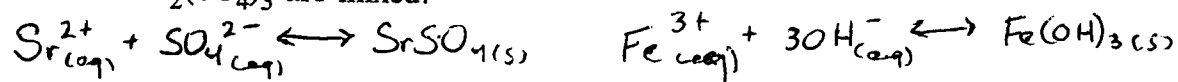


Describe each of the following compounds as soluble, or of low solubility

a.	CuBr	low
b.	CuCl ₂	
c.	NaOH	
d.	Na ₃ PO ₄	
e.	AgOH	low
f.	MgS	
g.	Sr(OH) ₂	
h.	PbSO ₄	low
i.	CdSO ₄	
j.	Rb ₂ SO ₄	
k.	CuOH	low
l.	BaSO ₄	low
m.	NH ₄ OH	
n.	Fr ₂ S	
o.	CuSO ₄	
p.	CrS	low
q.	(NH ₄) ₂ Cr ₂ O ₇	
r.	K ₂ CrO ₄	
s.	NaCl	
t.	KNO ₃	
u.	NH ₄ NO ₃	
v.	HCOOH	
w.	PbCO ₃	low
x.	PbNO ₃	
y.	Na ₂ SO ₃	

Solubility #1

1. Write the net ionic equation(s) for the reaction(s) when equal volumes of 0.2 M $\text{Sr}(\text{OH})_2$ and 0.2 M $\text{Fe}_2(\text{SO}_4)_3$ are mixed.



2. Describe the equilibrium that exists in a saturated solution of BaSO_4 in contact with some solid residue of BaSO_4 .

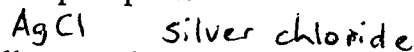


at equilibrium the $[\text{Ba}^{2+}]$ and $[\text{SO}_4^{2-}]$ are constant

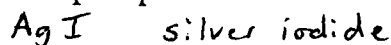
3. In an experiment, 0.1 M AgNO_3 is added to 0.1 M NaCl , resulting in the formation of a white precipitate. When 0.1 M NaI is added to this mixture, the white precipitate dissolves and a yellow precipitate forms.

- (a) Write the formula and name for each of the following:

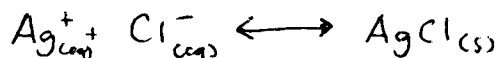
- (i) the white precipitate



- (ii) the yellow precipitate

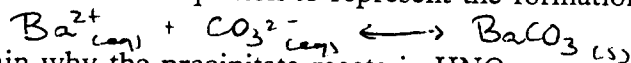


- (b) Write the net ionic equation to represent the formation of the more soluble precipitate.



4. A solution of $\text{Ba}(\text{NO}_3)_2$ is added to a solution of Na_2CO_3 , resulting in the formation of a white precipitate the reacts with HNO_3 .

- (a) Write a net ionic equation to represent the formation of the white precipitate.



- (b) Explain why the precipitate reacts in HNO_3 .



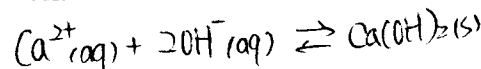
carbonates react with acids.

5. A student is given three beakers, each containing 100 mL of solution. The first beaker contains 0.20 M CaS ; the second contains 0.20 M CuSO_4 ; the third contains 0.20 M $\text{Sr}(\text{OH})_2$. The student is asked to select two solutions which, when combined, would result in the formation of a mixture containing a single precipitate.

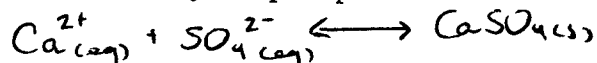
- (a) Which two solutions should the student use?



- (b) Write the net ionic equation for the precipitation reaction.



6. If a solution of calcium nitrate is added to a saturated solution of calcium sulphate, a precipitate is observed to form. Explain why this occurs, including any relevant equation(s) and identify the precipitate.



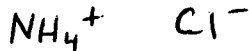
Adding $\text{Ca}(\text{NO}_3)_2$ increases the $[\text{Ca}^{2+}]$ causing the equilibrium to shift to the right and to form more solid product.

Key.

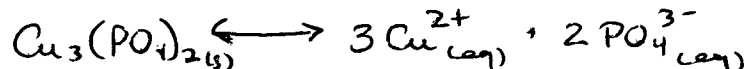
7. Write an equation that describes a saturated solution of NaCl.



8. A solution containing Al^{3+} , NH_4^+ and Mg^{2+} ions is added to a solution containing S^{2-} , Cl^- and OH^- ions. Identify the ions that do **not** form a precipitate.



9. Write an equation that describes the equilibrium present in a saturated solution of $\text{Cu}_3(\text{PO}_4)_2$.

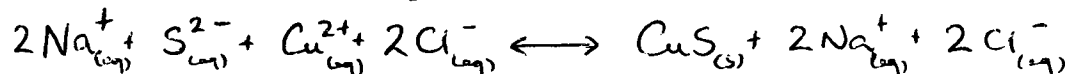


10. A 1.0 M solution of sodium sulphide is added to a 1.0 M solution of copper(II) chloride resulting in the formation of a precipitate.

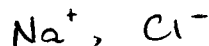
- (a) Identify the precipitate.



- (b) Write the **complete** ionic equation for the reaction.



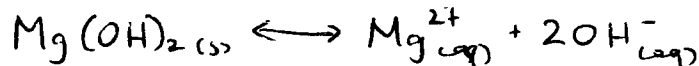
- (c) Identify all spectator ions.



11. Identify a salt that could be added to a saturated solution of BaSO_4 that would result in more solid barium sulphate forming.



12. Write an equation that describes a saturated solution of magnesium hydroxide.



13. When solid $\text{Ca}(\text{NO}_3)_2$ is added to a saturated solution of MgCO_3 , more MgCO_3 dissolves. Explain.

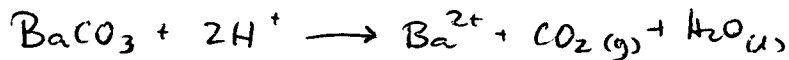
The Ca^{2+} reacts with the CO_3^{2-} to form solid CaCO_3 . This decreases the $[\text{CO}_3^{2-}]$ and more MgCO_3 dissolves.

14. A student mixes equal volumes of 0.20 M Na_2CO_3 and 0.20 M $\text{Ba}(\text{NO}_3)_2$, forming a white precipitate.

- (a) Write the net ionic equation for the precipitation reaction.



- (b) Explain why the precipitate dissolves when HCl is added.



Carbonates react with acid to form CO_2 and water.

15. Write an equation that represents the solubility equilibrium of a saturated solution of barium sulphate.



Solubility #2

- A solution contains SO_4^{2-} and Cl^- . Outline an experimental procedure to remove each ion individually from the solution, and identify the reagents used in the procedure.

 - ① Add either $\text{Ca}(\text{NO}_3)_2$ or $\text{Sr}(\text{NO}_3)_2$
or $\text{Ba}(\text{NO}_3)_2$ allow it to settle \rightarrow ppt's SO_4^{2-}
 - ② Add CuNO_3 to ppt Cl^- . Allow it to settle out
- A solution is known to contain Cu^+ , Be^{2+} and Sr^{2+} ions, each at a concentration of 0.20 M.

 - What compound could be used to precipitate the Sr^{2+} while leaving the other two cations in solution?
 Na_2SO_4
 - Write the net ionic equation for the reaction.
$$\text{Sr}^{2+}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) \rightleftharpoons \text{SrSO}_4(\text{s})$$
- A solution contains Pb^{2+} , Mg^{2+} , and Sr^{2+} . Outline a procedure to isolate the precipitate SrSO_4 from this solution.

 - 1) Use NaCl to ppt Pb^{2+} as PbCl_2
 - 2) Use Na_2SO_4 to ppt SrSO_4

or

 - 1) Use KOH to ppt both Pb^{2+} and Mg^{2+} as $\text{Pb}(\text{OH})_2$ and $\text{Mg}(\text{OH})_2$
 - 2) Use Na_2SO_4 to ppt SrSO_4
- Write the formula of two materials (other than water) that could be added to a saturated solution of Ag_2CO_3 to increase the amount of Ag_2CO_3 that will dissolve.

$\text{Ag}_2\text{CO}_3 \rightleftharpoons 2\text{Ag}^+ + \text{CO}_3^{2-}$ or halide ion $\text{Cl}^- + \text{Ag}^+ \rightleftharpoons \text{AgCl}$
add something to remove Ag^+ or CO_3^{2-} eg. Acid $\rightarrow \text{CO}_3^{2-} + 2\text{H}^+ \rightleftharpoons \text{CO}_2 + \text{H}_2\text{O}$
- A solution contains Ca^{2+} , Sr^{2+} and Pb^{2+} ions that must be separated.

 - Identify an anion that could be used to remove **only** the lead ion by precipitation.
 Cl^- , I^- , Br^- , S^{2-}
 - Identify an anion that could be used to separate Ca^{2+} from Sr^{2+} .
 OH^-

Key.

6. Use the table of solubilities to determine a scheme that allows the separation of Ba^{2+} , Cu^{2+} , and Br^- from each other.

- ① Add SO_4^{2-} to ppt Ba^{2+}
- ② Add S^{2-} to ppt Cu^{2+}
- ③ Add Ag^+ to ppt Br^-

7. Use the table of solubilities to describe how you would separate Mg^{2+} , Ba^{2+} and Ag^+ from each other.

- ① Add $\text{NaCl} \rightarrow \text{Ag}^+ + \text{Cl}^- \rightleftharpoons \text{AgCl}$
- ② Add $\text{Na}_2\text{SO}_4 \rightarrow \text{Ba}^{2+} + \text{SO}_4^{2-} \rightleftharpoons \text{BaSO}_4$
- ③ Add $\text{NaOH} \rightarrow \text{Mg}^{2+} + 2\text{OH}^- \rightleftharpoons \text{Mg}(\text{OH})_2$

8. Use the table of solubilities to outline a scheme to separate a mixture of Li^+ , Ag^+ , Sr^{2+} and Fe^{3+} from each other.

- ① Add $\text{NaCl} \rightarrow \text{Ag}^+ + \text{Cl}^- \rightleftharpoons \text{AgCl}$
- ② Add $\text{Na}_2\text{S} \rightarrow 2\text{Fe}^{3+} + 3\text{S}^{2-} \rightleftharpoons \text{Fe}_2\text{S}_3$
- ③ Add $\text{Na}_2\text{SO}_4 \rightarrow \text{Sr}^{2+} + \text{SO}_4^{2-} \rightleftharpoons \text{SrSO}_4$

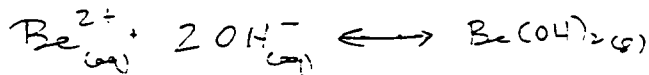
Li^+ is left.

9. A beaker contains OH^- ions and S^{2-} ions in solution both at a concentration of 0.10 M. You are asked to precipitate the OH^- ions while leaving the S^{2-} ions in solution.

(a) Which reagent could you use?



(b) Write the net ionic equation for the precipitation reaction.



Key

Solubility #3 - Assume that all solutions are at 25°C unless told otherwise.

1. Which one of the following solutions will contain the greatest silver ion concentration?

^{25°C}
 *saturated Ag_2CO_3 saturated AgCl ^{Ag_2CO_3}
 $[\text{Ag}^+] = 2.57 \times 10^{-4} \text{ M}$ $[\text{Ag}^+] = 1.34 \times 10^{-5} \text{ M}$

2. In which solution would you find the highest magnesium ion concentration: saturated MgCO_3 or saturated Mg(OH)_2 ?

$[\text{Mg}^{2+}] = 2.61 \times 10^{-3} \text{ M}$
 $[\text{Mg}^{2+}] = 1.11 \times 10^{-4} \text{ M}$

3. The K_{sp} values for various silver salts are given below:

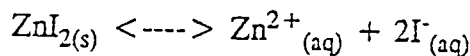
$[\text{Ag}^+] = 7.3 \times 10^{-5} \text{ M}$ ← x^2	AgBrO_3	5.3×10^{-5}	$[\text{Ag}^+] = 1.34 \times 10^{-5} \text{ M}$	$4x^3$	Ag_2CO_3	8.5×10^{-12}
$[\text{Ag}^+] = 7.4 \times 10^{-3} \text{ M}$ ← $4x^3$	<u>$\text{Ag}_2\text{Cr}_2\text{O}_7$</u>	2.0×10^{-7}		x^2	AgSCN	1.2×10^{-12}
	x^2	AgOH	1.5×10^{-10}	x^2	AgBr	5.4×10^{-13}
	x^2	AgCl	1.8×10^{-10}	x^2	AgI	8.5×10^{-17}
	$4x^3$	Ag_2CrO_4	1.1×10^{-12}	x^2	AgIO_3	3.2×10^{-8}

Determine which one of these compounds gives the highest silver ion concentration in a saturated solution and calculate this concentration.

- Calculate the solubility of BaSO_4 . $1.0 \times 10^{-5} \text{ M}$
- Calculate the mass of solid PbSO_4 that can be dissolved in 5.0 L of solution at 25°C.
- Calculate the solubility of PbI_2 at 25°C. 0.20 g
 $1.3 \times 10^{-3} \text{ M}$
- What is the solubility of calcium oxalate at 25°C in g/L? $6.1 \times 10^{-3} \text{ g/L}$
- How many milligrams of CaCO_3 would be dissolved in 1.0 L of saturated solution at 25°C. 7.1 mg/L
- What mass of PbI_2 will dissolve in 250 mL of water?
- What is the solubility of magnesium hydroxide? 0.15 g
 $1.11 \times 10^{-4} \text{ M}$
 - What is the $[\text{OH}^-]$ in a saturated solution of magnesium hydroxide?
 $2.22 \times 10^{-4} \text{ M}$

Solubility #4

- When excess Ag_2CO_3 solid is shaken with 1.00 L of 2.00 M K_2CO_3 , it is determined that 6.00×10^{-6} mol Ag_2CO_3 solid dissolves. Calculate the solubility product for silver carbonate.
- Calculate the K_{sp} for SrF_2 if the solubility is 0.122 g/L.
- A saturated solution of calcium hydroxide is found to have $[\text{OH}^-]$ of 2.09×10^{-2} M. Calculate the K_{sp} for $\text{Ca}(\text{OH})_2$.
- The solubility of Ag_2SO_4 is 0.62 g/L at 6°C . What is the K_{sp} at this temperature?
- Describe an analytical method which could be used to determine the K_{sp} of a saturated solution of silver sulphate other than drying the solid from a saturated solution.
- At 20°C the solubility of PbF_2 is 64 mg per 100 mL of solution. Determine the K_{sp} for lead II fluoride. $64 \text{ g/L} \times \frac{1 \text{ mol}}{245.2 \text{ g}} = 2.61 \times 10^{-3} \text{ mol/L}$ $[\text{Pb}^{2+}] = 2.61 \times 10^{-3} \text{ M}$
 $K_{sp} = [\text{Pb}^{2+}][\text{F}^-]^2 = 7.1 \times 10^{-8}$ $[\text{F}^-] = 5.22 \times 10^{-3} \text{ M}$
- The equilibrium in a saturated ZnI_2 solution is given by:



Predict the effect on the solubility of ZnI_2 of adding some:

- solid NaI. Solubility of ZnI_2 will decrease
- solid $\text{Zn}(\text{NO}_3)_2$ decrease
- solid NaOH $\text{Zn}(\text{OH})_2$ is produced; solubility will increase
- concentrated HCl no effect
- solid NH_4NO_3 no effect

#1 $[\text{Ag}^+] = \frac{2(6.00 \times 10^{-6} \text{ mol})}{1.00 \text{ L}} = 1.20 \times 10^{-5} \text{ M}$ $[\text{CO}_3^{2-}] = 2.00 \text{ M} + 6.00 \times 10^{-6} \text{ mol/L} = 2.00 \text{ M}$

$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}] = 2.88 \times 10^{-10}$$

#2 $0.122 \text{ g/L} \times \frac{1 \text{ mol}}{125.6 \text{ g}} = 9.71 \times 10^{-4} \text{ mol/L}$ $[\text{Sr}^{2+}] = 9.71 \times 10^{-4} \text{ M}$
 $[\text{F}^-] = 2(9.71 \times 10^{-4} \text{ M}) = 1.94 \times 10^{-3} \text{ M}$

$$K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 = 3.67 \times 10^{-9}$$

#3 $[\text{OH}^-] = 2.09 \times 10^{-2} \text{ M}$ $[\text{Ca}^{2+}] = \frac{2.09 \times 10^{-2} \text{ M}}{2} = 1.05 \times 10^{-2} \text{ M}$

$$K_{sp} = [\text{Ca}^{2+}][\text{OH}^-]^2 = 4.57 \times 10^{-6}$$

#4 $0.62 \text{ g/L} \times \frac{1 \text{ mol}}{311.9 \text{ g}} = 1.99 \times 10^{-3} \text{ mol/L}$ $[\text{Ag}^+] = 2(1.99 \times 10^{-3} \text{ M}) = 3.98 \times 10^{-3} \text{ M}$
 $[\text{SO}_4^{2-}] = 1.99 \times 10^{-3} \text{ M}$

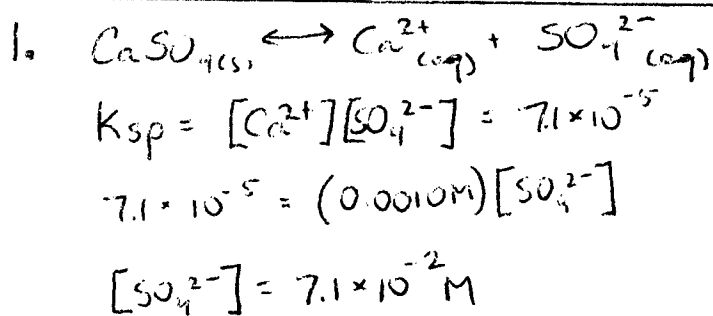
$$K_{sp} = [\text{Ag}^+]^2 [\text{SO}_4^{2-}] = 3.1 \times 10^{-8}$$

#5 Titrate a known volume of saturated Ag_2SO_4 solution with standardized NaCl to the stoichiometric point as determined by an indicator. The moles of Cl^- added = moles Ag^+ initially present. Then, $K_{sp} = [\text{Ag}^+]^2 [\text{SO}_4^{2-}]$

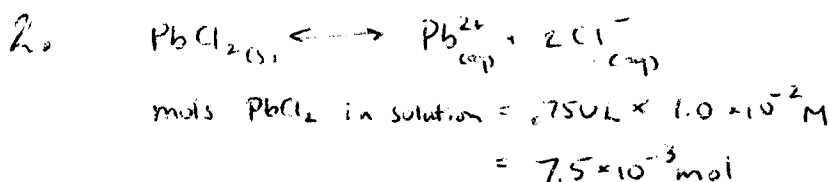
Answer Key.

Solubility #5

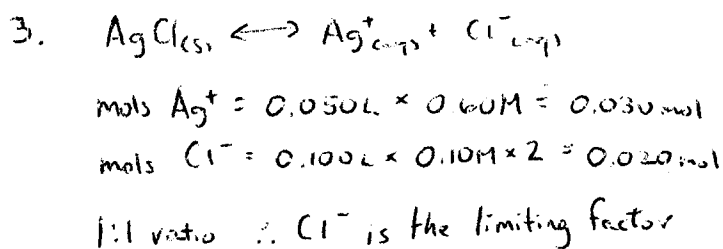
- What is the minimum mass of Na_2SO_4 crystal that must be dissolved in 5.0 L of 0.0010 M $\text{Ca}(\text{NO}_3)_2$ solution in order to initiate precipitation of calcium sulphate?
- 750 mL of a 1.0×10^{-2} M lead II chloride solution was cooled resulting in the precipitation of 1.80 g of solid PbCl_2 . What is the molarity of the cooled solution?
- 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.
- 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?
- 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.
- The $[\text{Ag}^+]$ of a solution is 4.0×10^{-3} M. Calculate the $[\text{Cl}^-]$ that must be exceeded before AgCl can precipitate.
- Calculate the solubility of PbI_2 .
- 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?



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

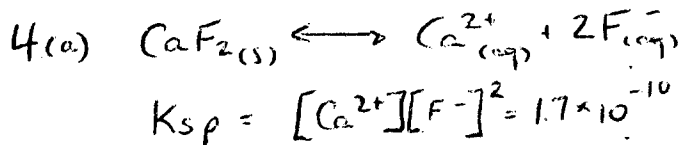


FW $\text{PbCl}_2 = 207.2 + 2(35.5) = 278.2 \text{ g/mol}$
 mols PbCl_2 ppt = $1.80 \text{ g} \div 278.2 \text{ g/mol} = 6.47 \times 10^{-3} \text{ mol}$
 mols left in solⁿ = $7.5 \times 10^{-3} - 6.47 \times 10^{-3} = 1.03 \times 10^{-3} \text{ mol}$
 $[\text{PbCl}_2] = \frac{1.03 \times 10^{-3} \text{ mol}}{0.750 \text{ L}} = 1.4 \times 10^{-3} \text{ M}$



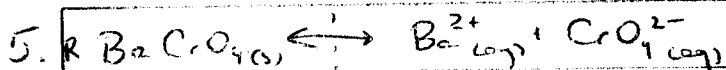
0.020 mol of precipitate will form
 FW $\text{AgCl} = 107.9 + 35.5 = 143.4 \text{ g/mol}$
 mass of $\text{AgCl} = 0.020 \text{ mol} \times 143.4 \text{ g/mol}$
 $= 2.9 \text{ g}$

Solubility #5 - Answer key



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

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



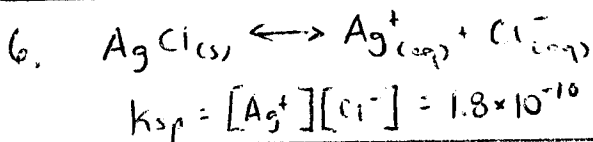
i	0.00	0.00
c	+x	+x
e	x	x

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

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

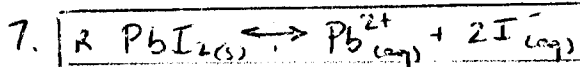
$$x = 1.1 \times 10^{-5} \text{ M}$$

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



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

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



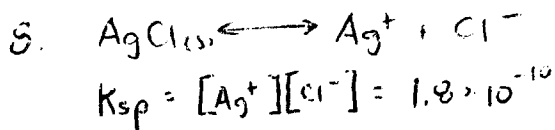
i	0.00	0.00
c	+x	+2x
e	x	2x

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

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

$$= 4x^3$$

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



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

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

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

then at eqm:

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

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

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

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

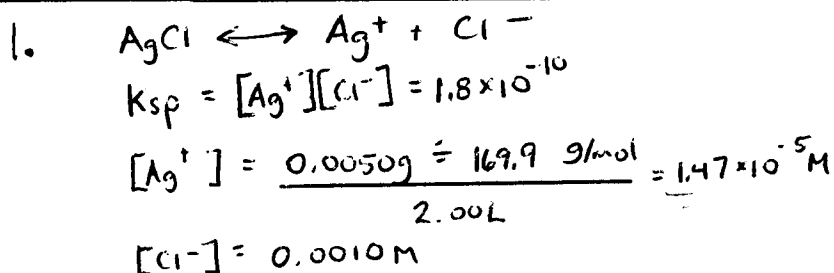
at eqm:

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

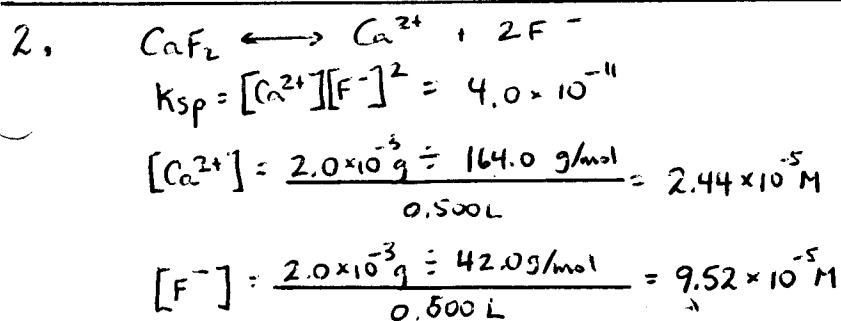
$$[\text{Cl}^{-}] = 0.11 \text{ M}$$

Solubility #6

1. Show whether or not a precipitate would be expected to form when 0.0050 g AgNO_3 crystals are added to 2.00 L of 0.0010 M NaCl .
2. 2.0 mg of $\text{Ca}(\text{NO}_3)_2$ and 2.0 mg NaF are dissolved and made up to 500 mL of solution. If the K_{sp} for CaF_2 is 4.0×10^{-11} , will a precipitate form?
3. Will a precipitate of AgCl form when 5.1 mg of AgNO_3 crystals are added to 3.0 L of 2.0×10^{-3} M NaCl ?
4. Show whether or not a precipitate of silver acetate forms when 15 mL of 1.0 M AgNO_3 is added to 45 mL of acetic acid in which the $[\text{CH}_3\text{COO}^-]$ is 5.2×10^{-3} M.
 $K_{sp} \text{CH}_3\text{COOAg} = 3.7 \times 10^{-3}$
5. Determine whether or not a precipitate of BaSO_4 will form when 0.15 g of K_2SO_4 solid is added to 2.0 L of 1.7×10^{-5} M BaCl_2 .
6. Explain why a precipitate of silver chloride will not be produced when 20 mL of 3.0×10^{-6} M AgNO_3 is mixed with 30 mL of 1.0×10^{-4} M NaCl .
7. When AgNO_3 crystals dissolve in a solution containing 0.010 M NaCl and 0.010 M Na_2CrO_4 , AgCl precipitates before the Ag_2CrO_4 . Explain this behavior.
8. A 0.010 M solution of AgNO_3 is added dropwise to a solution containing a mixture of carbonate and iodate ions, in which $[\text{CO}_3^{2-}] = 3.0 \times 10^{-3}$ M and $[\text{IO}_3^-] = 5.0 \times 10^{-3}$ M. Which substance precipitates first?
9. Will a precipitate of $\text{Al}(\text{OH})_3$ form when 0.50 L of 2.0×10^{-3} M AlCl_3 and 0.50 L of 4.0×10^{-2} M NaOH are mixed and diluted to 1000 L with water? $K_{sp} \text{Al}(\text{OH})_3 = 3.7 \times 10^{-15}$
10. Will a precipitate form when 400 mL of 0.0020 M $\text{Ba}(\text{OH})_2$ are mixed with 200 mL of 0.0020 M H_2SO_4 ?



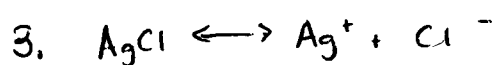
TIP = $(1.47 \times 10^{-5})(0.0010)$
 $= 1.5 \times 10^{-8}$
 Since TIP > K_{sp} a ppt will form.



TIP = $(2.44 \times 10^{-5})(9.52 \times 10^{-5})^2$
 $= 2.2 \times 10^{-13}$
 Since TIP < K_{sp} no ppt will form.

Solubility #6

Answer key



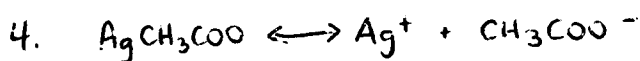
$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

$$[\text{Ag}^+] = \frac{5.1 \times 10^{-3} \text{ g} \div 169.9 \text{ g/mol}}{3.0 \text{ L}} = 1.00 \times 10^{-5} \text{ M}$$

$$[\text{Cl}^-] = 2.0 \times 10^{-3} \text{ M}$$

$$\begin{aligned} \text{TIP} &= (1.00 \times 10^{-5})(2.0 \times 10^{-3}) \\ &= 2.0 \times 10^{-8} \end{aligned}$$

TIP > K_{sp} a ppt will form.



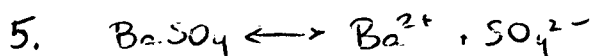
$$K_{sp} = [\text{Ag}^+][\text{CH}_3\text{COO}^-] = 3.7 \times 10^{-3}$$

$$[\text{Ag}^+] = \frac{15 \text{ mL} \times 1.0 \text{ M}}{15 \text{ mL} + 45 \text{ mL}} = 0.25 \text{ M}$$

$$[\text{CH}_3\text{COO}^-] = \frac{45 \text{ mL} \times 5.2 \times 10^{-3} \text{ M}}{45 \text{ mL} + 15 \text{ mL}} = 3.9 \times 10^{-3} \text{ M}$$

$$\begin{aligned} \text{TIP} &= (0.25)(3.9 \times 10^{-3}) \\ &= 9.8 \times 10^{-4} \end{aligned}$$

TIP < K_{sp} no ppt will form.



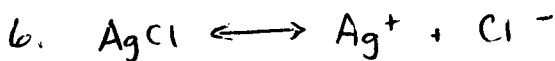
$$K_{sp} = [\text{Ba}^{2+}][\text{SO}_4^{2-}] = 1.1 \times 10^{-10}$$

$$[\text{Ba}^{2+}] = 1.7 \times 10^{-5} \text{ M}$$

$$[\text{SO}_4^{2-}] = \frac{0.15 \text{ g} \div 174.3 \text{ g/mol}}{2.0 \text{ L}} = 4.3 \times 10^{-4} \text{ M}$$

$$\begin{aligned} \text{TIP} &= (1.7 \times 10^{-5})(4.3 \times 10^{-4}) \\ &= 7.3 \times 10^{-9} \end{aligned}$$

TIP > K_{sp} a ppt will form



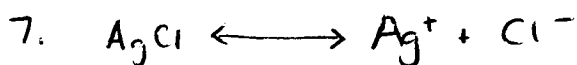
$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

$$[\text{Ag}^+] = \frac{20 \text{ mL} \times 3.0 \times 10^{-6} \text{ M}}{20 \text{ mL} + 30 \text{ mL}} = 1.2 \times 10^{-6} \text{ M}$$

$$[\text{Cl}^-] = \frac{30 \text{ mL} \times 1.0 \times 10^{-4} \text{ M}}{20 \text{ mL} + 30 \text{ mL}} = 6.0 \times 10^{-5} \text{ M}$$

$$\begin{aligned} \text{TIP} &= (1.2 \times 10^{-6})(6.0 \times 10^{-5}) \\ &= 7.2 \times 10^{-11} \end{aligned}$$

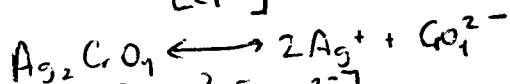
TIP < K_{sp} no ppt will form



$$K_{sp} = [\text{Ag}^+][\text{Cl}^-] = 1.8 \times 10^{-10}$$

$$[\text{Cl}^-] = 0.010 \text{ M}$$

$$[\text{Ag}^+] = \frac{K_{sp}}{[\text{Cl}^-]} = 1.8 \times 10^{-8} \text{ M to ppt AgCl}$$



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

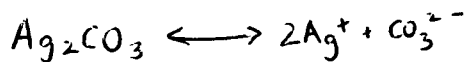
$$[\text{CrO}_4^{2-}] = 0.010 \text{ M}$$

$$[\text{Ag}^+] = \sqrt{\frac{K_{sp}}{[\text{CrO}_4^{2-}]}} = 1.05 \times 10^{-5} \text{ M to ppt Ag}_2\text{CrO}_4$$

When adding the AgNO_3 , the lower $[\text{Ag}^+]$ is reached before to higher one, thus the AgCl precipitates first.

Solubility #6

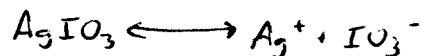
Answer key

8. For a precipitate of Ag_2CO_3 

$$K_{sp} = [\text{Ag}^+]^2 [\text{CO}_3^{2-}] = 8.5 \times 10^{-12}$$

$$[\text{CO}_3^{2-}] = 3.0 \times 10^{-3}$$

$$[\text{Ag}^+] = \sqrt{\frac{K_{sp}}{[\text{CO}_3^{2-}]}} = 5.3 \times 10^{-5} \text{ M}$$

For a precipitate of AgIO_3 

$$K_{sp} = [\text{Ag}^+][\text{IO}_3^-] = 3.2 \times 10^{-8}$$

$$[\text{IO}_3^-] = 5.0 \times 10^{-3} \text{ M}$$

$$[\text{Ag}^+] = \frac{K_{sp}}{[\text{IO}_3^-]} = 6.4 \times 10^{-6} \text{ M}$$

The AgIO_3 ppt's first9. $\text{Al}(\text{OH})_3 \rightleftharpoons \text{Al}^{3+} + 3\text{OH}^-$

$$K_{sp} = [\text{Al}^{3+}][\text{OH}^-]^3 = 3.7 \times 10^{-15}$$

$$[\text{Al}^{3+}] = \frac{0.50 \text{ L} \times 2.0 \times 10^{-3} \text{ M}}{1000 \text{ L}} = 1.0 \times 10^{-6} \text{ M}$$

$$[\text{OH}^-] = \frac{0.50 \text{ L} \times 4.0 \times 10^{-2} \text{ M}}{1000 \text{ L}} = 2.0 \times 10^{-5} \text{ M}$$

$$\begin{aligned} \text{TIP} &= (1.0 \times 10^{-6})(2.0 \times 10^{-5})^3 \\ &= 8.0 \times 10^{-21} \end{aligned}$$

TIP < K_{sp} so no ppt will form10. $\text{BaSO}_4 \rightleftharpoons \text{Ba}^{2+} + \text{SO}_4^{2-}$

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

$$[\text{Ba}^{2+}] = \frac{0.0020 \text{ M} \times 400 \text{ mL}}{400 \text{ mL} + 200 \text{ mL}} = 1.33 \times 10^{-3} \text{ M}$$

$$[\text{SO}_4^{2-}] = \frac{0.0020 \text{ M} \times 200 \text{ mL}}{400 \text{ mL} + 200 \text{ mL}} = 6.67 \times 10^{-4} \text{ M}$$

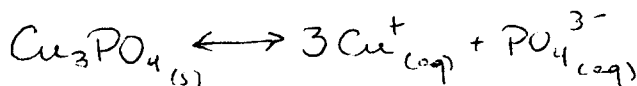
$$\begin{aligned} \text{TIP} &= (1.33 \times 10^{-3})(6.67 \times 10^{-4}) \\ &= 8.9 \times 10^{-7} \end{aligned}$$

TIP > K_{sp} so a ppt will form

Key.

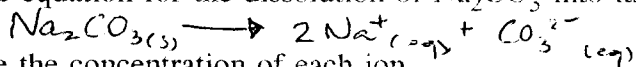
Solubility Open-ended Review

1. Write an equation that describes the equilibrium present in a saturated solution of $\text{Cu}_3(\text{PO}_4)_2$.



2. 53 g of Na_2CO_3 are dissolved in sufficient water to make 5.0 L of solution.

- A. Write the equation for the dissolution of Na_2CO_3 into its aqueous ions.



- B. Calculate the concentration of each ion.

$$\text{FW Na}_2\text{CO}_3 = 2(23.0) + 12.0 + 3(16.0) = 106 \text{ g/mol}$$

$$53 \text{ g} \div 106 \text{ g/mol} = 0.50 \text{ mol}$$

$$[\text{Na}_2\text{CO}_3] = 0.50 \text{ mol} / 5.0 \text{ L} = 0.10 \text{ M}$$

$$[\text{Na}^+] = 2(0.10 \text{ M}) = 0.20 \text{ M}$$

$$[\text{CO}_3^{2-}] = 0.10 \text{ M}$$

- C. Describe the changes in entropy and enthalpy as the Na_2CO_3 dissolves.

entropy increases as ions are formed.

enthalpy increases as the dissolving process is endothermic

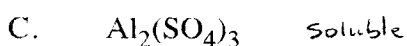
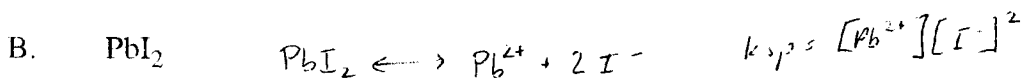
- D. When the solution was prepared, some doubt existed that Na_2SO_4 might have been used by mistake. Describe a suitable precipitation test that will confirm the presence of CO_3^{2-} ions in the solution.

Add any cation except: Alkali ions, H^+ , NH_4^+ which ppt neither

or Ag^+ , Ca^{2+} , Sr^{2+} , Ba^{2+} , Pb^{2+} which ppt both

possible answers include Al^{3+} , Fe^{2+} , Cu^{2+} etc.

3. Write an equilibrium expression and an equation that describes the equilibrium for only those salts that have low solubility:

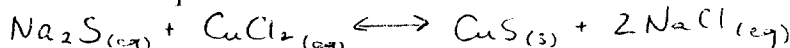


4. A 1.0 M solution of sodium sulphide is added to a 1.0 M solution of copper II chloride resulting in the formation of a precipitate.

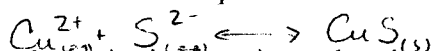
- A. Write the name and formula of the precipitate.

Copper II sulphide CuS

- B. Write the full equation for the reaction.

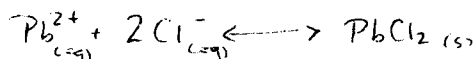


- C. Write the net ionic equation for the reaction.

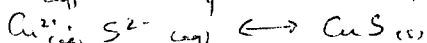
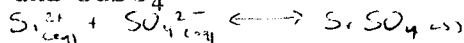


5. Write balanced net ionic equations showing the formation of each precipitate formed when equal volumes of the following 0.50 M solutions are mixed:

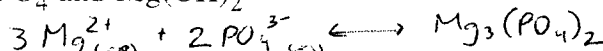
- A. MgCl_2 and $\text{Pb}(\text{NO}_3)_2$



- B. SrS and CuSO_4



- C. $(\text{NH}_4)_3\text{PO}_4$ and $\text{Mg}(\text{OH})_2$



6. Calculate the concentration of each ion in the following saturated solutions:

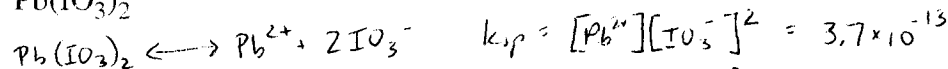
- A. $\text{Al}(\text{OH})_3$ $K_{sp} = 3.0 \times 10^{-33}$ $\text{Al}(\text{OH})_3 \rightleftharpoons \text{Al}^{3+} + 3\text{OH}^-$ $K_{sp} = [\text{Al}^{3+}] [\text{OH}^-]^3$

Let $x = \text{solubility}$ $3.0 \times 10^{-33} = x(3x)^3$ $[\text{Al}^{3+}] = 3.3 \times 10^{-9} \text{ M}$

then $[\text{Al}^{3+}] = x$ $x = 3.25 \times 10^{-9} \text{ M}$ $[\text{OH}^-] = 9.7 \times 10^{-9} \text{ M}$

$[\text{OH}^-] = 3x$

- B. $\text{Pb}(\text{IO}_3)_2$



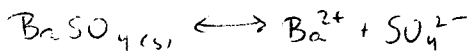
Let $x = \text{solubility}$ $3.7 \times 10^{-13} = x(2x)^2$ $[\text{Pb}^{2+}] = 4.5 \times 10^{-5} \text{ M}$

then $[\text{Pb}^{2+}] = x$ $x = 4.52 \times 10^{-5}$ $[\text{IO}_3^-] = 9.0 \times 10^{-5} \text{ M}$

$[\text{IO}_3^-] = 2x$

key.

7. A suspension of barium sulphate is used to improve the quality of X-rays in the digestive system. If the patient is required to drink 0.400 L of this suspension, calculate the actual mass in grams of the dissolved BaSO_4 .



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

Let x = solubility of BaSO_4

$$\text{then } [\text{Ba}^{2+}] = [\text{SO}_4^{2-}] = x$$

$$1.1 \times 10^{-10} = x^2$$

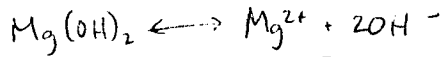
$$x = 1.05 \times 10^{-5} \text{ M}$$

$$\text{mols BaSO}_4 = 1.05 \times 10^{-5} \text{ M} \times 0.400 \text{ L} = 4.20 \times 10^{-6} \text{ mols}$$

$$\text{Fw BaSO}_4 = 137.3 + 32.1 + 4(16.0) = 233.4 \text{ g/mol}$$

$$\text{mass} = 4.20 \times 10^{-6} \text{ mol} \times 233.4 \text{ g/mol} = 9.8 \times 10^{-4} \text{ g}$$

8. Calculate the K_{sp} for Mg(OH)_2 if the solubility of magnesium hydroxide is 7.6 mg/L.



$$\text{Fw Mg(OH)}_2 = 24.3 + 2(16.0) + 2(1.0) = 58.3 \text{ g/mol}$$

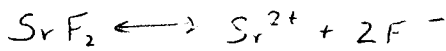
$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2$$

$$[\text{Mg}^{2+}] = 7.6 \text{ mg/L} \div 58.3 \text{ g/mol} \times 1000 \text{ mg/g} = 1.30 \times 10^{-4} \text{ M}$$

$$[\text{OH}^-] = 2(1.30 \times 10^{-4} \text{ M}) = 2.61 \times 10^{-4} \text{ M}$$

$$K_{sp} = (1.30 \times 10^{-4})(2.61 \times 10^{-4})^2 = 8.8 \times 10^{-12}$$

9. What maximum $[\text{F}^-]$ exists in a solution in which the $[\text{Sr}^{2+}] = 4.4 \times 10^{-3} \text{ M}$?

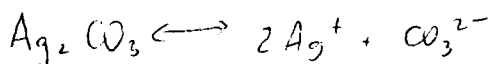


$$K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 = 4.3 \times 10^{-9}$$

$$[\text{F}^-] = 9.9 \times 10^{-4} \text{ M}$$

$$[\text{F}^-]^2 = \frac{4.3 \times 10^{-9}}{4.4 \times 10^{-3}} = 9.77 \times 10^{-7}$$

10. Show by calculation if a precipitate forms when 10.0 mL of 0.010 M AgNO_3 are mixed with an equal volume of 0.10 M Na_2CO_3 .



$$K_{sp} = [\text{Ag}^+]^2[\text{CO}_3^{2-}] = 8.5 \times 10^{-12}$$

$$[\text{Ag}^+] = \frac{10.0 \text{ mL} \times 0.010 \text{ M}}{20.0 \text{ mL}} = 0.0050 \text{ M}$$

$$[\text{CO}_3^{2-}] = \frac{10.0 \text{ mL} \times 0.10 \text{ M}}{20.0 \text{ mL}} = 0.050 \text{ M}$$

$$\text{TIP} = (0.0050)^2(0.050) = 1.25 \times 10^{-6}$$

TIP > K_{sp} \therefore a ppt forms.

11. A solution may contain Ba^{2+} and/or Al^{3+} . Describe a procedure to confirm the presence or absence of these ions.

① Add K_2SO_4 to ppt Ba^{2+}

② Add K_2SO_3 to ppt Al^{3+}

12.

- How many moles of PbI_2 would dissolve in water in which the $[\text{Pb}^{2+}] = 5.0 \times 10^{-6} \text{ M}$?

- no volume given omit question.

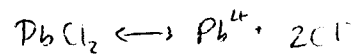
13. A 25.0 mL sample of saturated PbCl_2 solution is titrated to the endpoint with 48.1 mL of 0.015 M AgNO_3 solution. Calculate the K_{sp} of PbCl_2 .

$$\text{mols Ag}^+ = 48.1 \text{ mL} \times 0.015 \text{ M} = 0.722 \text{ mmol}$$

$$\text{mols Ag}^+ = \text{mols Cl}^-$$

$$[\text{Cl}^-] = \frac{0.722 \text{ mmol}}{25.0 \text{ mL}} = 2.89 \times 10^{-2} \text{ M}$$

$$[\text{Pb}^{2+}] = \frac{1}{2} (2.89 \times 10^{-2} \text{ M}) = 1.44 \times 10^{-2} \text{ M}$$



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

$$K_{sp} = (1.44 \times 10^{-2})(2.89 \times 10^{-2})^2$$

$$= 1.2 \times 10^{-5}$$

Chemistry 12
Solubility Review

Key.

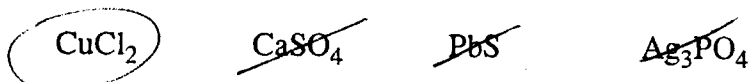
1. Molecular solutions do not conduct electricity because they contain

- A. molecules only
- B. cations and anions
- C. molecules and anions
- D. molecules and cations

2. To determine the solubility of a solute in water, a solution must be prepared that is

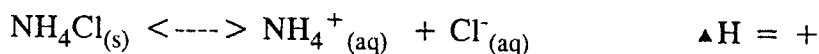
- A. saturated
- B. unsaturated
- C. concentrated
- D. supersaturated

3. Froms the list of salts below, how many are considered soluble at 25°C?



- A. none
- B. one
- C. two
- D. three

4. Consider the following equilibrium:



Which of the following will increase the solubility of ammonium chloride?

- A. stirring the solution
- B. adding water
- C. adding NH₄Cl
- D. heating

5. Na₂SO₄ solution is slowly added to a solution which contains 0.10 M Ba²⁺ and 0.10 M Pb²⁺. Which of the following statements describes the result of the addition of Na₂SO₄?

- A. BaSO₄ precipitates first because it is more soluble. 1.1×10^{-10}
- B. PbSO₄ precipitates first because it is more soluble. 1.8×10^{-8}
- C. BaSO₄ precipitates first because it is less soluble.
- D. PbSO₄ precipitates first because it is less soluble.

6. Identify the most soluble sulphide.

- A. HgS $K_{sp} = 1.6 \times 10^{-54}$
- B. PbS $K_{sp} = 7.0 \times 10^{-29}$
- C. FeS $K_{sp} = 3.7 \times 10^{-19}$
- D. MnS $K_{sp} = 2.3 \times 10^{-13}$

7. Four samples of a solution were analyzed and the following data were collected:

Anion added	observation
S ²⁻	nothing
SO ₄ ²⁻	precipitate
OH ⁻	nothing
CO ₃ ²⁻	precipitate

Which one of the following group II cations is found in the unknown solution?

- A. Mg²⁺
- B. Ca²⁺
- C. Sr²⁺
- D. Ba²⁺

8. The $[\text{OH}^-]$ is measured to be 3.3×10^{-3} mol/L in a 100 mL sample of a saturated solution of $\text{Al}(\text{OH})_3$. What is the solubility of $\text{Al}(\text{OH})_3$?
- A. 1.1×10^{-4} mol/L
 B. 3.3×10^{-4} mol/L
 C. 1.1×10^{-3} mol/L
 D. 3.3×10^{-3} mol/L
9. Which of the following salts has the lowest solubility?
- A. copper I chloride
 B. ammonium sulphide
 C. potassium hydroxide
 D. mercury II sulphate
10. The mixture that could produce a precipitate of two compounds is
- A. HgSO_4 and FeCl_2
 B. AgNO_3 and MgCl_2
 C. K_2CO_3 and CuSO_4
 D. ZnSO_4 and $\text{Ba}(\text{OH})_2$
11. In a saturated solution of zinc hydroxide, at 40°C , the $[\text{Zn}^{2+}] = 1.8 \times 10^{-5}$ M. What is the K_{sp} of the compound?
- A. 5.8×10^{-15}
 B. 2.3×10^{-14}
 C. 1.8×10^{-14}
 D. 6.5×10^{-10}
12. When equal volumes of 0.060 M AgNO_3 and 0.00090 M $\text{Ba}(\text{BrO}_3)_2$ are mixed, the trial ion product is
- A. less than K_{sp} and a precipitate forms.
 B. greater than K_{sp} and a precipitate forms.
 C. less than K_{sp} and no precipitate forms.
 D. greater than K_{sp} and no precipitate forms.
- TIP 27×10^{-5}
 $K_{\text{sp}} = 5.3 \times 10^{-5}$
13. What is the maximum amount of sodium sulphate that will dissolve in 1.0 L of 0.10 M $\text{Pb}(\text{NO}_3)_2$ without forming a precipitate?
- A. 1.8×10^{-8} mol
 B. 1.8×10^{-7} mol
 C. 1.3×10^{-4} mol
 D. 1.0×10^{-1} mol
14. Which one of the following equilibrium systems is described by a K_{sp} ?
- A. $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{CaO}(\text{s}) + \text{CO}_2(\text{g})$
 B. $\text{CaCO}_3(\text{s}) \rightleftharpoons \text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq})$
 C. $\text{Ca}^{2+}(\text{aq}) + \text{CO}_3^{2-}(\text{aq}) \rightleftharpoons \text{CaCO}_3(\text{s})$
 D. $\text{Ca}(\text{OH})_2(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq}) \rightleftharpoons \text{CaCO}_3(\text{s}) + 2\text{H}_2\text{O}(\text{l})$
15. In an experiment, a student mixes equal volumes of 0.0020 M Pb^{2+} ions with 0.0040 M I^- ions. What is TIP?
- A. 4.0×10^{-9}
 B. 3.2×10^{-8}
 C. 1.3×10^{-7}
 D. 8.0×10^{-6}
16. A 0.50 L solution of CuBr_2 contains 0.30 mol Br^- ions. What are the ionic concentrations in the solution?
- A. $[\text{Cu}^{2+}] = 0.15$ M $[\text{Br}^-] = 0.30$ M
 B. $[\text{Cu}^{2+}] = 0.30$ M $[\text{Br}^-] = 0.60$ M
 C. $[\text{Cu}^{2+}] = 0.60$ M $[\text{Br}^-] = 0.60$ M
 D. $[\text{Cu}^{2+}] = 0.60$ M $[\text{Br}^-] = 1.20$ M