UNIVERSITY OF
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COLLEGE OF ARTS AND SCIENCE
DEPARTMENT OF CHEMISTRY

CHEM131 HOMEWORK #14 KEY

8-71a. s = Ag_3PO_4 solubility, $[Ag^+] = 3s$.

	$Ag_3PO_4\ (s) \leftrightarrows$	$3Ag^+(aq)$	+ $PO_4^{3-}(aq)$
Initial		0	0
Change		+3s	+s
Equil.		3s	s

$K_{sp} = (3s)^3 s = 1.8*10^{-18} = 27s^4$

$s = 1.6*10^{-5}$ M = $1.6*10^{-5}$ mol/L*419 g/mol = $6.7*10^{-3}$ g/L.

All molar solubility problems proceed identically….

8-71b. s = $CaCO_3$ solubility, $[Ca^{2+}] = s$, $[CO_3^{2-}] = s$, $8.7*10^{-9} = s^2$,
$s = 9.3*10^{-5}$ M = $9.3*10^{-3}$ g/L.

8-71c. s = solubility of Hg_2Cl_2, $[Hg_2^{2+}] = s$, $[Cl^-] = 2s$, $1.1*10^{-18} = s*(2s)^2 = 4s^3$,
$s = 6.5*10^{-7}$ M = $3.1*10^{-4}$ g/L.

8-73a. $[Ca^{2+}] = 4.8*10^{-5}$ M; $[C_2O_4^{2-}] = 4.8*10^{-5}$; $K_{sp} = [Ca^{2+}][C_2O_4^{2-}] = 4.8*10^{-5} x 4.8*10^{-5} = 2.3*10^{-9}$.

8-73b. $[Pb^{2+}] = 2.14*10^{-2}$; $[Br^-] = 2[Pb^{2+}] = 4.28*10^{-2}$; $K_{sp} = [Pb^{2+}][Br^-]^2 = 2.14*10^{-2}$ x $1.83*10^{-3} = 3.92*10^{-5}$

8-73c. $K_{sp} = 1.32*10^{-5} x (3*1.32*10^{-5})^3 = 8.20*10^{-19}$

8-73d. $65.9*10^{-3}$g/L = $4.58*10^{-4}$ mol/L; $K_{sp} = (4.58*10^{-4})^2 = 2.10*10^{-7}$.

8-73e. 0.146g/0.1L = 1.46g/L = $3.28*10^{-3}$ M; $K_{sp} = 3.28*10^{-3}$ x $(2*3.28*10^{-3})^2 = 1.41*10^{-7}$.

8.84. We start with 50 mL x 0.10 M = 5.0 mmol Pb^{2+} and 50.0 mL x 1.0 M = 50 mmol Cl^-
$PbCl_2$ is the only possible precipitate, and the initial concentrations, $[Pb^{2+}]_0 = 0.050$ M, $[Cl^-]_0 = 0.5$ M since the volume doubles on mixing. $Q = 0.050*(0.50)^2 = 0.0125 > K_{sp}$ so a ppt. forms. Let's assume since K_{sp} is so small that the precipitation is complete. This is a stiochiometry problem – let's write out the mols of everything

	$PbCl_2\ (s)$	$\leftarrow$	$Pb^{2+}(aq)$	+ $2Cl^-(aq)$
Before:	0		5.0 mmol	50.0 mmol
Change:	+5.0 mmol		-5.0 mmol	-10.0 mmol
After:	5.0 mmol		0	40.0 mmol

Now do the equilibrium problem of the salt dissolving remembering that we have 100 mL of solution:

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$PbCl_2$ (*s*)	⇆	$Pb^{2+}(aq)$	+	$2Cl^-(aq)$
Initial		0		0.4 M = 40.0mmol/100 mL
Change		+s		+2s
Equil.		s		0.4 +2s

$K_{sp} = (0.4 + 2s)^2 s = 1.6*10^{-5}$ neglect the 2s relative to the 0.4:
$(0.4)^2 s = 1.6*10^{-5}$. $s = 1.0 * 10^{-4} = [Pb^{2+}]$; $[Cl^-] = 0.40 + s = 0.40$ mol/L to 2 sig figs

8.91.
Since the formation or stability constant for HgI_4^{2-} is **very** large, effectively all of the Hg^{2+} will be present in the form of the complex ion, and what we will do is calculate the problem assuming **ALL** the free Hg^{2+} concentration reacts to form the complex ion, then solve the equilibrium problem to find the trace amount of free Hg^{2+}. We solve it in 2 steps.

65 g KI x 1 mol KI/166.0 g x 1/0.5 L = 0.78 M KI = $[I^-]$

	Hg^{2+}	+ I^-	→	HgI_4^{2-}
Before:	5 mmol	390 mmol		0
Change	-5.0 mmol	-20 mmol		+5 mmol
After	0	370 mmol		5 mmol

	Hg^{2+}	+ I^-	⇆	HgI_4^{2-}
Initial	0	0.74 M		0.01 M
Change	X	X		-X
Equilibrium	X	0.74+x		0.01-X

$K_f = 1.0*10^{30} = (0.01-X)/(X*(0.74+X)^4)$
Make assumption that X is small relative to 0.01 and 0.74
$1.0*10^{30} = (0.01)/(X*(0.74)^4)$ or $X = 3.3 * 10^{-32}$ M. Our assumptions of two steps are reasonable since the Hg^{2+} concentration is so small (one ion per 10^7 L)

Next Homework: NONE! Enjoy your break.