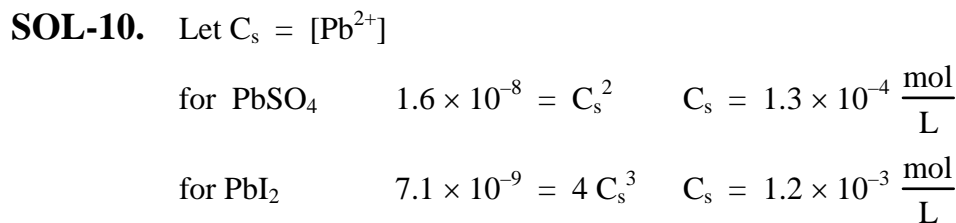
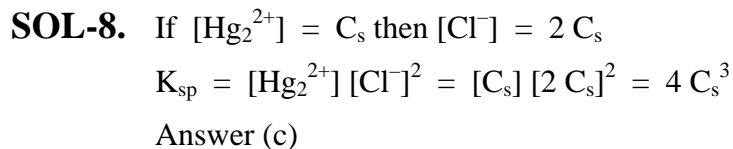
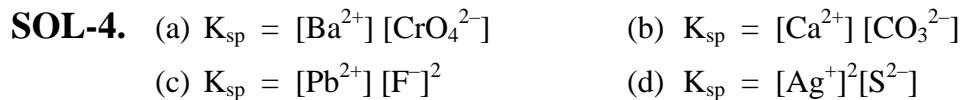
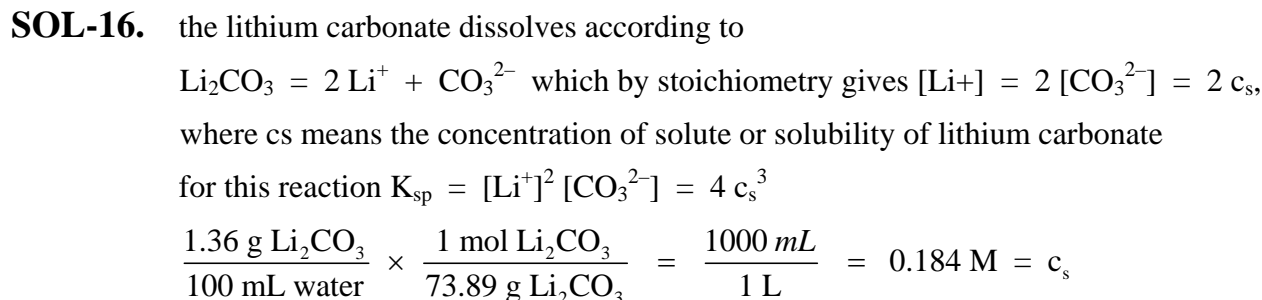


# Homework #11

## solution outlines

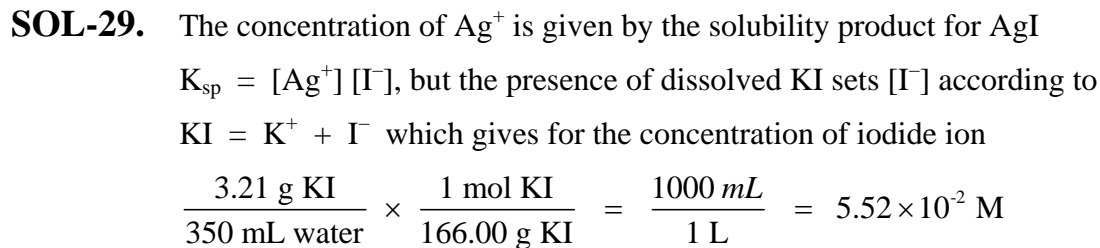
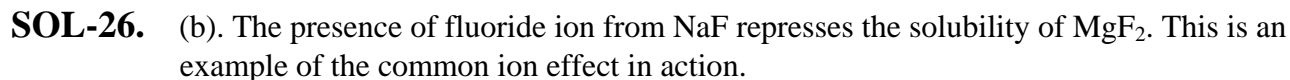


Lead (II) iodide is more soluble.



$$K_{sp} = 2.49 \times 10^{-2}$$

NOTE: By definition, molarity means moles of solute per liter of solution. In this instance we have taken 100 mL water to be 100 mL of solution. Stricly speaking, this is not correct, but the difference between the two is small enough to allow the approximation.



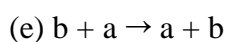
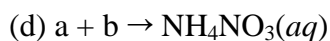
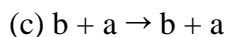
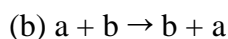
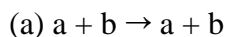
now return to the  $K_{sp}$  for AgI to solve for  $[\text{Ag}^+]$ :

$$[\text{Ag}^+] = K_{\text{sp}} / [\text{I}^-] = 8.3 \times 10^{-17} / 5.52 \times 10^{-2} = 1.50 \times 10^{-15} \text{ M}$$

11.5 (a) and (b) are Arrhenius acids.

11.6 Only (a) is an Arrhenius base.

11.13 In each reaction, let a represent Brønsted acid and b represent Brønsted base.



11.35  $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$

$\therefore \text{pH} = 3.72, \therefore [\text{H}_3\text{O}^+] = 1.9 \times 10^{-4} \text{ M}$

$\therefore \text{pH} + \text{pOH} = 14, \therefore \text{pOH} = 10.28 \text{ and } [\text{OH}^-] = 5.3 \times 10^{-11} \text{ M}.$

11.50 The weakest Brønsted acid is the one with the smallest acid dissociation constant. Of the acids listed, (d) phenol has the smallest  $K_a$  value.

11.52 The strongest base will result from the weakest acid. By this metric, (a)  $\text{CH}_3\text{COO}^-$  is the strongest base of those listed.

11.64 The leveling effect makes all strong acids of the same concentration (molarity) have the same hydronium concentration. Since both nitric acid and hydrobromic acid are strong acids, 0.10 M solutions of each will have the same hydronium concentration. In contrast, hydrofluoric acid ( $\text{HF}(aq)$ ) is not a strong acid, so a 0.10 M solution of it will have a lower hydronium concentration owing to incomplete dissociation.

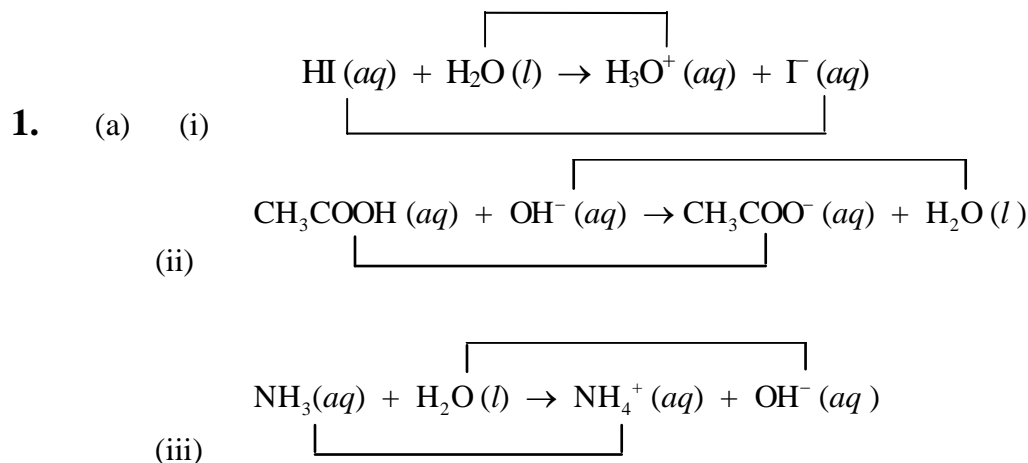
11.85 For a weak acid we can write  $[\text{H}_3\text{O}^+] = (K_a c_{\text{HA}})^{1/2}$ , where  $c_{\text{HA}}$  is the initial concentration of HA.

$\therefore [\text{H}_3\text{O}^+] = \sqrt{(1.8 \times 10^{-4})(0.10)} = 4.2 \times 10^{-3} \text{ M} = [\text{HCOO}^-]$

$\therefore [\text{HCOOH}] = 0.10 - 4.2 \times 10^{-3} = 9.6 \times 10^{-2} \text{ M}$

11.89 Assume that the dissociation of the second proton on ascorbic acid does not contribute to the species in solution.

	$\text{C}_6\text{H}_6\text{OH}_2$	$\text{H}_3\text{O}^+(aq)$	$\text{C}_6\text{H}_6\text{OH}^-$
initially	0.10 M	0	0
at equilibrium	0.10 M - 2.8%(0.10 M)	2.8%(0.10 M)	2.8%(0.10 M)
$K_a$	$= \frac{[\text{C}_6\text{H}_6\text{OH}^-][\text{H}_3\text{O}^+]}{[\text{C}_6\text{H}_6\text{OH}_2]} = \frac{(0.0028)^2}{(0.100 - 0.0028)} = 8.1 \times 10^{-5}$		



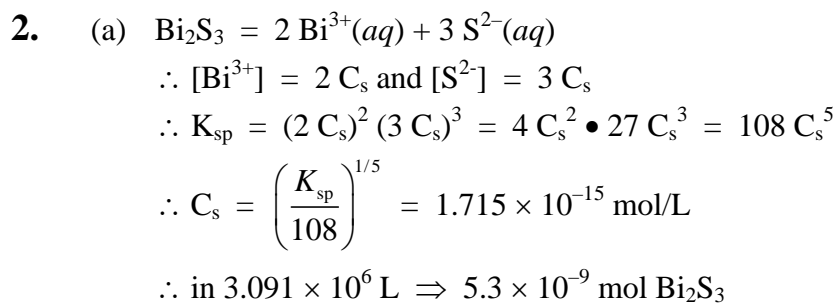
(b)  $\text{H}_3\text{O}^+$ ,  $\text{AlCl}_4^-$ , and  $\text{SiH}_4$ .  $\text{SiH}_4$  and  $\text{AlCl}_4^-$  cannot accept a proton due to lack of unpaired electrons. While  $\text{H}_3\text{O}^+$  does have a pair of nonbonding electrons and so, in principle, could accept another proton,  $\text{H}_4\text{O}^{2+}$  is not observed.

(c) HI is a strong acid  $\Rightarrow$  complete dissociation

$$\therefore 0.03091 \text{ M HI}(aq) \Rightarrow 0.03091 \text{ M} = [\text{H}^+] = [\text{I}^-]$$

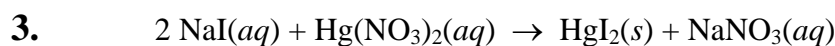
$$\therefore \text{pH} = -\log_{10}[\text{H}^+] = -\log_{10} 0.03091 = 1.51$$

$$\text{pOH} + \text{pH} = 14 \Rightarrow \text{pOH} = 12.49$$



(b)  $[\text{Bi}^{3+}] = 2 C_s = 1.06 \times 10^{-8} \text{ mol}$

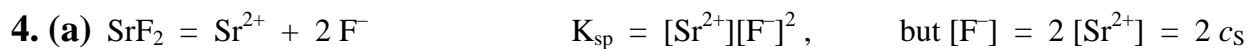
$N_{\text{Av}}$  ions/mol  $\Rightarrow 6.38 \times 10^{15} \text{ Bi}^{3+}$  ions in the  $3.091 \times 10^6$  liters of water of part (a)



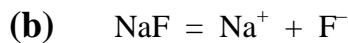
$$\frac{0.10 \text{ mol Hg}(\text{NO}_3)_2}{1 \text{ L}} \times 0.045 \text{ L} = 4.5 \times 10^{-3} \text{ mol Hg}(\text{NO}_3)_2$$

$$4.5 \times 10^{-3} \text{ mol Hg}(\text{NO}_3)_2 \times \frac{2 \text{ mol NaI}}{1 \text{ mol Hg}(\text{NO}_3)_2} = 9.00 \times 10^{-3} \text{ mol NaI}$$

$$\frac{9.00 \times 10^{-3} \text{ mol NaI}}{0.25 \frac{\text{mol NaI}}{\text{L}}} = 3.6 \times 10^{-2} \text{ L} \times \frac{1000 \text{ mL}}{1 \text{ L}} = 36 \text{ mL NaI}$$



$$\therefore K_{sp} = c_s (2 c_s)^2 = 4 c_s^3 \qquad \therefore c_s = \left( \frac{K_{sp}}{4} \right)^{1/3} = 8.49 \times 10^{-4} M$$



$\therefore [\text{F}^-] = 0.03 \text{ M}$ , which dominates the other equilibrium

$$\therefore K_{sp} = [\text{Sr}^{2+}][\text{F}^-]^2 \Rightarrow [\text{Sr}^{2+}] = \frac{K_{sp}}{[\text{F}^-]^2} = \frac{2.45 \times 10^{-9}}{(0.03)^2} = 2.72 \times 10^{-6} M$$

$$\therefore [\text{Sr}^{2+}] = c_s = 2.72 \times 10^{-6} M$$