3.091 OCW Scholar

Self-Assessment
Aqueous Solutions

Supplemental Exam Problems for Study
Solutions Key

Problem #1

(a) Identify the conjugate acid-base pairs in each equilibrium by drawing a line connecting each acid with its conjugate base, and identify the acid of each acid/base pair:

acid   acid

(i) $CH_3CO_2H(aq) + NH_3(aq) \rightleftharpoons CH_3CO_2^-(aq) + NH_4^+(aq)$

acid

(ii) $SbF_5(aq) + 2\,HF(aq) \rightleftharpoons H_2F^+(aq) + SbF_6^-(aq)$

acid

* There is another way to Think about This

(b) A 1.11 M solution of fluoroacetic acid, FCH_2CO_2H, is 5% dissociated in water.

(i) Calculate the value of the pK_a of FCH_2CO_2H.

let x be the extent of dissociation

$$FCH_2CO_2H + H_2O = FCH_2CO_2^- + H_3O^+$$

$$\therefore K_a = \frac{[FCH_2CO_2^-][H_3O^+]}{[FCH_2CO_2H]} = \frac{x^2}{1-x}$$ where $x = (0.05)(1.11) = 0.0555$

$$\text{so } K_a = \frac{(0.0555)^2}{1-0.0555} = 3.26 \times 10^{-3} \Rightarrow pK_a = -\log_{10}(K_a) = 2.49$$

(ii) Calculate the value of the pH of the solution.

$$pH = -\log_{10}[H_3O^+] = -\log_{10}[1.11M \times 0.05] = 1.26$$

* split the HF into 2 separate molecules:

acid base acid base

$$SbF_5 + HF + HF = H_2F^+ + SbF_6^-$$

Problem #2

Chromium hydroxide ($Cr(OH)_3$) dissolves in water according to

$$Cr(OH)_3 \;=\; Cr^{3+}(aq) \;+\; 3\,OH^{-}(aq) \qquad\qquad K_{sp} = 6.31 \times 10^{-31} \text{ at } 25^{\circ}C$$

Calculate the solubility of chromium hydroxide in 3.091 nM (3.091×10^{-9} M) NaOH(*aq*). Express your answer in moles of $Cr(OH)_3$ per liter of solution.

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

with 3.091 nM NaOH $\Rightarrow [OH^-] = 3.091 \times 10^{-9} M$

$$\therefore [Cr^{3+}] = K_{sp}/[OH^-]^3 = 6.31 \times 10^{-31}/(3.091 \times 10^{-9})^3$$

$$= 2.14 \times 10^{-5}$$

$\therefore$ Solubility of $Cr(OH)_3 = [Cr^{3+}] = 2.14 \times 10^{-5} M$

Problem #3

Comment on the solubility of iodine (I_2) in each of these ***liquids***: (1) carbon tetrachloride (CCl_4); (2) hydrogen fluoride (HF). State whether at room temperature you expect I_2 to be ***highly soluble*** or ***almost insoluble***, and explain why.

(1) I_2 in $CCl_4(\ell)$

Highly soluble because both I_2 and CCl_4 are nonpolar and therefore have the capability of mixing.

(2) I_2 in $HF(\ell)$

Amost insoluble because I_2 is a homopolar molecule and therefore nonpolar while HF is polar with hydrogen-bonding capability.

Problem #4

(a) The water dissociation equilibrium constant, K_w, expresses the relationship between hydronium (H_3O^+) and hydroxyl (OH^-) concentrations by the expression

$$K_w = [H_3O^+][OH^-]$$

Owing to the presence of dissolved salts the value of pK_w for seawater is 13.776 (not 14.00 as it is for pure water), where pK_w is defined as $-\log_{10}K_w$. Calculate the concentration of hydroxyl ions (OH^-) in seawater at a pH value of 7.00. Express your answer in moles OH^- per liter of solution (M).

$$K_w = [H^+][OH^-] = 10^{-13.776}$$

$$pH = 7.00 \Rightarrow [H^+] = 10^{-7}$$

$$\therefore [OH^-] = K_w / [H^+] = 10^{-13.776} / 10^{-7}$$

$$= 1.67 \times 10^{-7}\ M$$

(b) Would seawater at a pH value of 7.00 be classified acidic, basic, or neutral? Explain.

alkaline: pH of 7.00 is more alkaline than 13.776 / 2 = 6.89.

(c) Give an example of a dissolved salt that would cause the shift in the value of pK_w for seawater to 13.776 from the commonly accepted value of 14.00 which is valid for pure water. Justify your choice of salt.

Choose any hydroxide such as NaOH or $Ca(OH)_2$ the presence of OH^- operates through the common ion effect to shift the neutrality point of the acid-base equilibrium

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