Chemistry 12 Solubility of Compounds review Name Key.

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

a.	$CuBr$	low
b.	$CuCl_2$	
c.	$NaOH$	
d.	Na_3PO_4	
e.	$AgOH$	low
f.	MgS	
g.	$Sr(OH)_2$	
h.	$PbSO_4$	low
i.	$CdSO_4$	
j.	Rb_2SO_4	
k.	$CuOH$	low
l.	$BaSO_4$	low
m.	NH_4OH	
n.	Fr_2S	
o.	$CuSO_4$	
p.	CrS	low
q.	$(NH_4)_2Cr_2O_7$	
r.	K_2CrO_4	
s.	$NaCl$	
t.	KNO_3	
u.	NH_4NO_3	
v.	$HCOOH$	
w.	$PbCO_3$	low
x.	$PbNO_3$	
y.	Na_2SO_3	

Key

Solubility #1

1. Write the net ionic equation(s) for the reaction(s) when equal volumes of 0.2 M $Sr(OH)_2$ and 0.2 M $Fe_2(SO_4)_3$ are mixed.

$Sr^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \longleftrightarrow SrSO_{4(s)}$ $\quad Fe^{3+}_{(aq)} + 3OH^-_{(aq)} \longleftrightarrow Fe(OH)_{3(s)}$

2. Describe the equilibrium that exists in a saturated solution of $BaSO_4$ in contact with some solid residue of $BaSO_4$. $BaSO_{4(s)} \longleftrightarrow Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$.

at equilibrium the $[Ba^{2+}]$ and $[SO_4^{2-}]$ 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

      AgCl silver chloride

      (ii) the yellow precipitate

      AgI silver iodide

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

   $Ag^+_{(aq)} + Cl^-_{(aq)} \longleftrightarrow AgCl_{(s)}$

4. A solution of $Ba(NO_3)_2$ 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.

   $Ba^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \longleftrightarrow BaCO_{3(s)}$

   (b) Explain why the precipitate reacts in HNO_3.

   $BaCO_{3(s)} + 2H^+_{(aq)} \longleftrightarrow Ba^{2+}_{(aq)} + CO_{2(g)} + H_2O_{(l)}$

   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 $Sr(OH)_2$. 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?

   CaS and $Sr(OH)_2$

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

   $Ca^{2+}(aq) + 2OH^-(aq) \rightleftarrows Ca(OH)_{2(s)}$

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.

$Ca^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)} \longleftrightarrow CaSO_{4(s)}$

Adding $Ca(NO_3)_2$ increases the $[Ca^{2+}]$ 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.

$$NaCl_{(s)} \longleftrightarrow Na^+_{(aq)} + Cl^-_{(aq)}$$

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.

NH_4^+ Cl^-

9. Write an equation that describes the equilibrium present in a saturated solution of $Cu_3(PO_4)_2$.

$$Cu_3(PO_4)_{2(s)} \longleftrightarrow 3\,Cu^{2+}_{(aq)} + 2\,PO_4^{3-}{}_{(aq)}$$

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.

CuS

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

$$2\,Na^+_{(aq)} + S^{2-}_{(aq)} + Cu^{2+}_{(aq)} + 2\,Cl^-_{(aq)} \longleftrightarrow CuS_{(s)} + 2\,Na^+_{(aq)} + 2\,Cl^-_{(aq)}$$

(c) Identify **all** spectator ions.

Na^+, Cl^-

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

$Ba(NO_3)_2$ or Na_2SO_4

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

$$Mg(OH)_{2(s)} \longleftrightarrow Mg^{2+}_{(aq)} + 2\,OH^-_{(aq)}$$

13. When solid $Ca(NO_3)_2$ 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 $[CO_3^{2-}]$ and more $MgCO_3$ dissolves.

14. A student mixes equal volumes of 0.20 M Na_2CO_3 and 0.20 M $Ba(NO_3)_2$, forming a white precipitate.

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

$$Ba^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)} \longleftrightarrow BaCO_{3(s)}$$

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

$$BaCO_3 + 2H^+ \longrightarrow Ba^{2+} + CO_{2(g)} + H_2O_{(l)}$$

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.

$$BaSO_{4(s)} \longleftrightarrow Ba^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$$

Solubility #2

1. 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 $Ca(NO_3)_2$ or $Sr(NO_3)_2$
   or $Ba(NO_3)_2$ allow it to settle $\longrightarrow$ ppts SO_4^{2-}

   ② Add $CuNO_3$ to ppt Cl^-. Allow it to settle out

2. A solution is known to contain Cu^+, Be^{2+} and Sr^{2+} ions, each at a concentration of 0.20 M.

   (a) What compound could be used to precipitate the Sr^{2+} while leaving the other two cations in solution?

   Na_2SO_4

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

   $Sr^{2+}(aq) + SO_4^{2-}(aq) \longleftrightarrow SrSO_{4(s)}$

3. 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 $Pb(OH)_2$ and $Mg(OH$
   2) Use Na_2SO_4 to ppt $SrSO_4$

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.

   $Ag_2CO_3 \longleftrightarrow 2Ag^+ + CO_3^{2-}$ or halide ion $Cl^- + Ag^+ \longleftrightarrow AgCl_{(}$

   add something to remove Ag^+ or CO_3^{2-} eg Acid $\rightarrow CO_3^{2-} + 2H^+ \longleftrightarrow$ $CO_2 + H_2$

5. A solution contains Ca^{2+}, Sr^{2+} and Pb^{2+} ions that must be separated.

   (a) Identify an anion that could be used to remove **only** the lead ion by precipitation.

   Cl^-, I^-, Br^-, S^{2-}

   (b) 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 NaCl $\longrightarrow$ $Ag^+ + Cl^- \longleftrightarrow AgCl$

② Add Na_2SO_4 $\longrightarrow$ $Ba^{2+} + SO_4^{2-} \longleftrightarrow BaSO_4$

③ Add NaOH $\longrightarrow$ $Mg^{2+} + 2OH^- \longleftrightarrow Mg(OH)_2$

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 NaCl $\longrightarrow$ $Ag^+ + Cl^- \longleftrightarrow AgCl$

② Add Na_2S $\longrightarrow$ $2Fe^{3+} + 3S^{2-} \longleftrightarrow Fe_2S_3$

③ Add Na_2SO_4 $\longrightarrow$ $Sr^{2+} + SO_4^{2-} \longleftrightarrow SrSO_4$

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?

$Be(NO_3)_2$ ~ $Mg(NO_3)_2$ ~ $Ca(NO_3)_2$ ~ $Ba(NO_3)_2$

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

$Be^{2+}_{(aq)} + 2OH^-_{(aq)} \longleftrightarrow Be(OH)_{2(s)}$

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?

   $2x^2(x)$ Ag_2CO_3

   *saturated Ag_2CO_3 saturated AgCl

   $[Ag^+] = 2.57\times10^{-4}$ M $[Ag^+] = 1.34\times10^{-5}$ M

2. In which solution would you find the highest magnesium ion concentration: saturated $MgCO_3$ or ~~sturated MgF_2?~~ saturated $Mg(OH)_2$?

   *$MgCO_3$ $[Mg^{2+}] = 2.61\times10^{-3}$ M

   $Mg(OH)_2$ $[Mg^{2+}] = 1.11\times10^{-4}$ M

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

$[Ag^+] = 7.3\times10^{-3}$ M ← x^2	$AgBrO_3$	5.3×10^{-5}	$4x^3$	Ag_2CO_3	8.5×10^{-12}
$[Ag^+] = 7.4\times10^{-3}$ M ← $4x^3$	$Ag_2Cr_2O_7$ *	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.

4. Calculate the solubility of $BaSO_4$. 1.0×10^{-5} M

5. Calculate the mass of solid $PbSO_4$ that can be dissolved in 5.0 L of solution at 25°C.

   0.20 g.

6. Calculate the solubility of PbI_2 at 25°C.

   1.3×10^{-3} M

7. What is the solubility of calcium oxalate at 25°C in g/L?

   6.1×10^{-3} g/L

8. How many milligrams of $CaCO_3$ would be dissolved in 1.0 L of saturated solution at 25°C.

   7.1 mg/L

9. What mass of PbI_2 will dissolve in 250 mL of water?

   0.15 g

10. a. What is the solubility of magnesium hydroxide? 1.11×10^{-4} M

    b. What is the $[OH^-]$ in a saturated solution of magnesium hydroxide?

    2.22×10^{-4} M

Solubility #4

1. 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.

2. Calculate the K_{sp} for SrF_2 if the solubility is 0.122 g/L.

3. A saturated solution of calcium hydroxide is found to have $[OH^-]$ of 2.09×10^{-2} M. Calculate the K_{sp} for $Ca(OH)_2$.

4. The solubility of Ag_2SO_4 is 0.62 g/L at 6° C. What is the K_{sp} at this temperature?

5. 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.

6. At 20° C the solubility of PbF_2 is 64 mg per 100 mL of solution. Determine the K_{sp} for lead II fluoride. $0.64\ g/L \times \frac{1\ mol}{245.2\ g} = 2.61 \times 10^{-3}\ mol/L$ $[Pb^{2+}] = 2.61 \times 10^{-3}$ M
$K_{sp} = [Pb^{2+}][F^-]^2 = 7.1 \times 10^{-8}$ $[F^-] = 5.22 \times 10^{-3}$ M

7. The equilibrium in a saturated ZnI_2 solution is given by:

$$ZnI_{2(s)} \text{ <----> } Zn^{2+}_{(aq)} + 2I^-_{(aq)}$$

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

(a) solid NaI. Solubility of ZnI_2 will decrease
(b) solid $Zn(NO_3)_2$ decrease
(c) solid NaOH $Zn(OH)_2$ (s) produced; solubility will increase
(d) concentrated HCl no effect
(e) solid NH_4NO_3 no effect

#1 $[Ag^+] = \frac{2(6.00 \times 10^{-6}\ mol)}{1.00\ L} = 1.20 \times 10^{-5}$ M $[CO_3^{2-}] = 2.00M + 6.00 \times 10^{-6}\ mol/L = 2.00M$

$K_{sp} = [Ag^+]^2[CO_3^{2-}] = \underline{\underline{2.88 \times 10^{-10}}}$

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

$[F^-] = 2(9.71 \times 10^{-4} M) = 1.94 \times 10^{-3}$ M

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

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

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

#4 $0.62\ g/L \times \frac{1\ mol}{311.9\ g} = 1.99 \times 10^{-3}\ mol/L$ $[Ag^+] = 2(1.99 \times 10^{-3} M) = 3.98 \times 10^{-3}$ M
$[SO_4^{2-}] = 1.99 \times 10^{-3}$ M

$K_{sp} = [Ag^+]^2[SO_4^{2-}] = \underline{3.1 \times 10^{-8}}$

#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, $K_{sp} = [Ag^+]^2[SO_4^{2-}]$

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 $Ca(NO_3)_2$ 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 $[Ca^{2+}]$ of 1.0×10^{-3} M. (a) If the K_{sp} of CaF_2 is 1.7×10^{-10}, what is the maximum $[F^-]$ 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 $[Ba^{2+}]$ and the $[CrO_4^{2-}]$ at equilibrium.

6. The $[Ag^+]$ of a solution is 4.0×10^{-3} M. Calculate the $[Cl^-]$ 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?

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1. $CaSO_{4(s)} \leftrightarrow Ca^{2+}_{(aq)} + SO_{4\,(aq)}^{2-}$

$K_{sp} = [Ca^{2+}][SO_4^{2-}] = 7.1 \times 10^{-5}$

$7.1 \times 10^{-5} = (0.0010M)[SO_4^{2-}]$

$[SO_4^{2-}] = 7.1 \times 10^{-2}$ M

mols SO_4^{2-} = 7.1×10^{-2} mol/L × 5.0L = 0.355 mol

FW Na_2SO_4 = 2(23.0u) + 32.1u + 4(16.0u) = 142.1 g/mol

mass Na_2SO_4 = 0.355 mol × 142.1 g/mol = 50.g

2. $PbCl_{2(s)} \leftrightarrow Pb^{2+}_{(aq)} + 2Cl^-_{(aq)}$

mols $PbCl_2$ in solution = .750L × 1.0×10^{-2} M = 7.5×10^{-3} mol

FW $PbCl_2$ = 207.2u + 2(35.5u) = 278.2 g/mol

mols $PbCl_2$ ppt = 1.80g ÷ 278.2 g/mol = 6.47×10^{-3} mol

mols left in sol'n = $7.5 \times 10^{-3} - 6.47 \times 10^{-3} = 1.03 \times 10^{-3}$ mol

$[PbCl_2] = \frac{1.03 \times 10^{-3}\text{ mol}}{.750\text{ L}} = 1.4 \times 10^{-3}$ M

3. $AgCl_{(s)} \leftrightarrow Ag^+_{(aq)} + Cl^-_{(aq)}$

mols Ag^+ = 0.050L × 0.60M = 0.030 mol

mols Cl^- = 0.100L × 0.10M × 2 = 0.020 mol

1:1 ratio ∴ Cl^- is the limiting factor

0.020 mol of precipitate will form

FW AgCl = 107.9u + 35.5u = 143.4 g/mol

mass of AgCl = 0.020 mol × 143.4 g/mol = 2.9 g

Solubility #5 - Answer key

4(a) $CaF_{2(s)} \longleftrightarrow Ca^{2+}_{(aq)} + 2F^-_{(aq)}$

$K_{sp} = [Ca^{2+}][F^-]^2 = 1.7 \times 10^{-10}$

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

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

5.

R	$BaCrO_{4(s)} \longleftrightarrow$	$Ba^{2+}_{(aq)}$	$CrO_{4\,(aq)}^{2-}$
i		0.00	0.00
c		+ x	+ x
E		x	x

$K_{sp} = [Ba^{2+}][CrO_4^{2-}] = 1.2 \times 10^{-10}$

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

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

The $[Ba^{2+}]$ and $[CrO_4^{2-}]$ are $1.1 \times 10^{-5}\,M$ at equilibrium

6. $AgCl_{(s)} \longleftrightarrow Ag^+_{(aq)} + Cl^-_{(aq)}$

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

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

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

7.

R	$PbI_{2(s)} \longleftrightarrow$	$Pb^{2+}_{(aq)}$	$2I^-_{(aq)}$
i		0.00	0.00
c		+ x	+ 2x
E		x	2x

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

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

$= 4x^3$

$x = 1.3 \times 10^{-3}\,M$ = solubility of PbI_2

8. $AgCl_{(s)} \longleftrightarrow Ag^+ + Cl^-$

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

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

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

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

then at eqm:

$[Ag^+] = x$

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

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

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

at eqm:

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

$[Cl^-] = 0.11\,M$

Answer Key

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 $Ca(NO_3)_2$ 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 $[CH_3COO^-]$ is 5.2×10^{-3} M. K_{sp} CH_3COOAg = 3.7×10^{-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 $[CO_3^{2-}]$ = 3.0×10^{-3} M and $[IO_3^-]$ = 5.0×10^{-3} M. Which substance precipitates first?

9. Will a precipitate of $Al(OH)_3$ 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} $Al(OH)_3$ = 3.7×10^{-15}

10. Will a precipitate form when 400 mL of 0.0020 M $Ba(OH)_2$ are mixed with 200 mL of 0.0020 M H_2SO_4?

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1. $AgCl \leftrightarrow Ag^+ + Cl^-$

$K_{sp} = [Ag^+][Cl^-] = 1.8\times10^{-10}$

$[Ag^+] = \dfrac{0.0050\,g \div 169.9\ g/mol}{2.00\,L} = 1.47\times10^{-5}\,M$

$[Cl^-] = 0.0010\,M$

$TIP = (1.47\times10^{-5})(0.0010)$

$= 1.5\times10^{-8}$

Since TIP > ksp a ppt will form.

2. $CaF_2 \leftrightarrow Ca^{2+} + 2F^-$

$K_{sp} = [Ca^{2+}][F^-]^2 = 4.0\times10^{-11}$

$[Ca^{2+}] = \dfrac{2.0\times10^{-3}\,g \div 164.0\ g/mol}{0.500\,L} = 2.44\times10^{-5}\,M$

$[F^-] = \dfrac{2.0\times10^{-3}\,g \div 42.0\,g/mol}{0.500\,L} = 9.52\times10^{-5}\,M$

$TIP = (2.44\times10^{-5})(9.52\times10^{-5})^2$

$= 2.2\times10^{-13}$

Since TIP < Ksp no ppt will form.

Solubility #6 — Answer key

3. $AgCl \longleftrightarrow Ag^+ + Cl^-$

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

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

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

$TIP = (1.00 \times 10^{-5})(2.0 \times 10^{-3})$
$= 2.0 \times 10^{-8}$

TIP > ksp a ppt will form.

4. $AgCH_3COO \longleftrightarrow Ag^+ + CH_3COO^-$

$K_{sp} = [Ag^+][CH_3COO^-] = 3.7 \times 10^{-3}$

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

$[CH_3COO^-] = \dfrac{45\,mL \times 5.2 \times 10^{-3}\,M}{45\,mL + 15\,mL} = 3.9 \times 10^{-3}\,M$

$TIP = (0.25)(3.9 \times 10^{-3})$
$= 9.8 \times 10^{-4}$

TIP < ksp no ppt will form.

5. $BaSO_4 \longleftrightarrow Ba^{2+} + SO_4^{2-}$

$K_{sp} = [Ba^{2+}][SO_4^{2-}] = 1.1 \times 10^{-10}$

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

$[SO_4^{2-}] = \dfrac{0.15\,g \div 174.3\,g/mol}{2.0\,L} = 4.3 \times 10^{-4}\,M$

$TIP = (1.7 \times 10^{-5})(4.3 \times 10^{-4})$
$= 7.3 \times 10^{-9}$

TIP > ksp a ppt will form

6. $AgCl \longleftrightarrow Ag^+ + Cl^-$

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

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

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

$TIP = (1.2 \times 10^{-6})(6.0 \times 10^{-5})$
$= 7.2 \times 10^{-11}$

TIP < ksp no ppt will form

7. $AgCl \longleftrightarrow Ag^+ + Cl^-$

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

$[Cl^-] = 0.010\,M$

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

$Ag_2CrO_4 \longleftrightarrow 2Ag^+ + CrO_4^{2-}$

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

$[CrO_4^{2-}] = 0.010\,M$

$[Ag^+] = \sqrt{\dfrac{K_{sp}}{[CrO_4^{2-}]}} = 1.05 \times 10^{-5}\,M$ to ppt Ag_2CrO_4

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

Solubility #6 Answer key

8. For a precipitate of Ag_2CO_3

$Ag_2CO_3 \longleftrightarrow 2Ag^+ + CO_3^{2-}$

$Ksp = [Ag^+]^2[CO_3^{2-}] = 8.5\times10^{-12}$

$[CO_3^{2-}] = 3.0\times10^{-3}$

$[Ag^+] = \sqrt{\dfrac{ksp}{[CO_3^{2-}]}} = 5.3\times10^{-5}\,M$

For a precipitate of $AgIO_3$

$AgIO_3 \longleftrightarrow Ag^+ + IO_3^-$

$ksp = [Ag^+][IO_3^-] = 3.2\times10^{-8}$

$[IO_3^-] = 5.0\times10^{-3}\,M$

$[Ag^+] = \dfrac{ksp}{[IO_3^-]} = 6.4\times10^{-6}\,M$

The $AgIO_3$ ppts first

9. $Al(OH)_3 \longleftrightarrow Al^{3+} + 3OH^-$

$Ksp = [Al^{3+}][OH^-]^3 = 3.7\times10^{-15}$

$[Al^{3+}] = \dfrac{0.50\,L \times 2.0\times10^{-3}\,M}{1000\,L} = 1.0\times10^{-6}\,M$

$[OH^-] = \dfrac{0.50\,L \times 4.0\times10^{-2}\,M}{1000\,L} = 2.0\times10^{-5}\,M$

$TIP = (1.0\times10^{-6})(2.0\times10^{-5})^3$

$= 8.0\times10^{-21}$

TIP < ksp so no ppt will form

10. $BaSO_4 \longleftrightarrow Ba^{2+} + SO_4^{2-}$

$Ksp = [Ba^{2+}][SO_4^{2-}] = 1.1\times10^{-10}$

$[Ba^{2+}] = \dfrac{0.0020\,M \times 400\,mL}{400\,mL + 200\,mL} = 1.33\times10^{-3}\,M$

$[SO_4^{2-}] = \dfrac{0.0020\,M \times 200\,mL}{400\,mL + 200\,mL} = 6.67\times10^{-4}\,M$

$TIP = (1.33\times10^{-3})(6.67\times10^{-4})$

$= 8.9\times10^{-7}$

TIP > ksp so a ppt will form

Key.

Solubility Open-ended Review

1. Write an equation that describes the equilibrium present in a saturated solution of Cu_3PO_4.

$Cu_3PO_{4(s)} \longleftrightarrow 3Cu^+_{(aq)} + PO_{4(aq)}^{3-}$

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.

$Na_2CO_{3(s)} \longrightarrow 2Na^+_{(aq)} + CO_{3(aq)}^{2-}$

B. Calculate the concentration of each ion.

FW Na_2CO_3 = 2(23.0) + 12.0 + 3(16.0) = 106 g/mol

53 g ÷ 106 g/mol = 0.50 mol

$[Na_2CO_3]$ = 0.50 mol / 5.0 L = 0.10 M

$[Na^+]$ = 2(0.10 M) = 0.20 M

$[CO_3^{2-}]$ = 0.10 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:

A. Cu_2S $\quad Cu_2S \longleftrightarrow 2Cu^+ + S^{2-}$ $\quad k_{sp} = [Cu^+]^2[S^{2-}]$

B. PbI_2 $\quad PbI_2 \longleftrightarrow Pb^{2+} + 2I^-$ $\quad k_{sp} = [Pb^{2+}][I^-]^2$

C. $Al_2(SO_4)_3$ $\quad$ soluble

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 $\quad CuS$

B. Write the full equation for the reaction.

$Na_2S_{(aq)} + CuCl_{2(aq)} \longleftrightarrow CuS_{(s)} + 2NaCl_{(aq)}$

C. Write the net ionic equation for the reaction.

$Cu^{2+}_{(aq)} + S^{2-}_{(aq)} \longleftrightarrow CuS_{(s)}$

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 $Pb(NO_3)_2$

$Pb^{2+}_{(aq)} + 2Cl^-_{(aq)} \longleftrightarrow PbCl_{2(s)}$

B. SrS and $CuSO_4$

$Sr^{2+}_{(aq)} + SO_{4(aq)}^{2-} \longleftrightarrow SrSO_{4(s)}$

$Cu^{2+}_{(aq)} + S^{2-}_{(aq)} \longleftrightarrow CuS_{(s)}$

C. $(NH_4)_3PO_4$ and $Mg(OH)_2$

$3Mg^{2+}_{(aq)} + 2PO_{4(aq)}^{3-} \longleftrightarrow Mg_3(PO_4)_2$

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

A. $Al(OH)_3$ $\quad K_{sp} = 3.0 \times 10^{-33}$

$Al(OH)_3 \longleftrightarrow Al^{3+} + 3OH^-$ $\quad k_{sp} = [Al^{3+}][OH^-]^3$

Let x = solubility

then $[Al^{3+}] = x$

$[OH^-] = 3x$

$3.0 \times 10^{-33} = x(3x)^3$

$x = 3.25 \times 10^{-9}$ M

$[Al^{3+}] = 3.3 \times 10^{-9}$ M

$[OH^-] = 9.7 \times 10^{-9}$ M

B. $Pb(IO_3)_2$

$Pb(IO_3)_2 \longleftrightarrow Pb^{2+} + 2IO_3^-$ $\quad k_{sp} = [Pb^{2+}][IO_3^-]^2 = 3.7 \times 10^{-13}$

Let x = solubility

then $[Pb^{2+}] = x$

$[IO_3^-] = 2x$

$3.7 \times 10^{-13} = x(2x)^2$

$x = 4.52 \times 10^{-5}$

$[Pb^{2+}] = 4.5 \times 10^{-5}$ M

$[IO_3^-] = 9.0 \times 10^{-5}$ M

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 $BaSO_4$.

$BaSO_{4(s)} \longleftrightarrow Ba^{2+} + SO_4^{2-}$

$k_{sp} = [Ba^{2+}][SO_4^{2-}] = 1.1 \times 10^{-10}$

Let x = solubility of $BaSO_4$

then $[Ba^{2+}] = [SO_4^{2-}] = x$

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

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

mols $BaSO_4 = 1.05 \times 10^{-5}$ M $\times$ 0.400 L $= 4.20 \times 10^{-6}$ mols

FW $BaSO_4 = 137.3 + 32.1 + 4(16.0) = 233.4$ g/mol

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

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

$Mg(OH)_2 \longleftrightarrow Mg^{2+} + 2OH^-$

$k_{sp} = [Mg^{2+}][OH^-]^2$

FW $Mg(OH)_2 = 24.3 + 2(16.0) + 2(1.0) = 58.3$ g/mol

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

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

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

$= 8.8 \times 10^{-12}$

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

$SrF_2 \longleftrightarrow Sr^{2+} + 2F^-$

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

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

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

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.

$Ag_2CO_3 \longleftrightarrow 2Ag^+ + CO_3^{2-}$

$k_{sp} = [Ag^+]^2[CO_3^{2-}] = 8.5 \times 10^{-12}$

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

$[CO_3^{2-}] = \dfrac{10.0 \text{ mL} \times 0.10 \text{ M}}{20.0 \text{ mL}} = 0.050$ M

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

TIP > ksp ∴ 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 $[Pb^{2+}] = 5.0 \times 10^{-6}$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 $PbCl_2$.

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

mols Ag^+ = mols Cl^-

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

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

$PbCl_2 \longleftrightarrow Pb^{2+} + 2Cl^-$

$k_{sp} = [Pb^{2+}][Cl^-]^2$

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

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

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. molecues 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?

$CuCl_2$ ~~$CaSO_4$~~ ~~PbS~~ ~~Ag_3PO_4~~

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

4. Consider the following equilibrium:

$$NH_4Cl_{(s)} \text{ <----> } NH_4^+{}_{(aq)} + Cl^-{}_{(aq)} \qquad \Delta H = +$$

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

A. stirring the solution
B. adding water
C. adding NH_4Cl
D. heating

5. Na_2SO_4 solution is slowly added to a solution which contains 0.10 M Ba^{2+} and 0.10 M Pb^{2+}. Which of the following statements describes the result of the addition of Na_2SO_4?

A. $BaSO_4$ precipitates first because it is more soluble. 1.1×10^{-10}
B. $PbSO_4$ precipitates first because it is more soluble. 1.8×10^{-8}
C. $BaSO_4$ precipitates first because it is less soluble.
D. $PbSO_4$ precipitates first because it is less soluble.

6. Identify the most soluble sulphide.

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

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. Mg^{2+}
B. Ca^{2+}
C. Sr^{2+}
D. Ba^{2+}

8. The $[OH^-]$ is measured to be $3.3x10^{-3}$ mol/L in a 100 mL sample of a saturated solution of $Al(OH)_3$. What is the solubility of $Al(OH)_3$?

A. $1.1x10^{-4}$ mol/L
B. $3.3x10^{-4}$ mol/L
C. $1.1x10^{-3}$ mol/L
D. $3.3x10^{-3}$ 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 $Ba(OH)_2$

11. In a saturated solution of zinc hydroxide, at 40°C, the $[Zn^{2+}] = 1.8x10^{-5}$ M. What is the K_{sp} of the compound?

A. $5.8x10^{-15}$
B. $2.3x10^{-14}$
C. $1.8x10^{-14}$
D. $6.5x10^{-10}$

12. When equal volumes of 0.060 M $AgNO_3$ and 0.00090 M $Ba(BrO_3)_2$ 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 2.7×10^{-5}
ksp. 5.3×10^{-5}

13. What is the maximum amount of sodium sulphate that will dissolve in 1.0 L of 0.10 M $Pb(NO_3)_2$ without forming a precipitate?

A. $1.8x10^{-8}$ mol
B. $1.8x10^{-7}$ mol
C. $1.3x10^{-4}$ mol
D. $1.0x10^{-1}$ mol

14. Which one of the following equilibrium systems is described by a K_{sp} ?

A. $CaCO_{3(s)}$ <----> $CaO_{(s)} + CO_{2(g)}$
B. $CaCO_{3(s)}$ <----> $Ca^{2+}{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$
C. $Ca^{2+}{}_{(aq)} + CO_3{}^{2-}{}_{(aq)}$ <----> $CaCO_{3(s)}$
D. $Ca(OH)_{2(aq)} + H_2CO_{3(aq)}$ <----> $CaCO_{3(s)} + 2H_2O_{(l)}$

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. $4.0x10^{-9}$
B. $3.2x10^{-8}$
C. $1.3x10^{-7}$
D. $8.0x10^{-6}$

16. A 0.50 L solution of $CuBr_2$ contains 0.30 mol Br^- ions. What are the ionic concentrations in the solution?

A. $[Cu^{2+}] = 0.15$ M $[Br^-] = 0.30$ M
B. $[Cu^{2+}] = 0.30$ M $[Br^-] = 0.60$ M
C. $[Cu^{2+}] = 0.60$ M $[Br^-] = 0.60$ M
D. $[Cu^{2+}] = 0.60$ M $[Br^-] = 1.20$ M