

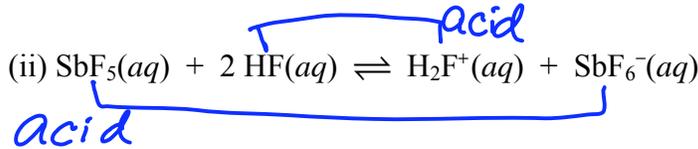
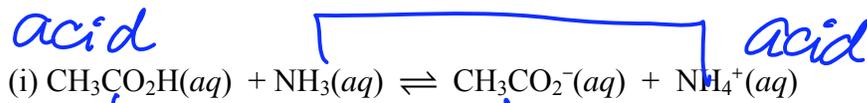
3.091 OCW Scholar

Self-Assessment Aqueous Solutions

Supplemental Exam Problems for Study Solutions Key

Problem #1

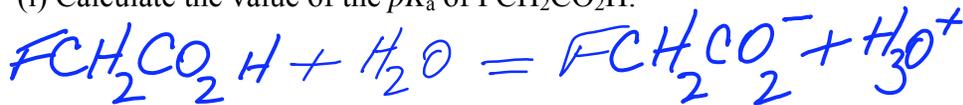
- (a) Identify the conjugate acid-base pairs in each equilibrium by drawing a line connecting each acid with its conjugate base, and identify the acid of each acid/base pair:



* there is another way to think about this

- (b) A 1.11 M solution of fluoroacetic acid, $\text{FCH}_2\text{CO}_2\text{H}$, is 5% dissociated in water.

- (i) Calculate the value of the pK_a of $\text{FCH}_2\text{CO}_2\text{H}$.



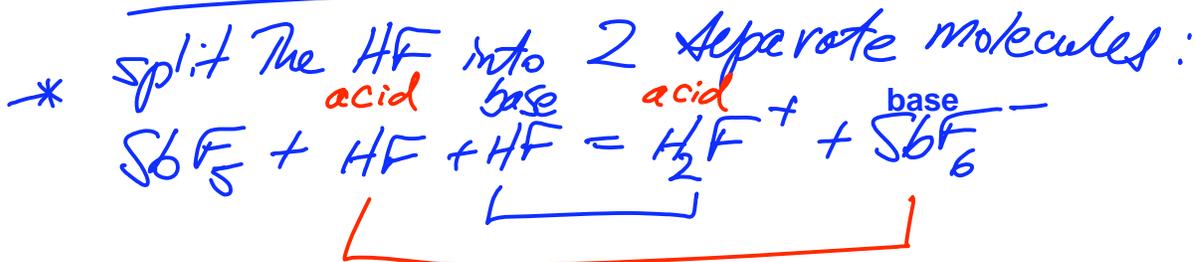
let x be the extent of dissociation

$$\therefore K_a = \frac{[\text{FCH}_2\text{CO}_2^-][\text{H}_3\text{O}^+]}{[\text{FCH}_2\text{CO}_2\text{H}]} = \frac{x^2}{1-x} \quad \text{where } x = (0.05)(1.11) = 0.0555$$

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

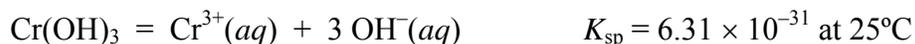
- (ii) Calculate the value of the pH of the solution.

$$pH = -\log_{10}[\text{H}_3\text{O}^+] = -\log_{10}[1.11 \text{ M} \times 0.05] = 1.26$$



Problem #2

Chromium hydroxide ($\text{Cr}(\text{OH})_3$) dissolves in water according to



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

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

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

$$\therefore [\text{Cr}^{3+}] = K_{\text{sp}} / [\text{OH}^{-}]^3 = 6.31 \times 10^{-31} / (3.091 \times 10^{-9})^3$$
$$= 2.14 \times 10^{-5}$$

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

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 $\text{CCl}_4(\ell)$

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

(2) I_2 in $\text{HF}(\ell)$

Almost 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

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$

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).

$$\begin{aligned}K_w &= [\text{H}^+][\text{OH}^-] = 10^{-13.776} \\pH = 7.00 &\Rightarrow [\text{H}^+] = 10^{-7} \\ \therefore [\text{OH}^-] &= K_w / [\text{H}^+] = 10^{-13.776} / 10^{-7} \\ &= 1.67 \times 10^{-7} \text{ M}\end{aligned}$$

- (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 $\text{Ca}(\text{OH})_2$ the presence of OH^- operates through the common ion effect to shift the neutrality point of the acid-base equilibrium

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