Chapter 15 **Extra Credit** **Name: ______________Key_______________**

1. Write the ionic equation for dissolution and the solubility product (K_{sp}) expression for each of the following slightly soluble ionic compounds:

(a) $PbCl_2$ $PbCl_2\,(s) \rightleftarrows Pb^{2+}\,(aq) + 2\,Cl^-\,(aq)$ $Ksp = [Pb^{2+}][Cl^-]^2$

(b) Ag_2S

(c) $Sr_3(PO_4)_2$

2. Use solubility products and predict which of the following salts is the most soluble, in terms of moles per liter, in pure water: CaF_2, Hg_2Cl_2, PbI_2, or $Sn(OH)_2$.

CaF_2 $Ksp = 4.0 \times 10^{-11}$

Hg_2Cl_2 $Ksp = 1.1 \times 10^{-18}$

PbI_2 $Ksp = 1.4 \times 10^{-8}$

$Sn(OH)_2$ $Ksp = 3 \times 10^{-27}$

Each of these break into a total of 3 ions (even mercury (I) chloride) so the Ksp's can be directly compared,

$Sn(OH)_2$ is less soluble than $Hg_2Cl_2 < CaF_2 < PbI_2$

3. Calculate the molar solubility of each. Look up Ksp in the appendix.

(a) PbI_2 $Ksp = 1.4 \times 10^{-8}$

$X = 0.0015\ M$

(b) Ag_2SO_4

Ksp = 1.2 x 10^{-5}

X = 0.014 M

4. Given the molar solubility, calculate K_{sp} for each of the slightly soluble solids indicated:

(a) AgBr: $x = 5.7 \times 10^{-7}$ *M*, Ksp = x^2

Ksp = 3.2 x 10^{-13}

(b) PbF_2: $x = 2.1 \times 10^{-3}$ *M*, Ksp = $x(2x)^2$ Ksp = 3.7 x 10^{-8}

5. The *Handbook of Chemistry and Physics* gives solubilities of the following compounds in grams per 100 mL of solution. Calculate the solubility product (Ksp) for each.

(a) $BaSiF_6$, 0.026 g/100 mL (contains SiF_6^{2-}) x = 0.00093M Ksp = 8.7×10^{-7}

(b) $Ce(IO_3)_4$, 1.5×10^{-2} g/100 mL MM $Ce(IO_3)_4$ = 840 g/mole

X = 0.00018 M Ksp = $[Ce^{4+}][IO_3^-]^4 = x(4x)^4 = 256\,(x)^5$

Ksp = 4.8×10^{-17}

6. Calculate the molar solubility of:

(a) AgCl(*s*) in pure water. Ksp = 1.6×10^{-10}

x = 1.3×10^{-5} M

AgCl(*s*) in 0.010 *M* NaCl

x = 1.6×10^{-8} M

How does the solubility change when a common ion is added? The solubility decreases.

(b) Calculate the molar solubility of: CaF_2(*s*) in 0.0125 *M* KF Ksp = 4.0×10^{-11} = $[Ca^{2+}][F^-]^2$ = $x(.0125 + 2x)^2$ = $x(.0125)^2$

$x = 2.6 \times 10^{-7}$ M

(c) Calculate the molar solubility of: $Ni(OH)_2$(*s*) in a solution with pH of 12.00. Ksp = 1.6×10^{-16}

$[OH^-] = 10^{-2}$ Ksp = $[Ni^{2+}][OH^-]^2$ = $(x)(.01 + 2x)^2$ = $x(.01)^2$ $x = 1.6 \times 10^{-12}$ M

7. Will a precipitate form given the concentrations indicated? (See appendix for K_{sp} values.)

(a) $CaCO_3$: $[Ca^{2+}] = 0.0020$ *M*, $[CO_3^{2-}] = 0.010$M Ksp = 8.7×10^{-9}

Q = 2.0×10^{-5} Q>Ksp, so a precipitate will form.

(b) $Mn(OH)_2$: $[Mn^{2+}] = 1.0 \times 10^{-4}$ *M*, $[OH^-] = 1.0 \times 10^{-5}$ *M*

Q = 1.0×10^{-14} Q<Ksp, a precipitate will not form.

8. Draw the Lewis Structure for each and Label the Lewis Acids and the Lewis Bases (reactants only)

(a) $BF_3 + F^- \longrightarrow BF_4^-$

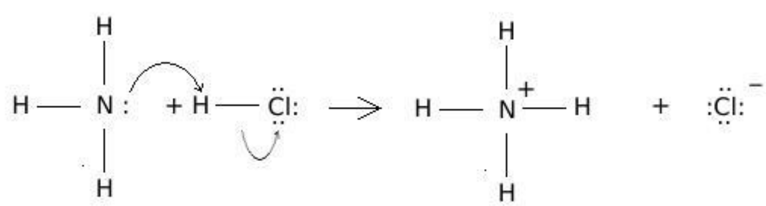
LA LB

(b) $Al(OH)_3 + OH^- \longrightarrow Al(OH)_4^-$

LA LB

Draw the Lewis Structure for each and Label the Lewis Acids and the Lewis Bases (reactants only) and predict the Products.

(a) $NH_3(g) + HCl(g) \longrightarrow$



LB LA

(b) $NH_4^+ + C_2H_5O^- \longrightarrow$
LA LB

9. A volume of .080 L of 2.0×10^{-3} M $Ba(NO_3)_2$ (aq) is added to .020 L of 5.0×10^{-3} M Li_2SO_4 (aq). Will a precipitate form?

For: $BaSO_4$: Ksp = 2.3×10^{-8}

Q = 1.6×10^{-6}. A precipitate will form.