbonate



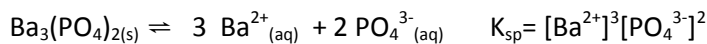
d) iron(II) sulfate



e) cobalt(II) nitrate

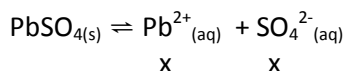


f) barium phosphate



2. What are the concentrations of the resulting ions in saturated aqueous solutions of the following compounds?

a) lead (II) sulfate



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

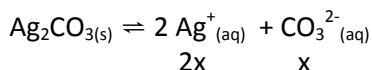
$$1.8 \times 10^{-8} = x^2$$

$$[\text{Pb}^{2+}] = 1.3 \times 10^{-4} \text{ M}, [\text{SO}_4^{2-}] = 1.3 \times 10^{-4} \text{ M}$$

$$x = 1.3 \times 10^{-4} \text{ M}$$

$$\text{therefore } [\text{Pb}^{2+}] = 1.3 \times 10^{-4} \text{ M and } [\text{SO}_4^{2-}] = 1.3 \times 10^{-4} \text{ M}$$

b) silver carbonate



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

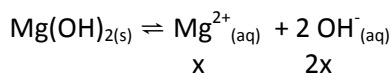
$$8.5 \times 10^{-12} = 4x^3$$

$$[\text{Ag}^+] = 2.6 \times 10^{-4} \text{ M}, [\text{CO}_3^{2-}] = 1.3 \times 10^{-4} \text{ M}$$

$$x = 1.3 \times 10^{-4} \text{ M}$$

$$\text{therefore } [\text{Ag}^+] = 2.6 \times 10^{-4} \text{ M and } [\text{CO}_3^{2-}] = 1.3 \times 10^{-4} \text{ M}$$

c) magnesium hydroxide



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

$$[\text{Mg}^{2+}] = 1.1 \times 10^{-4} \text{ M}, [\text{OH}^-] = 2.2 \times 10^{-4} \text{ M}$$

$$5.6 \times 10^{-12} = 4x^3$$

$$x = 1.1 \times 10^{-4} \text{ M}$$

$$\text{therefore } [\text{Mg}^{2+}] = 1.1 \times 10^{-4} \text{ M and } [\text{OH}^-] = 2.2 \times 10^{-4} \text{ M}$$

3. Calculate the K_{sp} values of the following substances from their solubility in water.

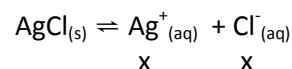
a) silver chloride with a solubility of $1.6 \times 10^{-3} \text{ g/L}$

$$[1.2 \times 10^{-10}]$$

$$\text{MM} = 143.32 \text{ g/mol}$$

$$x = \frac{1.6 \times 10^{-3} \text{ g/L}}{143.32 \text{ g/mol}}$$

$$x = 1.1 \times 10^{-5} \text{ mol/L}$$



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

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

$$K_{sp} = (1.1 \times 10^{-5})^2$$

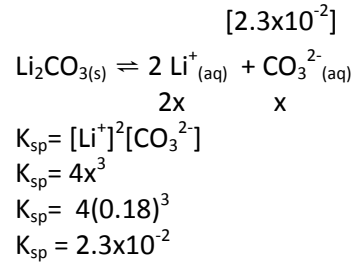
$$K_{sp} = 1.2 \times 10^{-10}$$

- b) lithium carbonate with a solubility of 13.0 g/L

$$MM = 73.89 \text{ g/mol}$$

$$x = \frac{13.0 \text{ g/L}}{73.89 \text{ g/mol}}$$

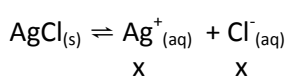
$$x = 0.18 \text{ mol/L}$$



4. Calculate the concentrations of ions in saturated aqueous solutions of the following.

- a) silver iodide

$$[\text{Ag}^+] = 9.2 \times 10^{-9} \text{ M}, [\text{I}^-] = 9.2 \times 10^{-9} \text{ M}$$



$$8.5 \times 10^{-17} = x^2$$

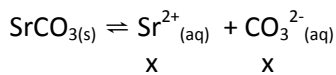
$$x = 9.2 \times 10^{-9}$$

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

$$\text{therefore } [\text{Ag}^+] = 9.2 \times 10^{-9} \text{ M and } [\text{Cl}^-] = 9.2 \times 10^{-9} \text{ M}$$

- b) strontium carbonate

$$[\text{Sr}^{2+}] = 2.4 \times 10^{-5} \text{ M}, [\text{CO}_3^{2-}] = 2.4 \times 10^{-5} \text{ M}$$



$$5.6 \times 10^{-10} = x^2$$

$$x = 2.4 \times 10^{-5} \text{ M}$$

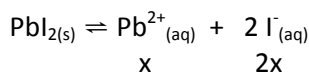
$$K_{sp} = [\text{Sr}^{2+}][\text{CO}_3^{2-}]$$

$$\text{therefore } [\text{Sr}^{2+}] = 2.4 \times 10^{-5} \text{ M and } [\text{CO}_3^{2-}] = 2.4 \times 10^{-5} \text{ M}$$

5. What are the ion concentrations in saturated solutions of the following:

- a) lead(II) iodide

$$[\text{Pb}^{2+}] = 1.3 \times 10^{-3} \text{ M}, [\text{I}^-] = 2.6 \times 10^{-3} \text{ M}$$



$$8.5 \times 10^{-9} = 4x^3$$

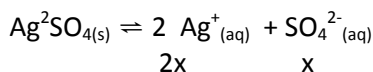
$$x = 1.3 \times 10^{-3} \text{ M}$$

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

$$\text{therefore } [\text{Pb}^{2+}] = 1.3 \times 10^{-3} \text{ M and } [\text{I}^-] = 2.6 \times 10^{-3} \text{ M}$$

- b) silver sulfate

$$[\text{Ag}^+] = 2.8 \times 10^{-2} \text{ M}, [\text{SO}_4^{2-}] = 1.1 \times 10^{-2} \text{ M}$$



$$1.2 \times 10^{-5} = 4x^3$$

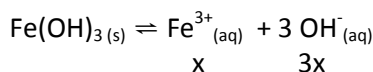
$$x = 5.7 \times 10^{-3} \text{ M}$$

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

$$\text{therefore } [\text{Ag}^+] = 0.011 \text{ M and } [\text{SO}_4^{2-}] = 1.1 \times 10^{-2} \text{ M}$$

- c) iron(III) hydroxide

$$[\text{Fe}^{3+}] = 9.9 \times 10^{-11} \text{ M}, [\text{OH}^-] = 3.0 \times 10^{-10} \text{ M}$$



$$2.6 \times 10^{-39} = 27x^4$$

$$x = 9.9 \times 10^{-11} \text{ M}$$

$$K_{sp} = [\text{Fe}^{3+}][\text{OH}^-]^3$$

$$\text{therefore } [\text{Fe}^{3+}] = 9.9 \times 10^{-11} \text{ M and } [\text{OH}^-] = 3.0 \times 10^{-10} \text{ M}$$

6. Calculate the
- K_{sp}
- values of the substances below from their solubility in water:

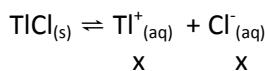
- a) thallium(I) chloride, 3.4 g/L at 25°C

$$[2.0 \times 10^{-4}]$$

$$MM = 239.85 \text{ g/mol}$$

$$x = \frac{3.4 \text{ g/L}}{239.85 \text{ g/mol}}$$

$$x = 0.014 \text{ mol/L}$$



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

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

$$K_{sp} = (0.014)^2$$

$$K_{sp} = 2.0 \times 10^{-4}$$

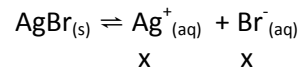
- b) silver bromide,
- 1.3×10^{-4}
- g/L at
- 20°C

MM = 187.80 g/mol

$$x = \frac{1.3 \times 10^{-4} \text{ g/L}}{187.80 \text{ g/mol}}$$

$$x = 6.9 \times 10^{-7} \text{ mol/L}$$

$$[4.8 \times 10^{-13}]$$



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

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

$$K_{sp} = (6.9 \times 10^{-7})^2$$

$$K_{sp} = 4.8 \times 10^{-13}$$

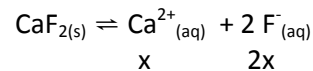
- c) calcium fluoride,
- 1.6×10^{-2}
- g/L at
- 20°C

MM = 78.80 g/mol

$$x = \frac{1.6 \times 10^{-2} \text{ g/L}}{78.80 \text{ g/mol}}$$

$$x = 2.0 \times 10^{-4} \text{ mol/L}$$

$$[3.4 \times 10^{-11}]$$



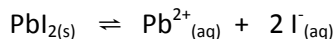
$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2$$

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

$$K_{sp} = 4(2.0 \times 10^{-4})^3$$

$$K_{sp} = 3.4 \times 10^{-11}$$

7. Calculate the maximum iodide ion concentration for lead(II) iodide dissolved in a
- 1.00×10^{-2}
- M solution of lead(II) nitrate.
- [9.2 × 10⁻⁴ M]



MR	1	1	2
I	-----	0.0100	0
C	-----	+x	+2x
E	-----	0.0100+x	2x

$$0.010/K_{sp} > 1000, \text{ approx. works}$$

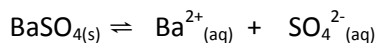
$$K_{sp} = [\text{Pb}^{2+}][\text{I}^-]^2$$

$$8.5 \times 10^{-9} = (1.00 \times 10^{-2} + x) [2x]^2$$

$$x = 4.6 \times 10^{-4}$$

$$[\text{I}^-] = 9.2 \times 10^{-4} \text{ M}$$

8. Calculate the maximum barium ion concentration in a 0.010 M aqueous solution of sodium sulfate. The
- K_{sp}
- of barium sulfate is
- 1.1×10^{-11}
- .
- [1.1 × 10⁻⁹]



MR	1	1	1
I	-----	0	0.010
C	-----	+x	+x
E	-----	x	0.010+x

$$0.010/K_{sp} > 1000, \text{ approx. works}$$

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

$$1.1 \times 10^{-11} = [\text{Ba}^{2+}](0.010 + x)$$

$$[\text{Ba}^{2+}] = 1.1 \times 10^{-9}$$

9. Calculate the maximum magnesium ion concentration in a 0.020 M aqueous solution of barium hydroxide. The
- K_{sp}
- of magnesium hydroxide is
- 1.2×10^{-11}
- .
- [7.5 × 10⁻⁹]



MR	1	1	2
I	-----	0	0.040
C	-----	+x	+2x
E	-----	x	0.040+2x

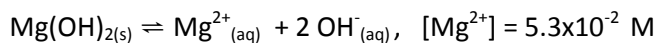
$$0.040/K_{sp} > 1000, \text{ approx. works}$$

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

$$1.2 \times 10^{-11} = [\text{Mg}^{2+}](0.040 + 2x)^2$$

$$[\text{Mg}^{2+}] = 7.5 \times 10^{-9} \text{ M}$$

10. Upon addition of hydroxide ions to sea water,
- Mg(OH)_2
- precipitates. If the magnesium ion concentration in sea water is
- 5.3×10^{-2}
- M, calculate the maximum hydroxide ion concentration in sea water.
- [1.0 × 10⁻⁵]



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

$$5.6 \times 10^{-12} = [5.3 \times 10^{-2}] [\text{OH}^-]^2$$

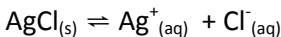
$$[\text{Mg}^{2+}] = 1.0 \times 10^{-5}$$

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11. A sample of sea water contains 0.53 M of Cl^- ions and 8.4×10^{-4} M of Br^- ions. What concentration of added Ag^+ ions would cause precipitation of AgCl and AgBr ? Which of these two halides would precipitate first? The K_{sp} for AgCl is 1.6×10^{-3} and the K_{sp} for AgBr is 6.5×10^{-13} . [Cl^- : 3.0×10^{-3} Br^- : 7.7×10^{-10} , AgBr precipitates first]

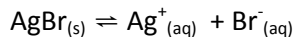
Since the K_{sp} for AgBr is smaller than the K_{sp} for AgCl , AgBr will precipitate first.

For Cl^- 

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

$$1.6 \times 10^{-3} = [\text{Ag}^+][0.53]$$

$$[\text{Ag}^+] = 3.0 \times 10^{-3} \text{ M}$$

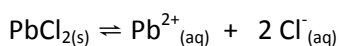
For Br^- 

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

$$6.5 \times 10^{-13} = [\text{Ag}^+][8.4 \times 10^{-4}]$$

$$[\text{Ag}^+] = 7.7 \times 10^{-10} \text{ M}$$

12. How many mg of Pb^{2+} must be present in 10.0 mL of 0.135 M NaCl solution for PbCl_2 to precipitate? [1.4 mg]



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

$$1.2 \times 10^{-5} = [\text{Pb}^{2+}][0.135]^2$$

$$[\text{Pb}^{2+}] = 6.6 \times 10^{-4} \text{ M}$$

Moles of Pb^{2+} in 10.0 mL

$$n = cv$$

$$n = (6.6 \times 10^{-4} \text{ M})(0.010 \text{ L})$$

$$n = 6.6 \times 10^{-6} \text{ mol}$$

mass of Pb^{2+}

$$\text{MM of Pb} = 207.20 \text{ g/mol}$$

$$m = n\text{MM}$$

$$m = (6.6 \times 10^{-6} \text{ mol})(207.20 \text{ g/mol})$$

$$m = 0.0014 \text{ g}$$

$$m = 1.4 \text{ mg}$$

13. Will a precipitate of CaF_2 form when 0.084 g of sodium fluoride is dissolved in 1.00 L of a 0.010 M aqueous solution of calcium chloride. K_{sp} for calcium fluoride is 3.9×10^{-11} . [$Q = 4.0 \times 10^{-8}$, yes a ppt will form]

$$\text{MM of NaF} = 41.99 \text{ g/mol}$$

$$n \text{ of NaF} = m/\text{MM}$$

$$= (0.084$$

$$\text{g})/(41.99 \text{ g/mol})$$

$$= 2.00 \times 10^{-3} \text{ mol}$$

$$[\text{F}^-] = 2.00 \times 10^{-3} \text{ M}$$

$$[\text{Ca}^{2+}] \text{ from } \text{CaCl}_2 = 0.010 \text{ M}$$

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

$$Q = [\text{Ca}^{2+}]_A [\text{F}^-]_A^2$$

$$Q = (0.010 \text{ M})(2.00 \times 10^{-3} \text{ M})^2$$

$$Q = 4.0 \times 10^{-8}$$

Q is greater than K_{sp} therefore a precipitate will form.

14. In which of the following reactions does a precipitate form?

a) 10.0 mL of 0.010 M AgNO_3 and 10.0 mL of 0.10 M Na_2SO_4 . (K_{sp} for $\text{Ag}_2\text{SO}_4 = 1.2 \times 10^{-5}$) [$Q = 1.25 \times 10^{-6}$]

$$[\text{Ag}^+] = (0.010 \text{ L})(0.010 \text{ M})/(0.020 \text{ L})$$

$$= 0.0050 \text{ M}$$

$$[\text{SO}_4^{2-}] = (0.010 \text{ L})(0.010 \text{ M})/(0.020 \text{ L})$$

$$= 0.050 \text{ M}$$



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

$$Q = [\text{Ag}^+]_A^2 [\text{SO}_4^{2-}]_A$$

$$Q = (0.0050 \text{ M})^2 (0.050 \text{ M})$$

$$Q = 1.25 \times 10^{-6}$$

Q is less than K_{sp} therefore no precipitate will form.

b) 1.0 mL of 0.10 M $\text{Ca}(\text{NO}_3)_2$ and 1.0 L of 0.010 M NaF . (K_{sp} for $\text{CaF}_2 = 3.45 \times 10^{-11}$) [$Q = 1.0 \times 10^{-8}$]

$$[\text{Ca}^{2+}] = (0.0010 \text{ L})(0.10 \text{ M})/(1.001 \text{ L})$$

$$= 1.0 \times 10^{-5} \text{ M}$$

$$[\text{F}^-] = (1.0 \text{ L})(0.010 \text{ M})/(1.001 \text{ L})$$

$$= 0.010 \text{ M}$$

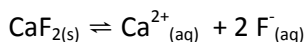
$$K_{sp} = [\text{Ca}^{2+}][\text{F}^-]^2 = 3.45 \times 10^{-11}$$

$$Q = [\text{Ca}^{2+}]_A [\text{F}^-]_A^2$$

$$Q = (1.0 \times 10^{-5} \text{ M})(0.010 \text{ M})^2$$

$$Q = 1.0 \times 10^{-8}$$

Q is greater than K_{sp} therefore a precipitate will form.



Name: _____

c) 5.0 mL of 0.0040 M AgNO_3 and 15 mL of a solution containing 1.5 mg of Br^- ions. $[Q=9.5 \times 10^{-7}]$

$$[\text{Ag}^+] = (0.0050 \text{ L})(0.0040 \text{ M}) / (0.020 \text{ L})$$

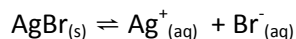
$$= 0.0010 \text{ M}$$

$$n \text{ of Br} = (0.0015 \text{ g}) / (79.90 \text{ g/mol})$$

$$= 1.9 \times 10^{-5} \text{ mol}$$

$$[\text{Br}^-] = (1.9 \times 10^{-5} \text{ mol}) / (0.020 \text{ L})$$

$$= 9.5 \times 10^{-4} \text{ M}$$



$$K_{sp} = [\text{Ag}^+][\text{Br}^-] = 5.4 \times 10^{-13}$$

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

$$Q = (0.0010 \text{ M})(9.5 \times 10^{-4} \text{ M})$$

$$Q = 9.5 \times 10^{-7} \text{ M}$$

Q is greater than K_{sp} therefore a precipitate will form.

15. Would you expect a precipitate of silver bromate ($K_{sp}=1.2 \times 10^{-11}$) to form when 50.0 mL of 0.0020 M silver nitrate is added to 250.0 mL of 0.020 M potassium bromate. $[Q=5.6 \times 10^{-6}]$

$$[\text{Ag}^+] = (0.050 \text{ L})(0.0020 \text{ M}) / (0.300 \text{ L})$$

$$= 3.3 \times 10^{-4} \text{ M}$$

$$[\text{BrO}_3^-] = (0.250 \text{ L})(0.020 \text{ M}) / (0.300 \text{ L})$$

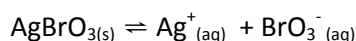
$$= 0.017 \text{ M}$$

$$K_{sp} = [\text{Ag}^+][\text{BrO}_3^-] = 1.2 \times 10^{-11}$$

$$Q = [\text{Ag}^+]_A [\text{BrO}_3^-]_A$$

$$Q = (3.3 \times 10^{-4} \text{ M})(0.017 \text{ M})$$

$$Q = 5.6 \times 10^{-6} \text{ M}$$

Q is greater than K_{sp} therefore a precipitate will form.

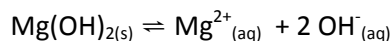
16. Will a precipitate of $\text{Mg}(\text{OH})_2$ form when 1.00 mL of 0.010 M $\text{Ca}(\text{OH})_2$ is added to 1.0 L of 0.20 M $\text{Mg}(\text{NO}_3)_2$. You may ignore the volume change caused by the addition of 1.00 mL. $[Q=8.0 \times 10^{-11}]$

$$[\text{Mg}^{2+}] = (1.0 \text{ L})(0.20 \text{ M}) / (1.00 \text{ L})$$

$$= 0.20 \text{ M}$$

$$[\text{OH}^-] = (2)(0.0010 \text{ L})(0.010 \text{ M}) / (1.00 \text{ L})$$

$$= 2.0 \times 10^{-5} \text{ M}$$



$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 3.45 \times 10^{-11}$$

$$Q = [\text{Mg}^{2+}]_A [\text{OH}^-]^2_A$$

$$Q = (0.20 \text{ M})(2.0 \times 10^{-5} \text{ M})^2$$

$$Q = 8.0 \times 10^{-11}$$

Q is greater than K_{sp} therefore a precipitate will form.

17. One litre of solution has 100.0 mg of Ba^{2+} and 10.0 g of Sr^{2+} . Within what range must the $[\text{CrO}_4^{2-}]$ be to precipitate barium without precipitating strontium. ($K_{sp} \text{ BaCrO}_4 = 1.2 \times 10^{-10}$ $K_{sp} \text{ SrCrO}_4 = 3.6 \times 10^{-5}$) $[1.6 \times 10^{-7} \text{ M} < [\text{CrO}_4^{2-}] < 3.3 \times 10^{-4} \text{ M}]$

Since the K_{sp} for BaCrO_4 is smaller than the K_{sp} for SrCrO_4 , BaCrO_4 will precipitate first.

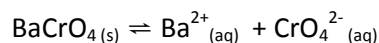
For Ba^{2+}

$$n \text{ of Ba}^{2+} = (0.100 \text{ g}) / (137.33 \text{ g/mol})$$

$$= 7.3 \times 10^{-4} \text{ mol}$$

$$[\text{Ba}^{2+}] = (7.3 \times 10^{-4} \text{ mol}) / (1.0 \text{ L})$$

$$= 7.3 \times 10^{-4} \text{ M}$$



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

$$1.2 \times 10^{-10} = [7.3 \times 10^{-4} \text{ M}][\text{CrO}_4^{2-}]$$

$$[\text{CrO}_4^{2-}] = 1.6 \times 10^{-7} \text{ M}$$

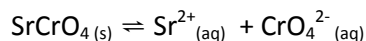
For Sr^{2+}

$$n \text{ of Ba}^{2+} = (10.0 \text{ g}) / (87.62 \text{ g/mol})$$

$$= 0.11 \text{ mol}$$

$$[\text{Ba}^{2+}] = (0.11 \text{ mol}) / (1.0 \text{ L})$$

$$= 0.11 \text{ M}$$



$$K_{sp} = [\text{Sr}^{2+}][\text{CrO}_4^{2-}]$$

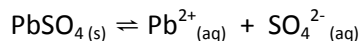
$$3.6 \times 10^{-5} = [0.11 \text{ M}][\text{CrO}_4^{2-}]$$

$$[\text{CrO}_4^{2-}] = 3.3 \times 10^{-4} \text{ M}$$

Therefore $1.6 \times 10^{-7} \text{ M} < [\text{CrO}_4^{2-}] < 3.3 \times 10^{-4} \text{ M}$

Name: _____

18. A 0.010 M aqueous solution of Na_2SO_4 is added one drop at a time to 1.00 L of 0.0010 M lead(II) nitrate. What is the minimum volume of sodium sulfate that must be added to form a precipitate of lead(II) sulfate. The K_{sp} for lead(II) sulfate is 1.3×10^{-8} . [1.3 mL]



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

$$1.3 \times 10^{-8} = (0.0010 \text{ M}) [\text{SO}_4^{2-}]$$

$$[\text{SO}_4^{2-}] = 1.3 \times 10^{-5}$$

To determine the volume of 0.01M Na_2SO_4

$$C_1 V_1 = C_2 V_2$$

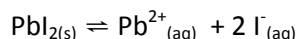
$$V_1 = C_2 V_2 / C_1$$

$$V_1 = (1.3 \times 10^{-5})(1.00 \text{ L}) / (0.010 \text{ M})$$

$$V_1 = 1.3 \times 10^{-3} \text{ L} = 1.3 \text{ mL}$$

19. Compare the molar solubility of PbI_2 in a) pure water and b) in 0.10 M NaI. [a) 1.3×10^{-3} , b) 8.5×10^{-7}]

a) in pure water

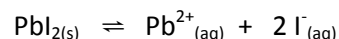


$$x \quad 2x$$

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

$$8.5 \times 10^{-9} = 4x^3$$

$$x = 1.3 \times 10^{-3}$$

Therefore the molar solubility in water is $1.3 \times 10^{-3} \text{ M}$.b) in 0.10 M NaI $K_{sp} = 8.5 \times 10^{-9}$ 

MR	1	1	2
I	-----	0	0.10
C	-----	+x	+x
E	-----	x	0.10+x

0.10/ K_{sp} > 1000, approx. works

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

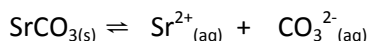
$$8.5 \times 10^{-9} = x(0.10+x)^2$$

$$8.5 \times 10^{-9} = x(0.10+x)^2$$

$$x = 8.5 \times 10^{-7}$$

Therefore the molar solubility in 0.10 NaI is $8.5 \times 10^{-7} \text{ M}$

20. How many grams of SrCO_3 will dissolve in 250 mL of 0.080 mol/L SrNO_3 ? [2.6 $\times 10^{-7}$ g]



MR	1	1	1
I	-----	0.080	0
C	-----	+x	+x
E	-----	0.080+x	x

0.080/ K_{sp} > 1000, approx. works $K_{sp} = 5.6 \times 10^{-10}$

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

$$5.6 \times 10^{-10} = (0.080+x)(x)$$

$$x = 7.0 \times 10^{-9}$$

$$n = cv$$

$$= (7.0 \times 10^{-9})(0.250 \text{ L})$$

$$= 1.8 \times 10^{-9} \text{ mol}$$

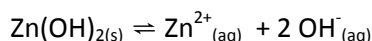
$$\text{MM} = 147.63 \text{ g/mol}$$

$$m = n \text{MM}$$

$$m = (1.8 \times 10^{-9} \text{ mol})(147.63 \text{ g/mol})$$

$$m = 2.6 \times 10^{-7} \text{ g}$$

21. For each of the following substances, calculate the milligrams per millilitre of metallic ion that can remain at equilibrium in a solution having a $[\text{OH}^{-}] = 1.0 \times 10^{-4}$.

a) $\text{Zn}(\text{OH})_2$, $K_{sp} = 4.3 \times 10^{-17}$ [2.8 $\times 10^{-7}$ mg/mL]

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

$$4.3 \times 10^{-17} = [\text{Zn}^{2+}](1.0 \times 10^{-4})^2$$

$$[\text{Zn}^{2+}] = 4.3 \times 10^{-9} \text{ M}$$

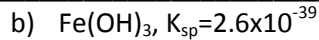
$$\text{sol} = (4.3 \times 10^{-9} \text{ mol/L})(65.38 \text{ g/mol})$$

$$= 2.8 \times 10^{-7} \text{ g/L}$$

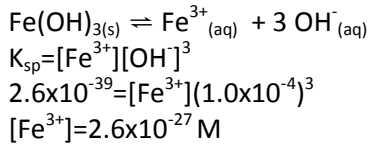
$$= 2.8 \times 10^{-7} \text{ mg/mL}$$

Name: _____

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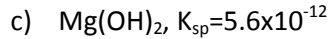
$[1.5 \times 10^{-25} \text{ mg/mL}]$



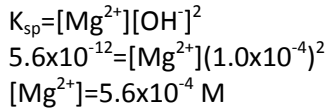
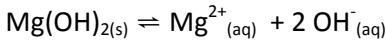
$$\text{sol} = (2.6 \times 10^{-27} \text{ mol/L})(55.85 \text{ g/mol})$$

$$= 1.5 \times 10^{-25} \text{ g/L}$$

$$= \mathbf{1.5 \times 10^{-25} \text{ mg/mL}}$$



$[1.4 \times 10^{-2} \text{ mg/mL}]$

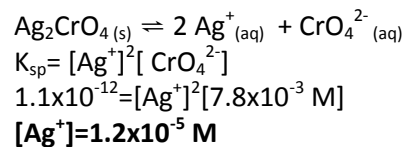
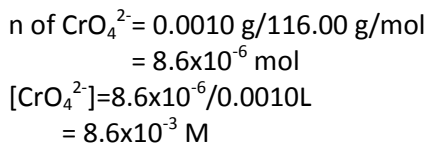
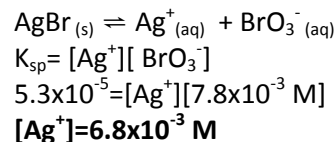
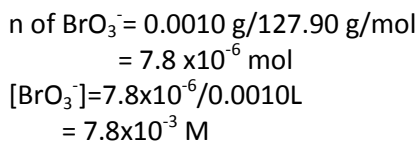
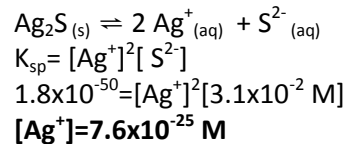
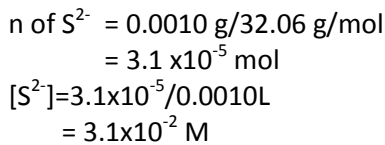
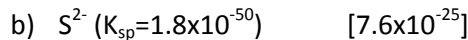
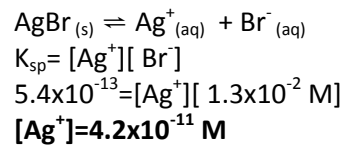
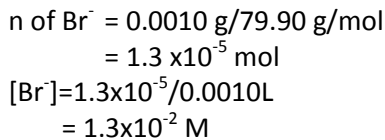


$$\text{sol} = (5.6 \times 10^{-4} \text{ mol/L})(24.31 \text{ g/mol})$$

$$= 1.4 \times 10^{-2} \text{ g/L}$$

$$= \mathbf{1.4 \times 10^{-2} \text{ mg/mL}}$$

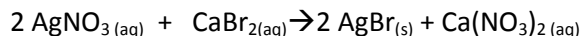
22. Calculate the $[\text{Ag}^{+}]$ needed to begin precipitation of each of the following anions from solutions containing 1 mg of anion per mL of solution.



Name: _____

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23. What is the solubility in mol/L of AgBr in a solution resulting from the addition of 50.0 mL of 0.010 CaBr₂ to 50.0 mL of 0.0080 M AgNO₃? ($K_{sp}=5.4 \times 10^{-13}$) [9.0x10⁻¹¹]



c	0.0080	0.010
V	0.050	0.050
n	4.0x10 ⁻⁴	5.0x10 ⁻⁴
LF	4.0x10 ⁻⁴ /2= 2.0x10 ⁻⁴	5.0x10 ⁻⁴ /1 =5.0x10 ⁻⁴

Limiting

Excess CaBr₂ = 5.0x10⁻⁴-2.0x10⁻⁴ (amount used)
= 3.0x10⁻⁴ mol

Initial [Br⁻] = 2(3.0x10⁻⁴ mol)/0.100 L = 6.0x10⁻³ M

	$\text{AgBr}_{(\text{s})} \rightleftharpoons \text{Ag}^+_{(\text{aq})} + \text{Br}^-_{(\text{aq})}$		
MR	1	1	1
I	---	0	6.0x10 ⁻³
C	---	+x	+x
E	---	x	6.0x10 ⁻³ +x

$K_{sp} = 5.4 \times 10^{-13}$
0.0060/ K_{sp} >1000
Approx works

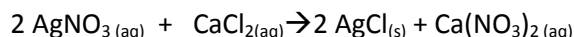
$$K_{sp} = [\text{Ag}^+][\text{Br}^-]$$

$$5.4 \times 10^{-13} = (x)(6.0 \times 10^{-3} + x)$$

$$x = 9.0 \times 10^{-11}$$

The solubility of AgBr is 9.0x10⁻¹¹

24. If 50.0 mL of 0.10 M AgNO₃ is added to 150 mL of 0.10 M CaCl₂, what is the resulting concentration of each ion in the final solution? ($K_{sp}=1.8 \times 10^{-10}$) [Ag⁺=1.4x10⁻⁹M, Cl⁻=0.125 M, NO₃⁻=0.025M, Ca²⁺=0.075 M]



c	0.10	0.10
V	0.050	0.150
n	0.0050	0.015
LF	0.0050/2= 0.0025	0.015/1 =0.015

Limiting

Excess CaCl₂ = 0.015-0.0025 (amount used)
= 0.0125 mol

Initial [Cl⁻] = 2(0.0125)/0.200 L = 0.125 M

	$\text{AgCl}_{(\text{s})} \rightleftharpoons \text{Ag}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$		
MR	1	1	1
I	---	0	0.125
C	---	+x	+x
E	---	x	0.125+x

$K_{sp} = 1.8 \times 10^{-10}$
0.125/ K_{sp} >1000
Approx works

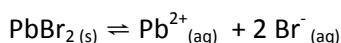
$$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$$

$$1.8 \times 10^{-10} = (0.125 + x)(x)$$

$$x = 1.4 \times 10^{-9}$$

[Ag⁺] = 1.4x10⁻⁹ M [Cl⁻] = 0.125 M
[NO₃⁻] = (0.050 L)(0.10 M)/0.200 L = 0.025 M
[Ca²⁺] = (0.150 L)(0.10 M)/0.200 L = 0.075 M

25. What volume of 0.10 M CaBr₂ must be added to 100.0 mL of 0.10 M of Pb(NO₃)₂, before a precipitate of PbBr₂ ($K_{sp}=1.4 \times 10^{-8}$) starts to form. Assume the total volume remains at 100.0 mL. [0.19 mL]



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

$$1.4 \times 10^{-8} = (0.10 \text{ M})[\text{Br}^-]^2$$

$$[\text{Br}^-] = 3.7 \times 10^{-4}$$

To determine the volume of 0.10 M CaBr₂

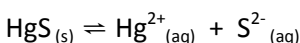
$$C_1 V_1 = C_2 V_2$$

$$V_1 = C_2 V_2 / C_1$$

$$V_1 = (3.7 \times 10^{-4})(0.100 \text{ L}) / 2(0.10 \text{ M})$$

$$V_1 = 1.9 \times 10^{-4} \text{ L} = \mathbf{0.19 \text{ mL}}$$

26. How many litres of water at 25°C must be added to mercury(II) sulfide ($K_{sp}=3.0 \times 10^{-54}$) in order for 1 mercury atom to be present in solution? (HINT: You will need Avogadro's number) [1000 L]



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

$$3.0 \times 10^{-54} = x^2$$

$$x = 1.7 \times 10^{-27}$$

$$[\text{Hg}^{2+}] = 1.7 \times 10^{-27}$$

$$n \text{ Hg}^{2+} = 1 \text{ atom} / N_A (6.02 \times 10^{23})$$

$$= 1.7 \times 10^{-24} \text{ mol}$$

$$c = n / V$$

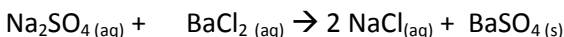
$$V = n / c$$

$$V = 1.7 \times 10^{-24} \text{ mol} / 1.7 \times 10^{-27}$$

$$\mathbf{V = 1.0 \times 10^3 \text{ L}}$$

Name: _____

27. A 100.0 mL sample of 1.00 M Na_2SO_4 is added to 200.0 mL of 1.00 M BaCl_2 . Determine the mass of BaSO_4 that precipitates from solution and the concentration of all ions at equilibrium. [$m=23.3$ g, $\text{Ba}^{2+}=0.333$ M, $\text{Cl}^-=1.33$ M, $\text{SO}_4^{2-}=3.33 \times 10^{-10}$ M, $\text{Na}^+=0.667$ M]

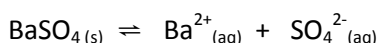


c	1.00	1.00
V	0.100	0.200
n	0.100	0.200
LF	$0.100/1=0.100$	$0.200/2=0.100$

Limiting

Excess $\text{BaCl}_2 = 0.200 - 0.100$ (amount used)
 $= 0.100$ mol

Initial $[\text{Ba}^{2+}] = 0.100 \text{ mol} / 0.300 \text{ L} = 0.33 \text{ M}$



MR	1	1	1
I	---	0.33	0
C	---	+x	+x
E	---	$0.33+x$	x

$K_{sp} = 1.1 \times 10^{-10}$
 $0.33/K_{sp} > 1000$
 Approx works

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

$$1.1 \times 10^{-10} = (0.33+x)(x)$$

$$x = 3.3 \times 10^{-10}$$

To find mass of pptn of BaSO_4 made = 0.100 moln dissolved = $(3.3 \times 10^{-10} \text{ M})(0.300 \text{ L}) = 9.9 \times 10^{-11} \text{ mol}$

$$n \text{ of ppt} = 0.100 - 9.9 \times 10^{-11} \text{ mol}$$

$$= 0.100 \text{ mol}$$

$$m = n \text{ MM}$$

$$= (0.100 \text{ mol})(233.39 \text{ g/mol})$$

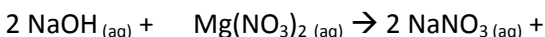
$$= 23.3 \text{ g}$$

$$[\text{Ba}^{2+}] = 0.33 \text{ M} \quad [\text{SO}_4^{2-}] = 3.3 \times 10^{-10} \text{ M}$$

$$[\text{Cl}^-] = 2(0.20 \text{ L})(1.00 \text{ M}) / 0.300 \text{ L} = 1.33 \text{ M}$$

$$[\text{Na}^+] = 2(1.00 \text{ M})(0.100 \text{ L}) / 0.300 \text{ L} = 0.67 \text{ M}$$

28. A 1.50 L sample of 0.250 M NaOH is added to 1.00 L of 0.150 M $\text{Mg}(\text{NO}_3)_2$. Calculate the mass of $\text{Mg}(\text{OH})_2$ that precipitates and the concentration of all ions in solution at equilibrium. [$m=8.75$ g, $\text{Mg}^{2+}=6.22 \times 10^{-9}$ M, $\text{OH}^-=0.0300$ M, $\text{NO}_3^-=0.120$ M, $\text{Na}^+=0.150$ M]

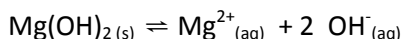


c	0.250	0.150
V	1.50 L	1.00
n	0.375	0.150
LF	$0.375/2=0.188$	$0.150/1=0.150$

Limiting

Excess $\text{NaOH} = 0.375 - 0.300$ (amount used)
 $= 0.075$ mol

Initial $[\text{OH}^-] = 0.075 \text{ mol} / 2.50 \text{ L} = 0.0300 \text{ M}$



MR	1	1	2
I	---	0	0.0300
C	---	+x	+2x
E	---	x	$0.0300+2x$

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

$$5.6 \times 10^{-12} = (x)(0.0300+2x)^2$$

$$x = 6.22 \times 10^{-9}$$

To find mass of pptn of $\text{Mg}(\text{OH})_2$ made = 0.150 moln dissolved = $(6.22 \times 10^{-9} \text{ M})(2.50 \text{ L}) = 1.6 \times 10^{-8} \text{ mol}$

$$n \text{ of ppt} = 0.150 - 1.6 \times 10^{-8} \text{ mol}$$

$$= 0.150 \text{ mol}$$

$$m = n \text{ MM}$$

$$= (0.150 \text{ mol})(58.33 \text{ g/mol})$$

$$= 8.75 \text{ g}$$

$$[\text{Mg}^{2+}] = 6.22 \times 10^{-9} \text{ M} \quad [\text{OH}^-] = 0.0300 \text{ M}$$

$$[\text{NO}_3^-] = 2(1.00 \text{ L})(0.150 \text{ M}) / 2.50 \text{ L} = 0.120 \text{ M}$$

$$[\text{Na}^+] = (1.50 \text{ L})(0.250 \text{ M}) / 2.50 \text{ L} = 0.150 \text{ M}$$

$K_{sp} = 5.6 \times 10^{-12}$
 $0.0300/K_{sp} > 1000$
 Approx works