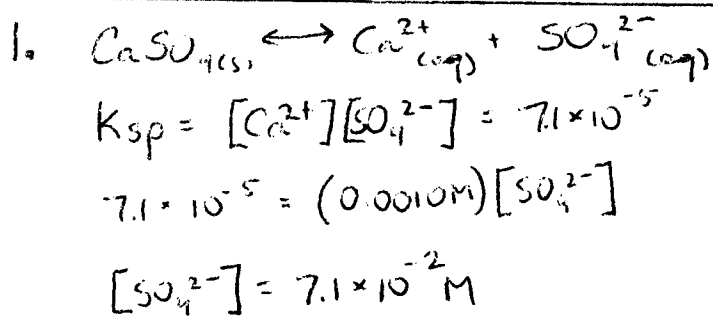


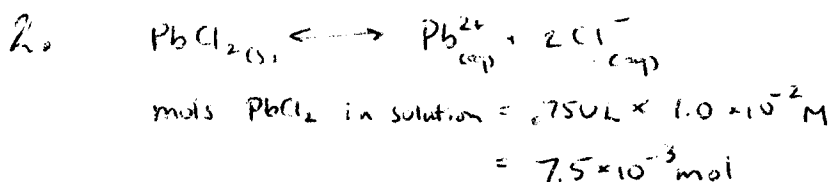
Answer Key.

Solubility #5

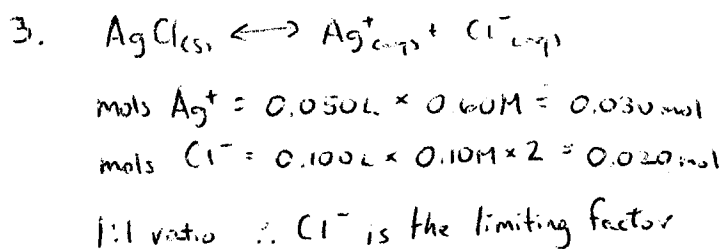
1. What is the minimum mass of Na_2SO_4 crystal that must be dissolved in 5.0 L of 0.0010 M $\text{Ca}(\text{NO}_3)_2$ solution in order to initiate precipitation of calcium sulphate?
2. 750 mL of a 1.0×10^{-2} M lead II chloride solution was cooled resulting in the precipitation of 1.80 g of solid PbCl_2 . What is the molarity of the cooled solution?
3. Determine the mass of silver chloride precipitated when 50 mL of 0.60 M AgNO_3 solution is mixed with 100 mL of 0.10 M CaCl_2 solution. Assume AgCl has negligible solubility.
4. A given sample of water with temporary hardness has a $[\text{Ca}^{2+}]$ of 1.0×10^{-3} M.
(a) If the K_{sp} of CaF_2 is 1.7×10^{-10} , what is the maximum $[\text{F}^-]$ that can be attained in temporary hard water before CaF_2 would precipitate?
5. A solution is prepared by adding 1.5 mol BaCrO_4 to water to make 1.0 L of solution. Calculate the $[\text{Ba}^{2+}]$ and the $[\text{CrO}_4^{2-}]$ at equilibrium.
6. The $[\text{Ag}^+]$ of a solution is 4.0×10^{-3} M. Calculate the $[\text{Cl}^-]$ that must be exceeded before AgCl can precipitate.
7. Calculate the solubility of PbI_2 .
8. 30 mL of 0.10 M AgNO_3 is added to 70 mL of 0.10 M CaCl_2 . What is the concentration of each ion in the solution once precipitation stops and equilibrium is established?



$\text{mols } \text{SO}_4^{2-} = 7.1 \times 10^{-2} \text{ mol/L} \times 5.0 \text{ L} = 0.355 \text{ mol}$
 $\text{FW } \text{Na}_2\text{SO}_4 = 2(23.0) + 32.1 + 4(16.0) = 142.1 \text{ g/mol}$
 $\text{mass } \text{Na}_2\text{SO}_4 = 0.355 \text{ mol} \times 142.1 \text{ g/mol} = 50.9 \text{ g}$

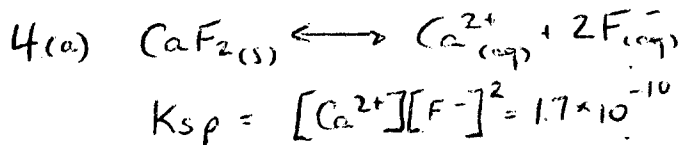


$\text{FW } \text{PbCl}_2 = 207.2 + 2(35.5) = 278.2 \text{ g/mol}$
 $\text{mols } \text{PbCl}_2 \text{ ppt} = 1.80 \text{ g} \div 278.2 \text{ g/mol} = 6.47 \times 10^{-3} \text{ mol}$
 $\text{mols left in sol}^n = 7.5 \times 10^{-3} - 6.47 \times 10^{-3} = 1.03 \times 10^{-3} \text{ mol}$
 $[\text{PbCl}_2] = \frac{1.03 \times 10^{-3} \text{ mol}}{0.750 \text{ L}} = 1.4 \times 10^{-3} \text{ M}$



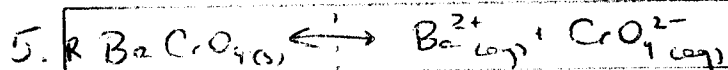
0.020 mol of precipitate will form
 $\text{FW } \text{AgCl} = 107.9 + 35.5 = 143.4 \text{ g/mol}$
 $\text{mass of } \text{AgCl} = 0.020 \text{ mol} \times 143.4 \text{ g/mol}$
 $= 2.9 \text{ g}$

Solubility #5 - Answer key



$$1.7 \times 10^{-10} = (1.0 \times 10^{-3})[\text{F}^{-}]^2$$

$$[\text{F}^{-}] = 4.1 \times 10^{-4} \text{ M}$$



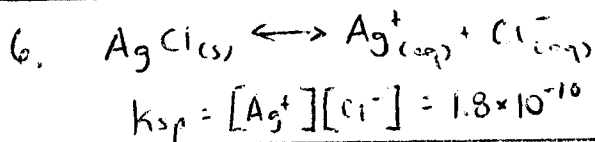
i	0.00	0.00
c	+x	+x
e	x	x

$$K_{sp} = [\text{Ba}^{2+}][\text{CrO}_4^{2-}] = 1.2 \times 10^{-10}$$

$$1.2 \times 10^{-10} = x^2$$

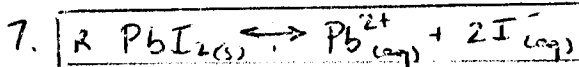
$$x = 1.1 \times 10^{-5} \text{ M}$$

The $[\text{Ba}^{2+}]$ and $[\text{CrO}_4^{2-}]$ are $1.1 \times 10^{-5} \text{ M}$ at equilibrium



$$1.8 \times 10^{-10} = (4.0 \times 10^{-3})[\text{Cl}^{-}]$$

$$[\text{Cl}^{-}] = 4.5 \times 10^{-8} \text{ M}$$



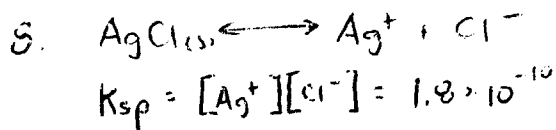
i	0.00	0.00
c	+x	+2x
e	x	2x

$$K_{sp} = [\text{Pb}^{2+}][\text{I}^{-}]^2 = 8.5 \times 10^{-9}$$

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

$$= 4x^3$$

$$x = 1.3 \times 10^{-3} \text{ M} = \text{solubility of PbI}_2$$



$$[\text{Ag}^{+}] = \frac{30 \text{ mL} \times 0.10 \text{ M}}{100 \text{ mL}} = 0.030 \text{ M}$$

$$[\text{Cl}^{-}] = \frac{70 \text{ mL} \times 0.10 \text{ M} \times 2}{100 \text{ mL}} = 0.140 \text{ M}$$

If all the Ag^{+} precipitates $[\text{Cl}^{-}] = 0.140 \text{ M} - 0.030 \text{ M} = 0.11 \text{ M}$

then at eqm:

$$[\text{Ag}^{+}] = x$$

$$[\text{Cl}^{-}] = 0.11 \text{ M} + x \approx 0.11 \text{ M} \text{ if } x \text{ is small}$$

$$1.8 \times 10^{-10} = x(0.11 \text{ M})$$

$$x = 1.64 \times 10^{-9} \text{ M}$$

at eqm:

$$[\text{Ag}^{+}] = 1.6 \times 10^{-9} \text{ M}$$

$$[\text{Cl}^{-}] = 0.11 \text{ M}$$