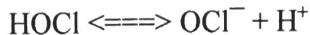


### Equilibrium Practice #3

1)  $K_a$ ,  $K_b$

20 minutes



Hypochlorous acid, HOCl, is a weak acid commonly used as a bleaching agent. The acid-dissociation constant,  $K_a$ , for the reaction represented above is  $3.2 \times 10^{-8}$ .

(a) Calculate the  $[\text{H}^+]$  of a 0.14-molar solution of HOCl.

(b) Write the correctly balanced net ionic equation for the reaction that occurs NaOCl is dissolved in water and calculate the numerical value of the equilibrium constant for the reaction.

(c) Calculate the pH of a solution made by combining 40.0 milliliters of 0.14-molar HOCl and 10.0 milliliters of 0.56-molar NaOH.

(d) How many millimoles of solid NaOH must be added to 50.0 milliliters of 0.20-molar HOCl to obtain a buffer solution that has a pH of 7.49? Assume that the addition of the solid NaOH results in a negligible change in volume.

(e) Household bleach is made by dissolving chlorine gas in water, as represented below.

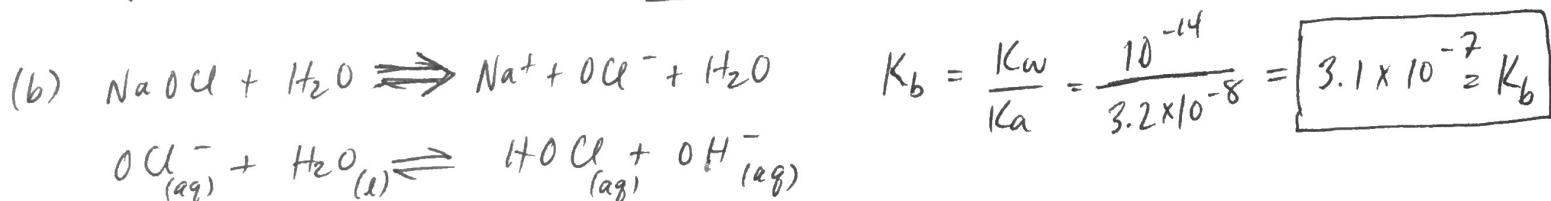


Calculate the pH of such a solution if the concentration of HOCl in the solution is 0.065 molar.

(a)	R	$\text{HOCl} \rightleftharpoons \text{OCl}^- + \text{H}^+$	$K_a = 3.2 \times 10^{-8} = \frac{x^2}{0.14-x}$
	I	0.14 M	0
C	-x	+x	+x
E	0.14-x	x	x

$x = [\text{H}^+] = 6.7 \times 10^{-5} \text{ M}$

$\text{pH} = -\log [\text{H}^+] = 4.2$



(c)  $(0.0400 \text{ L})(0.14 \text{ M}) = 0.00560 \text{ mol HOCl} \therefore (0.0100 \text{ L})(0.56 \text{ M}) = 0.00560 \text{ mol NaOH}$   
 $\therefore @ \text{ equivalence point!} \text{ and } \therefore \text{produced } 0.00560 \text{ mol OCl}^-$

R	$\text{OCl}^- + \text{H}_2\text{O} \rightleftharpoons \text{HOCl} + \text{OH}^-$	$K_b = \frac{x^2}{0.112-x}$
I	$\frac{0.00560 \text{ mol}}{0.0500 \text{ L}}$	0
C	-x	+x
E	$0.112-x$	x

$x = [\text{OH}^-] = 1.86 \times 10^{-4} \text{ M}$

$\text{pOH} = -\log [\text{OH}^-] = 3.73 \therefore \boxed{\text{pH} = 14 - 3.73 = 10.3}$  (10.27)

$$(d) (0.050 \text{ M})(0.20 \text{ M}) = 0.0100 \text{ mol HCl} = 10.0 \text{ mmol HCl}$$

$$\text{pH} = \text{pK}_a + \log \frac{[\text{base}]}{[\text{acid}]}$$

$$\text{pK}_a = -\log K_a = -\log(3.2 \times 10^{-8})$$

$$\text{pH} = 7.49 \quad \longleftrightarrow \quad \text{pK}_a = 7.49$$

this tells us we are at the  $\frac{1}{2}$  equivalence point and  $\therefore$  mols HA ~~has been reduced~~ ~~reduced by  $\frac{1}{2}$~~  of the original amount (or is  $\frac{1}{2}$  neutralized)

initial mmol HCl = 10.0 mmol  $\therefore$  5.0 mmol has reacted to reach  $\frac{1}{2}$  equivalence so we needed to add 5.0 mmol NaOH



for every mole of HOCl produced there is also 1 mol  $\text{H}^+$  produced  $\therefore [\text{HOCl}] = [\text{H}^+] = 0.065 \text{ M}$

$$\text{pH} = -\log [\text{H}^+] = -\log (0.065 \text{ M})$$

$$\boxed{\text{pH} = 1.19}$$

2)  $K_{sp}$

20 minutes

a) The solubility of  $\text{Cu}(\text{OH})_2$  is  $1.72 \times 10^{-6}$  gram per 100. milliliters of solution at  $25^\circ\text{C}$ .

(i) Write the balanced chemical equation for the dissociation of  $\text{Cu}(\text{OH})_2(s)$  in aqueous solution.

(ii) Calculate the solubility (in moles per liter) of  $\text{Cu}(\text{OH})_2$  at  $25^\circ\text{C}$ .

(iii) Calculate the value of the solubility-product constant,  $K_{sp}$ , for  $\text{Cu}(\text{OH})_2$  at  $25^\circ\text{C}$ .

b) The value of the solubility-product constant,  $K_{sp}$ , for  $\text{Zn}(\text{OH})_2$  is  $7.7 \times 10^{-17}$  at  $25^\circ\text{C}$ .

(i) Calculate the solubility (in moles per liter) of  $\text{Zn}(\text{OH})_2$  at  $25^\circ\text{C}$  in a solution with a pH of 9.35.

(ii) At  $25^\circ\text{C}$ , 50.0 milliliters of 0.100-molar  $\text{Zn}(\text{NO}_3)_2$  is mixed with 50.0 milliliters of 0.300-molar  $\text{NaOH}$ . Calculate the molar concentration of  $\text{Zn}^{2+}(\text{aq})$  in the resulting solution once equilibrium has been established. Assume that volumes are additive.



$$(\text{ii}) \left( \frac{1.72 \times 10^{-6}}{100 \text{ mL}} \text{ g} \right) \left( \frac{1000 \text{ mL}}{1 \text{ L}} \right) \left( \frac{1 \text{ mol}}{97.566 \text{ g}} \right) = \boxed{1.76 \times 10^{-7} \text{ M}}$$

$$(\text{iii}) K_{sp} = [\text{Cu}^{2+}][\text{OH}^-]^2 = x(2x)^2 = 4x^3 = (1.76 \times 10^{-7})(2 \cdot 1.76 \times 10^{-7})^2$$
$$\boxed{K_{sp} = 2.18 \times 10^{-20}}$$

	$\text{Zn}(\text{OH})_2(s)$	$\text{Zn}^{2+} + 2\text{OH}^-$
I	X	
C		$+ x$
E		$x$

$$\text{pH} = 9.35 \rightarrow 10^{-9.35}$$

$$\text{pOH} = 4.65 \rightarrow 10^{-4.65} = 2.24 \times 10^{-5} \text{ M} = [\text{OH}^-]$$

$$K_{sp} = [\text{Zn}^{2+}][\text{OH}^-]^2 = x(2.24 \times 10^{-5} + 2x)^2 = 7.7 \times 10^{-17}$$

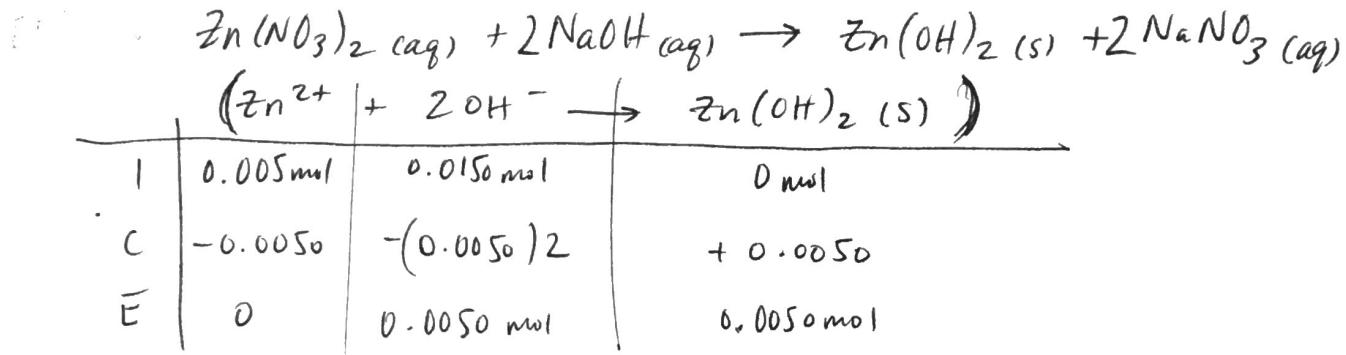
$$x = [\text{Zn}^{2+}] = \text{molar solubility} = \boxed{1.49 \times 10^{-7} \text{ M}}$$

$$\text{or } x(2.24 \times 10^{-5})^2 = 7.7 \times 10^{-17}$$

$$\boxed{x = 1.53 \times 10^{-7} \text{ M}} \quad \leftarrow \text{by assuming } 2x \text{ is a negligible amount}$$

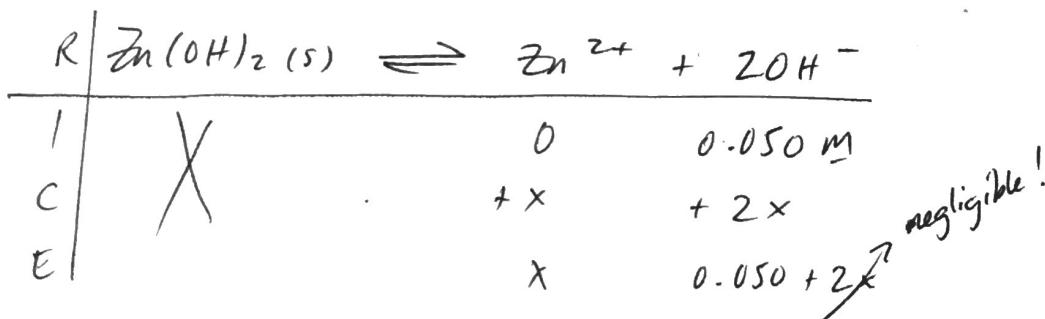
$$(b) \text{ (ii)} \quad (0.050 \text{ L})(0.100 \text{ M}) = 0.0050 \text{ mol } \text{Zn}(\text{NO}_3)_2 \text{ (aq)}$$

$$(0.050 \text{ L})(0.300 \text{ M}) = 0.0150 \text{ mol } \text{NaOH} \text{ (aq)}$$



$$V_{\text{total}} = (0.050 \text{ L}) + (0.050 \text{ L}) = 0.10 \text{ L}$$

$$\therefore [\text{OH}^-] = \frac{0.0050 \text{ mol}}{0.10 \text{ L}} = 0.050 \text{ M OH}^-$$



$$K_{\text{sp}} = [\text{Zn}^{2+}][\text{OH}^-]^2 = x(0.050)^2 = 7.7 \times 10^{-17}$$

$$x = [\text{Zn}^{2+}] = 3.1 \times 10^{-14} \text{ M}$$

3) Buffer  
20 Minutes

A buffer solution contains 0.40 mole of formic acid, HCOOH, and 0.60 mole of sodium formate, HCOONa, in 1.00 liter of solution. The ionization constant,  $K_a$ , of formic acid is  $1.8 \times 10^{-4}$

$$\boxed{pH = 3.92}$$

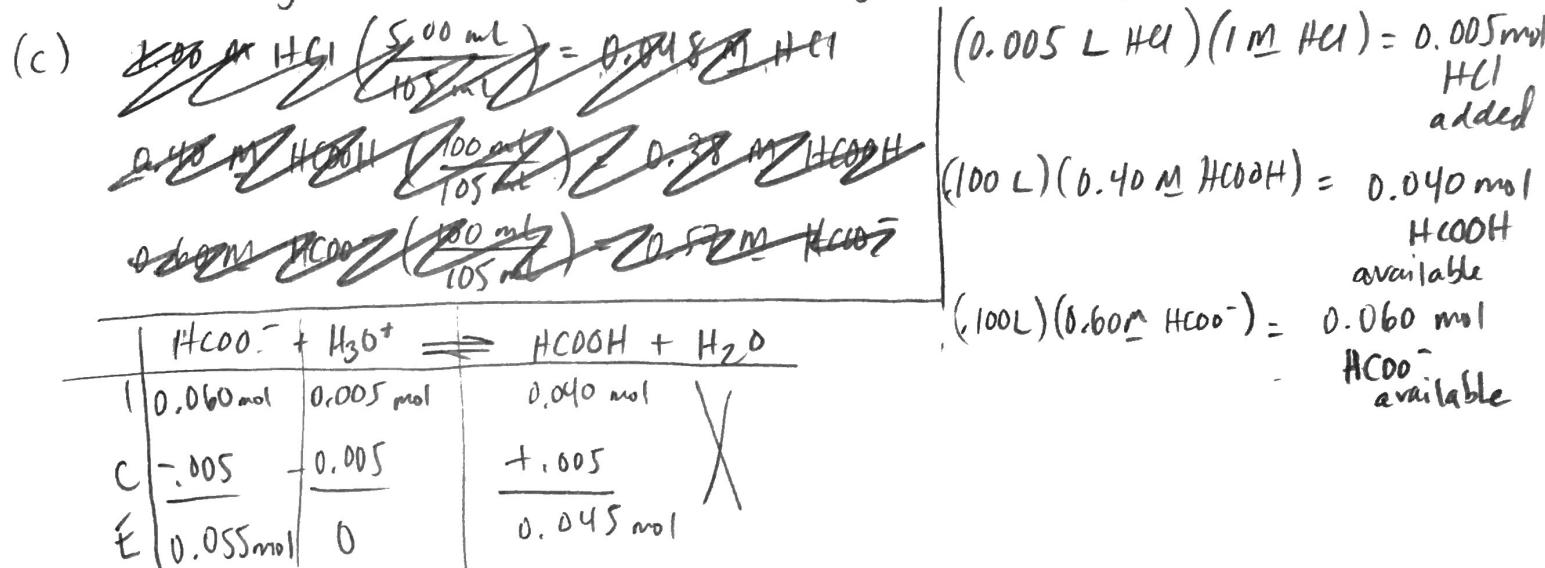
(a) Calculate the pH of this solution.  $pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log(1.8 \times 10^{-4}) + \log \left(\frac{0.60}{0.40}\right)$

(b) If 100. milliliters of this buffer solution is diluted to a volume of 1.00 liter with pure water, the pH does not change. Discuss why the pH remains constant on dilution.

(c) A 5.00-milliliter sample of 1.00-molar HCl is added to 100. milliliters of the original buffer solution. Calculate the  $[\text{H}_3\text{O}^+]$  of the resulting solution.

(d) A 800-milliliter sample of 2.00-molar formic acid is mixed with 200. milliliters of 4.80-molar NaOH. Calculate the  $[\text{H}_3\text{O}^+]$  of the resulting solution.

(b) no change in mole ratio by diluting  $\therefore$  no change in pH



$$pH = pK_a + \log \left( \frac{[\text{base}]}{[\text{acid}]} \right) = -\log(1.8 \times 10^{-4}) + \log \left( \frac{0.055 \text{ mol}}{0.045 \text{ mol}} \right)$$

$$pH = 3.83$$

$$[\text{H}_3\text{O}^+] = 10^{-3.83} = \boxed{1.5 \times 10^{-4} \text{ M}}$$



or

$$pH = pK_a + \log \frac{[\text{base}]}{[\text{acid}]} = -\log(1.8 \times 10^{-4}) + \log \left( \frac{0.52 \text{ M}}{0.48 \text{ M}} \right) = 3.83$$

$$(d) (0.800 \text{ L}) (2.00 \text{ M}) = 1.6 \text{ mol HCOOH}$$

$$(0.200 \text{ L}) (4.80 \text{ M}) = 0.96 \text{ mol NaOH}$$

0.64 mol HCOOH remaining

$$\frac{0.64 \text{ mol HCOOH}}{1.00 \text{ L}} = 0.64 \text{ M HCOOH}$$

R	HCOOH	$\rightleftharpoons$	$\text{COO}^-$	$\text{H}^+$
I	0.64 M		0	0
C	-x		+x	+x
E	$0.64 - x$		x	x

$$K_a = 1.8 \times 10^{-4} = \frac{x^2}{0.64 - x}$$

$$x = [\text{H}_3\text{O}^+] = 1.1 \times 10^{-2} \text{ M}$$

4) 10 minutes

$\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$  Ammonia is a weak base that dissociates in water as shown above. At 25°C, the base dissociation constant,  $K_b$ , for  $\text{NH}_3$  is  $1.8 \times 10^{-5}$ .

- Determine the hydroxide ion concentration and the percentage dissociation of a 0.150 molar solution of ammonia at 25°C.
- Determine the pH of a solution prepared by adding 0.0500 mole of solid ammonium chloride to 100. millilitres of a 0.150 molar solution of ammonia.
- Determine the pH of a 0.234 M ammonium chloride solution

