Name: ________________________________

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

      $Na_2S_{(s)} \rightleftharpoons 2\ Na^+_{(aq)} + S^{2-}_{(aq)}$ $K_{sp} = [Na^+]^2[S^{2-}]$
   b) calcium iodide

      $CaI_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2\ I^-_{(aq)}$ $K_{sp} = [Ca^{2+}][I^-]^2$
   c) lithium carbonate

      $Li_2CO_{3(s)} \rightleftharpoons 2\ Li^+_{(aq)} + CO_3^{2-}{}_{(aq)}$ $K_{sp} = [Li^+]^2[CO_3^{2-}]$
   d) iron(II) sulfate

      $FeSO_{4(s)} \rightleftharpoons Fe^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$ $K_{sp} = [Fe^{2+}][SO_4^{2-}]$
   e) cobalt(II) nitrate

      $Co(NO_3)_{2(s)} \rightleftharpoons Co^{2+}_{(aq)} + 2\ NO_3^-{}_{(aq)}$ $K_{sp} = [Co^{2+}][NO_3^-]^2$
   f) barium phosphate

      $Ba_3(PO_4)_{2(s)} \rightleftharpoons 3\ Ba^{2+}_{(aq)} + 2\ PO_4^{3-}{}_{(aq)}$ $K_{sp} = [Ba^{2+}]^3[PO_4^{3-}]^2$

2. What are the concentrations of the resulting ions in saturated aqueous solutions of the following compounds?
   a) lead (II) sulfate [Pb^{2+}=1.3x10^{-4} M, SO_4^{2-}= 1.3x10^{-4} M]

      $PbSO_{4(s)} \rightleftharpoons Pb^{2+}_{(aq)} + SO_4^{2-}{}_{(aq)}$

      x x

      $K_{sp} = [Fe^{2+}][SO_4^{2-}]$

      $1.8x10^{-8} = x^2$

      x=1.3x10^{-4} M

      therefore [Pb^{2+}]=1.3x10^{-4} M and [SO_4^{2-}]= 1.3x10^{-4} M
   b) silver carbonate [Ag^+=2.6x10^{-4} M, CO_3^{2-}= 1.3x10^{-4} M]

      $Ag_2CO_{3(s)} \rightleftharpoons 2\ Ag^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

      2x x

      $K_{sp} = [Ag^+]^2[CO_3^{2-}]$

      $8.5x10^{-12} = 4x^3$

      x=1.3x10^{-4} M

      therefore [Ag^+]=2.6x10^{-4} M and [CO_3^{2-}]= 1.3x10^{-4} M
   c) magnesium hydroxide [Mg^{2+}=1.1x10^{-4} M, OH^-= 2.2x10^{-4} M]

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

      x 2x

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

      $5.6x10^{-12} = 4x^3$

      x=1.1x10^{-4} M

      therefore [Mg^{2+}]=1.1x10^{-4} M and [OH^-]= 2.2x10^{-4} M

3. Calculate the K_{sp} values of the following substances from their solubility in water.
   a) silver chloride with a solubility of 1.6x10^{-3}g/L [1.2x10^{-10}]

      MM=143.32g/mol

      $$x = \frac{1.6x10^{-3}\ g/L}{143.32\ g/mol}$$

      $$x = 1.1x10^{-5}\ mol/L$$

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

      x x

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

      $K_{sp} = x^2$

      $K_{sp} = (1.1x10^{-5})^2$

      $K_{sp} = 1.2x10^{-10}$

b) lithium carbonate with a solubility of 13.0 g/L [$2.3x10^{-2}$]

MM=73.89 g/mol

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

$$x = 0.18\ mol/L$$

$Li_2CO_{3(s)} \rightleftharpoons 2\ Li^+_{(aq)} + CO_3^{2-}{}_{(aq)}$

2x x

$K_{sp}= [Li^+]^2[CO_3^{2-}]$

$K_{sp}= 4x^3$

$K_{sp}= 4(0.18)^3$

$K_{sp} = 2.3x10^{-2}$

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

a) silver iodide [Ag^+=$9.2x10^{-9}$ M, I^-= $9.2x10^{-2}$ M]

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

x x

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

$8.5x10^{-17}= x^2$

x= $9.2x10^{-9}$

therefore [Ag^+]=$9.2x10^{-9}$ M and [Cl^-]= $9.2x10^{-9}$ M

b) strontium carbonate [Sr^{2+}=$2.4x10^{-5}$ M, CO_3^{2-}= $2.4x10^{-5}$ M]

$SrCO_{3(s)} \rightleftharpoons Sr^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$

x x

$K_{sp}= [Sr^{2+}][CO_3^{2-}]$

$5.6x10^{-10}= x^2$

x=$2.4x10^{-5}$ M

therefore [Sr^{2+}]=$2.4x10^{-5}$ M and [CO_3^{2-}]= $2.4x10^{-5}$ M

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

a) lead(II) iodide [Pb^{2+}=$1.3x10^{-3}$ M, I^-= $2.6x10^{-3}$ M]

$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2\ I^-_{(aq)}$

x 2x

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

$8.5x10^{-9}= 4x^3$

x=$1.3x10^{-3}$ M

therefore [Pb^{2+}]=$1.3x10^{-3}$ M and [I^-]= $2.6x10^{-6}$ M

b) silver sulfate [Ag^+=$2.8x10^{-2}$ M, SO_4^{2-}= $1.1x10^{-2}$ M]

$Ag^2SO_{4(s)} \rightleftharpoons 2\ Ag^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

2x x

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

$1.2x10^{-5}= 4x^3$

x=$5.7x10^{-3}$ M

therefore [Ag^+]=0.011 M and [SO_4^{2-}]= $1.1x10^{-2}$ M

c) iron(III) hydroxide [Fe^{3+}=$9.9x10^{-11}$ M, OH^-= $3.0x10^{-10}$ M]

$Fe(OH)_{3\ (s)} \rightleftharpoons Fe^{3+}_{(aq)} + 3\ OH^-_{(aq)}$

x 3x

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

$2.6x10^{-39}= 27x^4$

x=$9.9x10^{-11}$ M

therefore [Fe^{3+}]=$9.9x10^{-11}$ M and [OH^-]= $3.0x10^{-10}$ 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 [$2.0x10^{-4}$]

MM=239.85 g/mol

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

$$x = 0.014\ mol/L$$

$TlCl_{(s)} \rightleftharpoons Tl^+_{(aq)} + Cl^-_{(aq)}$

x x

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

$K_{sp}= x^2$

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

$K_{sp} = 2.0x10^{-4}$

b) silver bromide, $1.3x10^{-4}$ g/L at 20°C [$4.8x10^{-13}$]

MM=187.80 g/mol

$$x = \frac{1.3x10^{-4}\ g/L}{187.80\ g/mol}$$

$$x = 6.9x10^{-7}\ mol/L$$

$AgBr_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + Br^{-}_{(aq)}$

x x

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

$K_{sp}= x^2$

$K_{sp}= (6.9x10^{-7})^2$

$K_{sp} = 4.8x10^{-13}$

c) calcium fluoride, $1.6x10^{-2}$ g/L at 20°C [$3.4x10^{-11}$]

MM=78.80 g/mol

$$x = \frac{1.6x10^{-2}\ g/L}{78.80\ g/mol}$$

$$x = 2.0x10^{-4}\ mol/L$$

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

x 2x

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

$K_{sp}= 4x^3$

$K_{sp}= 4(2.0x10^{-4})^3$

$K_{sp} = 3.4x10^{-11}$

7. Calculate the maximum iodide ion concentration for lead(II) iodide dissolved in a $1.00x10^{-2}$ M solution of lead(II) nitrate. [$9.2x10^{-4}$ M]

$PbI_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2\ I^{-}_{(aq)}$

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

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

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

$8.5x10^{-9}= (1.00x10^{-2}+ \cancel{x})\ [2x]^2$

$x=4.6x10^{-4}$

$[I^{-}]=9.2x10^{-4}$ 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 $1.1x10^{-11}$. [$1.1x10^{-9}$]

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

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

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

$K_{sp}= [Ba^{2+}][SO_4^{2-}]$

$1.1x10^{-11}=[Ba^{2+}](0.010 + \cancel{x})$

$[Ba^{2+}]= 1.1x10^{-9}$

9. Calculate the maximum magnesium ion concentration in a 0.020 M aqueous solution of barium hydroxide. The K_{sp} of magnesium hydroxide is $1.2x10^{-11}$. [$7.5x10^{-9}$]

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

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

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

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

$1.2x10^{-11}= [Mg^{2+}](0.040+ \cancel{2x})^2$

$[Mg^{2+}] = 7.5x10^{-9}$ M

10. Upon addition of hydroxide ions to sea water, $Mg(OH)_2$ precipitates. If the magnesium ion concentration in sea water is $5.3x10^{-2}$ M, calculate the maximum hydroxide ion concentration in sea water. [$1.0x10^{-5}$]

$Mg(OH)_{2(s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2\ OH^{-}_{(aq)}$, $[Mg^{2+}] = 5.3x10^{-2}$ M

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

$5.6x10^{-12}= [5.3x10^{-2}]\ [OH^{-}]^2$

$[Mg^{2+}] = 1.0x10^{-5}$

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11. A sample of sea water contains 0.53 M of Cl^- ions and $8.4x10^{-4}$ 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 $1.6x10^{-3}$ and the K_{sp} for AgBr is $6.5x10^{-13}$. [Cl^-:$3.0x10^{-3}$ Br^-:$7.7x10^{-10}$, AgBr precipitates first]

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

For Cl^-

$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$
$K_{sp}= [Ag^+][Cl^-]$
$1.6x10^{-3}=[Ag^+][0.53]$
$[Ag^+]=3.0x10^{-3}$ M

For Br^-

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$
$K_{sp}= [Ag^+][Br^-]$
$6.5x10^{-13}=[Ag^+][8.4x10^{-4}]$
$[Ag^+]=7.7x10^{-10}$ 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]

$PbCl_{2(s)} \rightleftharpoons Pb^{2+}_{(aq)} + 2\ Cl^-_{(aq)}$
$K_{sp}= [Pb^{2+}][Cl^-]^2$
$1.2x10^{-5}= [Pb^{2+}][0.135]^2$
$[Pb^{2+}]=6.6x10^{-4}$ M

Moles of Pb^{2+} in 10.0 mL
n=cv
n=($6.6x10^{-4}$ M)(0.010 L)
n=$6.6x10^{-6}$ mol

mass of Pb^{2+}
MM of Pb =207.20 g/mol
m=nMM
m=($6.6x10^{-6}$ mol)(207.20 g/mol)
m=0.0014 g
m=1.4 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 $3.9x10^{-11}$. [Q=$4.0x10^{-8}$, yes a ppt will form]

MM of NaF =41.99 g/mol
n of NaF =m/MM
= (0.084 g)/(41.99 g/mol)
= $2.00x10^{-3}$ mol

$[F^-]=2.00x10^{-3}$ M
$[Ca^{2+}]$ from $CaCl_2$=0.010 M

$K_{sp}=[Ca^{2+}][F^-]^2$
$Q=[Ca^{2+}]_A[F^-]_A^2$

Q = (0.010 M)($2.00x10^{-3}$ M)^2
Q=$4.0x10^{-8}$
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 Ag_2SO_4=$1.2x10^{-5}$) [Q=$1.25x10^{-6}$]

$[Ag^+]$=(0.010 L)(0.010 M)/(0.020 L)
= 0.0050 M

$[SO_4^{2-}]$=(0.010 L)(0.010 M)/(0.020 L)
= 0.050 M

$Ag_2SO_{4(s)} \rightleftharpoons 2\ Ag^+_{(aq)} + SO_4^{2-}{}_{(aq)}$

$K_{sp}=[Ag^+][SO_4^{2-}]$
$Q= [Ag^+]_A^2[SO_4^{2-}]_A$
Q= (0.0050 M)^2(0.050 M)
Q = $1.25x10^{-6}$
Q is less than K_{sp} therefore no precipitate will form.

   b) 1.0 mL of 0.10 M $Ca(NO_3)_2$ and 1.0 L of 0.010 M NaF. (K_{sp} for CaF_2 3.45×10^{-11}) [Q=$1.0x10^{-8}$]

$[Ca^{2+}]$=(0.0010 L)(0.10 M)/(1.001 L)
= $1.0x10^{-5}$ M

$[F^-]$=(1.0 L)(0.010 M)/(1.001 L)
= 0.010 M

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

$K_{sp}=[Ca^{2+}][F^-]^2 =3.45\times10^{-11}$
$Q= [Ca^{2+}]_A[F^-]^2_A$
Q= ($1.0x10^{-5}$ M)(0.010 M)^2
Q = $1.0x10^{-8}$
Q is greater than K_{sp} therefore a precipitate will form.

c) 5.0 mL of 0.0040 M $AgNO_3$ and 15 mL of a solution containing 1.5 mg of Br^- ions. [Q=$9.5x10^{-7}$]

$[Ag^+]$=(0.0050 L)(0.0040 M)/(0.020 L)
= 0.0010 M

n of Br = (0.0015 g)/(79.90 g/mol)
= $1.9x10^{-5}$ mol
$[Br^-]$=($1.9x10^{-5}$ mol)/(0.020 L)
= $9.5x10^{-4}$ M

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$

K_{sp}= $[Ag^+][Br^-]$ = $5.4x10^{-13}$
Q=$[Ag^+]_A[Br^-]_A$
Q=(0.0010 M)($9.5x10^{-4}$ M)
Q=$9.5x10^{-7}$ M
Q is greater than K_{sp} therefore a precipitate will form.

15. Would you expect a precipitate of silver bromate (K_{sp}=$1.2x10^{-11}$) 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=$5.6x10^{-6}$]

$[Ag^+]$=(0.050 L)(0.0020 M)/(0.300 L)
= $3.3x10^{-4}$ M

$[BrO_3^-]$=(0.250 L)(0.020 M)/(0.300 L)
= 0.017 M

$AgBrO_{3(s)} \rightleftharpoons Ag^+_{(aq)} + BrO_3^-{}_{(aq)}$

K_{sp}= $[Ag^+][BrO_3^-]$ = $1.2x10^{-11}$
Q=$[Ag^+]_A[BrO_3^-]_A$
Q=($3.3x10^{-4}$ M)(0.0167 M)
Q=$5.6x10^{-6}$ M
Q is greater than K_{sp} therefore a precipitate will form.

16. Will a precipitate of $Mg(OH)_2$ form when 1.00 mL of 0.010 M $Ca(OH)_2$ is added to 1.0 L of 0.20 M $Mg(NO_3)_2$. You may ignore the volume change caused by the addition of 1.00 mL. [Q=$8.0x10^{-11}$]

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

$[OH^-]$=(2)(0.0010 L)(0.010 M)/(1.00 L)
= $2.0x10^{-5}$ M

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

K_{sp}=$[Mg^{2+}][OH^-]^2$ =3.45×10^{-11}
Q= $[Mg^{2+}]_A[OH^-]^2_A$
Q= (0.20 M)($2.0x10^{-5}$ M)^2
Q =$8.0x10^{-11}$
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 $[CrO_4^{2-}]$ be to precipitate barium without precipitating strontium. (K_{sp} $BaCrO_4$ =$1.2x10^{-10}$ K_{sp} $SrCrO_4$ =$3.6x10^{-5}$) [$1.6x10^{-7}$ M < $[CrO_4^{2-}]$ < $3.3x10^{-4}$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 of Ba^{2+} = (0.100 g)/137.33g/mol)
= $7.3x10^{-4}$ mol
$[Ba^{2+}]$=($7.3x10^{-4}$ mol)/1.0 L
= $7.3x10^{-4}$ M
$BaCrO_{4\,(s)} \rightleftharpoons Ba^{2+}_{(aq)} + CrO_4^{2-}{}_{(aq)}$
K_{sp}= $[Ba^{2+}][CrO_4^{2-}]$
$1.2x10^{-10}$ =[$7.3x10^{-4}$ M] $[CrO_4^{2-}]$
$[CrO_4^{2-}]$=$1.6x10^{-7}$ M

For Sr^{2+}

n of Ba^{2+} = (10.0 g)/87.62g/mol)
= 0.11 mol
$[Ba^{2+}]$=(0.11 mol)/1.0 L
= 0.11 M
$SrCrO_{4\,(s)} \rightleftharpoons Sr^{2+}_{(aq)} + CrO_4^{2-}{}_{(aq)}$
K_{sp}= $[Sr^{2+}][CrO_4^{2-}]$
$3.6x10^{-5}$ =[0.11 M] $[CrO_4^{2-}]$
$[CrO_4^{2-}]$=$3.3x10^{-4}$ M

Therefore $1.6x10^{-7}$ M < $[CrO_4^{2-}]$ < $3.3x10^{-4}$M

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 $1.3x10^{-8}$. [1.3 mL]

$PbSO_{4\,(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + SO_4{}^{2-}{}_{(aq)}$
$K_{sp}= [Pb^{2+}][SO_4{}^{2-}]$
$1.3x10^{-8}=(0.0010\text{ M})\,[SO_4{}^{2-}]$
$[SO_4{}^{2-}]=1.3x10^{-5}$

To determine the volume of 0.01M Na_2SO_4
$c_1v_1=c_2v_2$
$v_1=c_2v_2/c_1$
$v_1=(1.3x10^{-5})(1.00L)/(0.010\text{ M})$
$v_1=1.3x10^{-3}$ L = 1.3 mL

19. Compare the molar solubility of PbI_2 in a) pure water and b) in 0.10 M NaI. [a) $1.3x10^{-3}$, b) $8.5x10^{-7}$]

a) in pure water
$PbI_{2(s)} \rightleftharpoons Pb^{2+}{}_{(aq)} + 2\,I^-{}_{(aq)}$
x 2x
$K_{sp}= [Pb^{2+}][I^-]^2$
$8.5x10^{-9}= 4x^3$
$x= 1.3x10^{-3}$
Therefore the molar solubility in water is $1.3x10^{-3}$ M.

b) in 0.10 M NaI $K_{sp}= 8.5x10^{-9}$

	$PbI_{2(s)}$ $\rightleftharpoons$	$Pb^{2+}{}_{(aq)}$ +	$2\,I^-{}_{(aq)}$
MR	1	1	2
I	------	0	0.10
C	-------	+x	+x
E	-------	x	0.10+x

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

$K_{sp}= [Pb^{2+}][I^-]^2$
$8.5x10^{-9}= x(0.10+x)^2$
$8.5x10^{-9}= x(0.10+ \cancel{x})^2$
$x= 8.5x10^{-7}$

Therefore the molar solubility in 0.10 NaI is $8.5x10^{-7}$M

20. How many grams of $SrCO_3$ will dissolve in 250 mL of 0.080 mol/L $SrNO_3$? [$2.6x10^{-7}$ g]

	$SrCO_{3(s)}$ $\rightleftharpoons$	$Sr^{2+}{}_{(aq)}$ +	$CO_3{}^{2-}{}_{(aq)}$
MR	1	1	1
I	------	0.080	0
C	-------	+x	+x
E	-------	0.080+x	x

$0.080/K_{sp}$ >1000, approx. works $K_{sp}=5.6x10^{-10}$

$K_{sp}= [Sr^{2+}][CO_3{}^{2-}]$
$5.6x10^{-10}=(0.080+ \cancel{x})(x)$
$x= 7.0x10^{-9}$

n=cv
$=(7.0x10^{-9})(0.250\text{ L})$
$= 1.8x10^{-9}$ mol

MM=147.63 g/mol
m=nMM
$m=(1.8x10^{-9}\text{ mol})(147.63\text{ g/mol})$
$m=2.6x10^{-7}$ 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 $[OH^-]=1.0x10^{-4}$.
   a) $Zn(OH)_2$, $K_{sp}=4.3x10^{-17}$ [$2.8x10^{-7}$mg/mL]

$Zn(OH)_{2(s)} \rightleftharpoons Zn^{2+}{}_{(aq)} + 2\,OH^-{}_{(aq)}$
$K_{sp}=[Zn^{2+}][OH^-]^2$
$4.3x10^{-17}=[Zn^{2+}](1.0x10^{-4})^2$
$[Zn^{2+}]=4.3x10^{-9}$ M

sol = $(4.3x10^{-9}\text{ mol/L})(65.38\text{ g/mol})$
= $2.8x10^{-7}$ g/L
=$2.8x10^{-7}$ mg/mL

b) $Fe(OH)_3$, $K_{sp}=2.6x10^{-39}$ [$1.5x10^{-25}$ mg/mL]

$Fe(OH)_{3(s)} \rightleftharpoons Fe^{3+}_{(aq)} + 3\ OH^-_{(aq)}$
$K_{sp}=[Fe^{3+}][OH^-]^3$
$2.6x10^{-39}=[Fe^{3+}](1.0x10^{-4})^3$
$[Fe^{3+}]=2.6x10^{-27}$ M

sol = ($2.6x10^{-27}$mol/L)(55.85 g/mol)
= $1.5x10^{-25}$ g/L
=$1.5x10^{-25}$ mg/mL

c) $Mg(OH)_2$, $K_{sp}=5.6x10^{-12}$ [$1.4x10^{-2}$ mg/mL]

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

$K_{sp}=[Mg^{2+}][OH^-]^2$
$5.6x10^{-12}=[Mg^{2+}](1.0x10^{-4})^2$
$[Mg^{2+}]=5.6x10^{-4}$ M

sol = ($5.6x10^{-4}$ mol/L)(24.31 g/mol)
= $1.4x10^{-2}$ g/L
=$1.4x10^{-2}$ mg/mL

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

a) Br^- [$4.2x10^{-11}$]

n of Br^- = 0.0010 g/79.90 g/mol
= 1.3 $x10^{-5}$ mol
$[Br^-]=1.3x10^{-5}/0.0010L$
= $1.3x10^{-2}$ M

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$
$K_{sp}= [Ag^+][Br^-]$
$5.4x10^{-13}=[Ag^+][1.3x10^{-2}\ M]$
$[Ag^+]=4.2x10^{-11}$ M

b) S^{2-} ($K_{sp}=1.8x10^{-50}$) [$7.6x10^{-25}$]

n of S^{2-} = 0.0010 g/32.06 g/mol
= 3.1 $x10^{-5}$ mol
$[S^{2-}]=3.1x10^{-5}/0.0010L$
= $3.1x10^{-2}$ M

$Ag_2S_{(s)} \rightleftharpoons 2\ Ag^+_{(aq)} + S^{2-}_{(aq)}$
$K_{sp}= [Ag^+]^2[S^{2-}]$
$1.8x10^{-50}=[Ag^+]^2[3.1x10^{-2}\ M]$
$[Ag^+]=7.6x10^{-25}$ M

c) BrO_3^- [$6.8x10^{-3}$]

n of BrO_3^-= 0.0010 g/127.90 g/mol
= 7.8 $x10^{-6}$ mol
$[BrO_3^-]=7.8x10^{-6}/0.0010L$
= $7.8x10^{-3}$ M

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + BrO_3^-{}_{(aq)}$
$K_{sp}= [Ag^+][BrO_3^-]$
$5.3x10^{-5}=[Ag^+][7.8x10^{-3}\ M]$
$[Ag^+]=6.8x10^{-3}$ M

d) CrO_4^{2-} [$1.2x10^{-5}$]

n of CrO_4^{2-}= 0.0010 g/116.00 g/mol
= $8.6x10^{-6}$ mol
$[CrO_4^{2-}]=8.6x10^{-6}/0.0010L$
= $8.6x10^{-3}$ M

$Ag_2CrO_{4\ (s)} \rightleftharpoons 2\ Ag^+_{(aq)} + CrO_4^{2-}{}_{(aq)}$
$K_{sp}= [Ag^+]^2[CrO_4^{2-}]$
$1.1x10^{-12}=[Ag^+]^2[7.8x10^{-3}\ M]$
$[Ag^+]=1.2x10^{-5}$ M

23. What is the solubility in mol/L of AgBr in a solution resulting from the addition of 50.0 mL of 0.010 $CaBr_2$ to 50.0 mL of 0.0080 M $AgNO_3$? (K_{sp}=5.4x10^{-13}) [9.0x10^{-11}]

	2 $AgNO_{3\,(aq)}$ +	$CaBr_{2(aq)}$ → 2 $AgBr_{(s)}$ + $Ca(NO_3)_{2\,(aq)}$
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} Limiting	5.0x10^{-4}/1 =5.0x10^{-4}

Excess $CaBr_2$ = 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

	$AgBr_{(s)}$ ⇌	$Ag^+_{(aq)}$ +	$Br^-_{(aq)}$
MR	1	1	1
I	---	0	6.0x10^{-3}
C	---	+x	+x
E	---	x	6.0x10^{-3}+x

K_{sp}= 5.4x10^{-13}
0.0060/Ksp>1000
Approx works

K_{sp}= [Ag^+][Br^-]
5.4x10^{-13}=(x)(6.0x10^{-3}+ $\cancel{x}$)
x=9.0x10^{-11}

The solubility of AgBr is 9.0x10^{-11}

24. If 50.0 mL of 0.10 M $AgNO_3$ is added to 150 mL of 0.10 M $CaCl_2$, what is the resulting concentration of each ion in the final solution? (K_{sp}=1.8x10^{-10}) [Ag^+ =1.4x10^{-9}M, Cl^-=0.125 M, NO_3^-=0.025M, Ca^{2+}=0.075 M]

	2 $AgNO_{3\,(aq)}$ +	$CaCl_{2(aq)}$ → 2 $AgCl_{(s)}$ + $Ca(NO_3)_{2\,(aq)}$
c	0.10	0.10
V	0.050	0.150
n	0.0050	0.015
LF	0.0050/2= 0.0025 Limiting	0.015/1 =0.015

Excess $CaCl_2$ = 0.015-0.0025 (amount used)
= 0.0125 mol

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

	$AgCl_{(s)}$ ⇌	$Ag^+_{(aq)}$ +	$Cl^-_{(aq)}$
MR	1	1	1
I	---	0	0.125
C	---	+x	+x
E	---	x	0.125+x

K_{sp}= 1.8x10^{-10}
0.125/Ksp>1000
Approx works

K_{sp}= [Ag^+][Br^-]
1.8x10^{-10}=(0.125+ $\cancel{x}$)(x)
x=1.4x10^{-9}

[Ag^+] = 1.4x10^{-9} M [Cl^-]= 0.125 M
[NO_3^-]= (0.050 L)(0.10 M)/0.200 L = 0.025 M
[Ca^{2+}]= (0.150 L)(0.10 M)/0.200L = 0.075 M

25. What volume of 0.10 M $CaBr_2$ must be added to 100.0 mL of 0.10 M of $Pb(NO_3)_2$, before a precipitate of $PbBr_2$ (K_{sp}=1.4x10^{-8}) starts to form. Assume the total volume remains at 100.0 mL. [0.19 mL]

$PbBr_{2\,(s)}$ ⇌ $Pb^{2+}_{(aq)}$ + 2 $Br^-_{(aq)}$
K_{sp}= [Pb^{2+}][Br^-]^2
1.4x10^{-8}=(0.10 M)[Br^-]^2
[Br^-]=3.7x10^{-4}

To determine the volume of 0.10 M $CaBr_2$
$c_1V_1=c_2V_2$
$V_1=c_2V_2/c_1$
V_1=(3.7x10^{-4})(0.100 L)/2(0.10 M)
V_1=1.9x10^{-4} L = **0.19 mL**

26. How many litres of water at 25°C must be added to mercury(II) sulfide (K_{sp}=3.0x10^{-54}) in order for 1 mercury atom to be present in solution? (HINT: You will need Avogadro's number) [1000 L]

$HgS_{(s)}$ ⇌ $Hg^{2+}_{(aq)}$ + $S^{2-}_{(aq)}$
x x
K_{sp}= [Hg^{2+}][S^{2-}]
3.0x10^{-54}= x^2
x=1.7x10^{-27}
[Hg^{2+}]=1.7x10^{-27}

n Hg^{2+} = 1 atom/N_A (6.02x10^{23})
= 1.7x10^{-24} mol

c=n/V
V=n/c
V=1.7x10^{-24} mol/1.7x10^{-27}
V=1.0x10^3 L

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, Ba^{2+} =0.333 M, Cl^- =1.33 M, SO_4^{2-}= 3.33 X10^{-10} M, Na^+=0.667 M]

	$Na_2SO_{4\,(aq)}$ +	$BaCl_{2\,(aq)}$ → 2 $NaCl_{(aq)}$ + $BaSO_{4\,(s)}$
c	1.00	1.00
V	0.100	0.200
n	0.100	0.200
LF	0.100/1= 0.100	0.200/2 =0.200
	Limiting	

Excess $BaCl_2$ = 0.200-0.100 (amount used)
= 0.100 mol
Initial [Ba^{2+}] = 0.100 mol/0.300 L = 0.33 M

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

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

K_{sp}= 1.1 x 10^{-10}
0.33/Ksp>1000
Approx works

K_{sp}= [Ba^{2+}][SO_4^{2-}]
1.1 x 10^{-10}=(0.33+ $\cancel{x}$)(x)
x=3.3x10^{-10}

To find mass of ppt
n of $BaSO_4$ made = 0.100 mol
n dissolved = (3.3x10^{-10} M)(0.300 L) = 9.9x10^{-11} mol

n of ppt = 0.100 – 9.9x10^{-11} mol
= 0.100 mol

m = n MM
= (0.100 mol)(233.39 g/mol)
= 23.3 g

[Ba^{2+}] = 0.33 M   [SO_4^{2-}]= 3.3x10^{-10} M
[Cl^-]= 2(0.20 L)(1.00 M)/0.300 L = 1.33 M
[Na^+]= 2(1.00 M)(0.100 L)/0.300L = 0.67 M

28. A 1.50 L sample of 0.250 M NaOH is added to 1.00 L of 0.150 M $Mg(NO_3)_2$. Calculate the mass of $Mg(OH)_2$ that precipitates and the concentration of all ions in solution at equilibrium. [m=8.75 g, Mg^{2+} =6.22x10^{-9} M, OH^-=0.0300 M, NO_3^-= 0.120 M, Na^+=0.150 M]

2 $NaOH_{(aq)}$ + $Mg(NO_3)_{2\,(aq)}$ → 2 $NaNO_{3\,(aq)}$ + $Mg(OH)_{2\,(s)}$

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 NaOH = 0.375-0.300 (amount used)
= 0.075 mol
Initial [OH^-] = 0.075 mol/2.50 L = 0.0300 M

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

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

K_{sp}= 5.6 x 10^{-12}
0.0300/Ksp>1000
Approx works

K_{sp}= [Mg^{2+}][OH^-]^2
5.6 x 10^{-12}=(x)(0.0300+ $\cancel{2x}$)^2
x=6.22x10^{-9}

To find mass of ppt
n of $Mg(OH)_2$ made = 0.150 mol
n dissolved = (6.22x10^{-9} M)(2.50 L) = 1.6x10^{-8} mol

n of ppt = 0.150 – 1.6x10^{-8} mol
= 0.150 mol

m = n MM
= (0.150 mol)(58.33 g/mol)
= **8.75 g**

[Mg^{2+}] = 6.22x10^{-9}M   [OH^-]= 0.0300 M
[NO_3^-]= 2(1.00 L)(0.150 M)/2.50 L = 0.120 M
[Na^+]= (1.50 L)(0.250 M)/ 2.50 L = 0.150 M