

* While I prefer you turn in a hard copy of the worksheet, I will accept scanned copies sent to my email address, joensj@fiu.edu

Section: (circle one) M,W,F Tu,Tr

For problems involving calculations you must show your work for credit. Unless otherwise stated, you may assume $T = 25.0\text{ }^{\circ}\text{C}$.

1) A chemist has the following two solutions

Solution A 0.0428 M HCl (a strong acid)
Solution B 0.1425 M KOH (a strong soluble base)

a) What are the values for pH for solution A and solution B?

The reactions are $\text{HCl(aq)} + \text{H}_2\text{O}(\ell) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$

$\text{KOH(s)} \rightarrow \text{K}^+(\text{aq}) + \text{OH}^-(\text{aq})$

For solution A $[\text{H}_3\text{O}^+] = [\text{HCl}] = 0.0428\text{ M}$ $\text{pH} = -\log_{10}(0.0428) = 1.37$

For solution B $[\text{OH}^-] = [\text{KOH}] = 0.1425\text{ M}$ $\text{pOH} = -\log_{10}(0.1425) = 0.85$

$\text{pH} = 14.00 - \text{pOH} = 14.00 - 0.85 = 13.15$

b) A new solution, solution C, is formed by combining 10.00 mL of solution A with 15.00 mL of solution B. What is the pH for solution C?

The reaction taking place when we add solutions A and B together is

$\text{HCl(aq)} + \text{KOH(aq)} \rightarrow \text{KCl(aq)} + \text{H}_2\text{O}(\ell)$

moles HCl = 0.01000 L soln $\frac{0.0428\text{ mol HCl}}{\text{L}} = 4.28 \times 10^{-4}\text{ mol HCl}$

moles KOH = 0.01500 L soln $\frac{0.1425\text{ mol KOH}}{\text{L}} = 2.138 \times 10^{-3}\text{ mol KOH}$

Based on the number of moles and the reaction, HCl is the limiting reactant, and KOH is in excess. The number of moles of excess KOH is

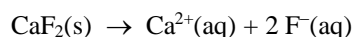
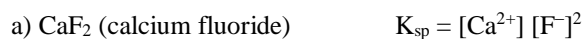
moles excess KOH = $2.138 \times 10^{-3}\text{ mol} - 4.28 \times 10^{-4}\text{ mol} = 1.71 \times 10^{-3}\text{ mol KOH}$ in excess

The volume of the solution is $V = 10.00\text{ mL} + 15.00\text{ mL} = 25.00\text{ mL} = 0.02500\text{ L}$

So $[\text{KOH}]_{\text{excess}} = \frac{1.71 \times 10^{-3}\text{ mol KOH}}{0.02500\text{ L}} = 0.0684\text{ M KOH}$

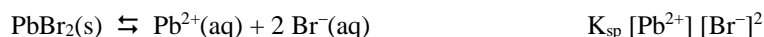
$\text{pOH} = -\log_{10}(0.0684) = 1.16$ $\text{pH} = 14.00 - 1.16 = 12.84$

2) Give the correct expressions for K_{sp} for the following slightly soluble ionic compounds in water.



3) The solubility product for lead II bromide (PbBr_2 , MW = 367.0 g/mol) is $K_{sp} = 4.6 \times 10^{-6}$ at $T = 25.0^\circ\text{C}$.

a) Give the reaction corresponding to PbBr_2 dissolving in water, and the corresponding expression for K_{sp} .



b) What are the molar solubility and solubility by mass for lead II bromide in pure water?

	Initial	Change	Equilibrium
Pb^{2+}	0	x	x
Br^-	0	2x	2x

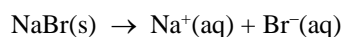
$(x)(2x)^2 = 4x^3 = 4.6 \times 10^{-6}$ $x^3 = (4.6 \times 10^{-6})/4 = 1.15 \times 10^{-6}$
 $x = (1.15 \times 10^{-6})^{1/3} = 1.05 \times 10^{-2}$

molar solubility = $x = 1.05 \times 10^{-2}$ M

solubility by mass = $1.05 \times 10^{-2} \frac{\text{mol}}{\text{L}} \frac{367.0 \text{ g}}{\text{mol}} = 3.85 \text{ g/L}$

c) What are the molar solubility and the solubility by mass for lead II bromide when added to a 0.088 M solution of sodium bromide (NaBr , MW = 102.9 g/mol), a soluble ionic compound?

NaBr is a soluble ionic compound. When added to water it dissociates by the process



Because the solution has $[\text{NaBr}] = 0.088 \text{ M}$, there is an initial concentration of Br^- ion, $[\text{Br}^-] = 0.088 \text{ M}$. Other than that, the problem is the same as in part b.

	Initial	Change	Equilibrium
Pb^{2+}	0	x	x
Br^-	0.088	2x	0.088 + 2x

$(x)(0.088 + 2x)^2 = 4.6 \times 10^{-6}$ Assume $x \ll 0.088$, then

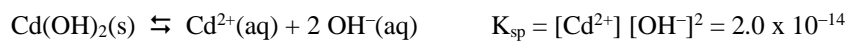
$x(0.088)^2 = 4.6 \times 10^{-6}$ $x = (4.6 \times 10^{-6})/(0.088)^2 = 5.9 \times 10^{-4}$ The assumption that $x \ll 0.088$ was good.

molar solubility = $x = 5.9 \times 10^{-4}$ M

solubility by mass = $5.9 \times 10^{-4} \frac{\text{mol}}{\text{L}} \frac{367.0 \text{ g}}{\text{mol}} = 0.22 \text{ g/L}$

4) What is the molar solubility for cadmium hydroxide ($\text{Cd}(\text{OH})_2$, $\text{MW} = 146.4 \text{ g/mol}$) in a $\text{pH} = 10.00$ buffer solution? The solubility product for cadmium hydroxide is $K_{\text{sp}} = 2.0 \times 10^{-14}$

The solubility reaction is



Normally we would at this point make an ICE table, but since the solution is a buffer solution we can assume the pH is constant.

$$\text{pH} = 10.00 \quad \text{pOH} = 14.00 - 10.00 = 4.00 \quad [\text{OH}^{-}] = 10^{-\text{pOH}} = 10^{-4.00} = 1.0 \times 10^{-4} \text{ M}$$

$$[\text{Cd}^{2+}] [\text{OH}^{-}]^2 = 2.0 \times 10^{-14} \quad [\text{Cd}^{2+}] = \frac{2.0 \times 10^{-14}}{[\text{OH}^{-}]^2} = \frac{2.0 \times 10^{-14}}{(1.0 \times 10^{-4})^2} = 2.0 \times 10^{-6} \text{ M}$$

Since for every Cd^{2+} ion that forms a $\text{Cd}(\text{OH})_2$ must dissociate, the molar solubility is

$$\text{molar solubility} = x = 2.0 \times 10^{-6} \text{ M}$$