

**10. Refer to Section 14.1.**

Calculate  $[H^+]$ , and from that, pH and  $[OH^-]$ .

a.  $[H^+] = K_a \times \frac{[HB]}{[B^-]} = 6.2 \times 10^{-8} \times \frac{0.335}{0.335} = 6.2 \times 10^{-8} M$

$pH = -\log[H^+] = -\log(6.2 \times 10^{-8}) = 7.21$

$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{6.2 \times 10^{-8}} = 1.6 \times 10^{-7} M$

b.  $[H^+] = K_a \times \frac{[HB]}{[B^-]} = 6.2 \times 10^{-8} \times \frac{0.335}{0.100} = 2.1 \times 10^{-7} M$

$pH = -\log[H^+] = -\log(2.1 \times 10^{-7}) = 6.68$

$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{2.1 \times 10^{-7}} = 4.8 \times 10^{-8} M$

c.  $[H^+] = K_a \times \frac{[HB]}{[B^-]} = 6.2 \times 10^{-8} \times \frac{0.335}{0.0750} = 2.8 \times 10^{-7} M$

$pH = -\log[H^+] = -\log(2.8 \times 10^{-7}) = 6.55$

$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{2.8 \times 10^{-7}} = 3.6 \times 10^{-8} M$

d.  $[H^+] = K_a \times \frac{[HB]}{[B^-]} = 6.2 \times 10^{-8} \times \frac{0.335}{0.0300} = 6.9 \times 10^{-7} M$

$pH = -\log[H^+] = -\log(6.9 \times 10^{-7}) = 6.16$

$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{6.9 \times 10^{-7}} = 1.4 \times 10^{-8} M$

**14. Refer to Section 14.1.**

$$[\text{HF}] = 0.0237 \text{ M}$$

$$[\text{KF}] = \frac{0.037 \text{ mol.}}{0.135 \text{ L}} = 0.27 \text{ M}$$

$$[\text{H}^+] = K_a \times \frac{[\text{HB}]}{[\text{B}^-]} = 6.9 \times 10^{-4} \times \frac{0.0237}{0.27} = 6.0 \times 10^{-5} \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = -\log(6.0 \times 10^{-5}) = 4.22$$

**18. Refer to Section 14.1, Example 14.2 and Table 13.2.**

Calculate the moles of the acid and base, then calculate the  $[\text{H}^+]$  and pH.

- a. Calculate  $[\text{H}^+]$  from pH, then calculate the acid/base ratio of the buffer.

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-3.0} = 1 \times 10^{-3} \text{ M}$$

$$[\text{H}^+] = K_a \times \frac{[\text{HB}]}{[\text{B}^-]} \Rightarrow 1 \times 10^{-3} = 1.9 \times 10^{-4} \times \frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]}$$

$$\frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5$$

- b.  $\text{NaCHO}_2$  will dissociate to  $\text{Na}^+$  and  $\text{CHO}_2^-$ . Use this fact and the equation from part (a) to calculate  $[\text{HCHO}_2]$ .

$$\frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5 \Rightarrow \frac{[\text{HCHO}_2]}{0.139} = 5$$

$[\text{HCHO}_2] = 0.7 \text{ M}$ , thus 0.7 mol. would need to be added to the liter of solution.

$$c. \frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5 \Rightarrow \frac{0.159}{[\text{CHO}_2^-]} = 5$$

$$[\text{CHO}_2^-] = 0.03 \text{ M}$$

$$0.3500 \text{ L} \times \frac{0.03 \text{ mol. CHO}_2^-}{1 \text{ L}} \times \frac{1 \text{ mol. NaCHO}_2}{1 \text{ mol. CHO}_2^-} \times \frac{68.01 \text{ g NaCHO}_2}{1 \text{ mol. NaCHO}_2} = 0.7 \text{ g NaCHO}_2$$

$$d. \frac{[\text{HCHO}_2]}{[\text{CHO}_2^-]} = 5 \Rightarrow \frac{[\text{HCHO}_2]}{0.500} = 5$$

$[\text{HCHO}_2] = 2 \text{ M}$ , since the volume of the formate solution is one liter, 2 mol of  $\text{HCHO}_2$  is needed.

$$2 \text{ mol. HCHO}_2 \times \frac{1 \text{ L}}{0.236 \text{ mol. HCHO}_2} = 1 \times 10^1 \text{ L}$$

**20. Refer to Section 14.1.**

$$a. 5.50 \text{ g} \times \frac{1 \text{ mol. NH}_4\text{Cl}}{53.49 \text{ g NH}_4\text{Cl}} = 0.103 \text{ mol. NH}_4\text{Cl}$$

$$[\text{H}^+] = K_a \times \frac{n_{\text{NH}_4^+}}{n_{\text{NH}_3}} = 5.6 \times 10^{-10} \times \frac{0.103 \text{ mol. NH}_4^+}{0.0188 \text{ mol. NH}_3} = 3.1 \times 10^{-9} \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = -\log(3.1 \times 10^{-9}) = 8.51$$

- b. Note that volume was not used in the calculations above. Thus we would expect the pH to remain the same, independent of the volume of solution.

**38. Refer to Section 14.1.**

a.  $[H^+] = 10^{-\text{pH}} = 10^{-7.40} = 4.0 \times 10^{-8} M$

$$[H^+] = K_a \times \frac{[H_2PO_4^-]}{[HPO_4^{2-}]} \Rightarrow 4.0 \times 10^{-8} = 6.2 \times 10^{-8} \times \frac{[H_2PO_4^-]}{[HPO_4^{2-}]}$$

$$\frac{[H_2PO_4^-]}{[HPO_4^{2-}]} = 0.65$$

b.  $[H^+] = 10^{-\text{pH}} = 10^{-6.80} = 1.6 \times 10^{-7} M$

$$[H^+] = K_a \times \frac{n_{H_2PO_4^-}}{n_{HPO_4^{2-}}} \Rightarrow 1.6 \times 10^{-7} = 6.2 \times 10^{-8} \times \frac{n_{H_2PO_4^-}}{n_{HPO_4^{2-}}}$$

$$\frac{n_{H_2PO_4^-}}{n_{HPO_4^{2-}}} = 2.6$$

To calculate the amount of  $HPO_4^{2-}$  ions that were converted, set  $n_{HPO_4^{2-}} = 1$  in part (a) above. Then  $n_{H_2PO_4^-} = 0.65$ . Then set  $x = \text{moles } HPO_4^{2-} \text{ converted}$ . Substituting gives:

$$2.6 = \frac{0.65 + x}{1.0 - x}$$

$x = 0.54$ , thus 54% is converted.

c.  $[H^+] = 10^{-\text{pH}} = 10^{-7.80} = 1.6 \times 10^{-8} M$

$$[H^+] = K_a \times \frac{n_{H_2PO_4^-}}{n_{HPO_4^{2-}}} \Rightarrow 1.6 \times 10^{-8} = 6.2 \times 10^{-8} \times \frac{n_{H_2PO_4^-}}{n_{HPO_4^{2-}}}$$

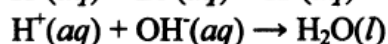
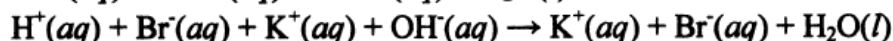
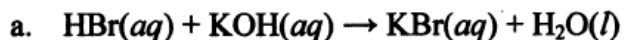
$$\frac{n_{H_2PO_4^-}}{n_{HPO_4^{2-}}} = 0.26$$

To calculate the amount of  $H_2PO_4^{2-}$  ions that were converted, set  $n_{H_2PO_4^-} = 1$  in part (a) above, then  $n_{HPO_4^{2-}} = 1.5$ . Then set  $x = \text{moles } HPO_4^{2-} \text{ converted}$ . Substituting gives:

$$2.6 = \frac{1 - x}{1.5 + x}$$

$x = 0.48$ , thus 48% is converted.

**44. Refer to Section 14.3 and Example 14.5.**



- b. At the equivalence point, all the base has reacted with the acid. Thus, the species present are  $\text{K}^+$ ,  $\text{Br}^-$ , and  $\text{H}_2\text{O}$ .

c. 
$$0.03500 \text{ L} \times \frac{0.257 \text{ mol. HBr}}{1 \text{ L}} \times \frac{1 \text{ mol. KOH}}{1 \text{ mol. HBr}} \times \frac{1 \text{ L}}{0.1375 \text{ mol. KOH}} = 0.0654 \text{ L}$$

- d. Before KOH is added, pH can be calculated from the HBr. Since HBr is a strong acid,

$$[\text{H}^+] = [\text{HBr}]$$

$$\text{pH} = -\log(0.257) = 0.590$$

- e. At half-way to the equivalence point, half of the HBr is neutralized. That means half as much base as needed to react the equivalence point will have been added. The total volume will be  $0.03500 \text{ L} + \frac{1}{2}(0.0654 \text{ L}) = 0.0677 \text{ L}$ .

$$0.03500 \text{ L} \times \frac{0.257 \text{ mol. HBr}}{1 \text{ L}} \times \frac{1}{2} = 0.00450 \text{ mol. HBr} = 0.00450 \text{ mol. H}^+$$

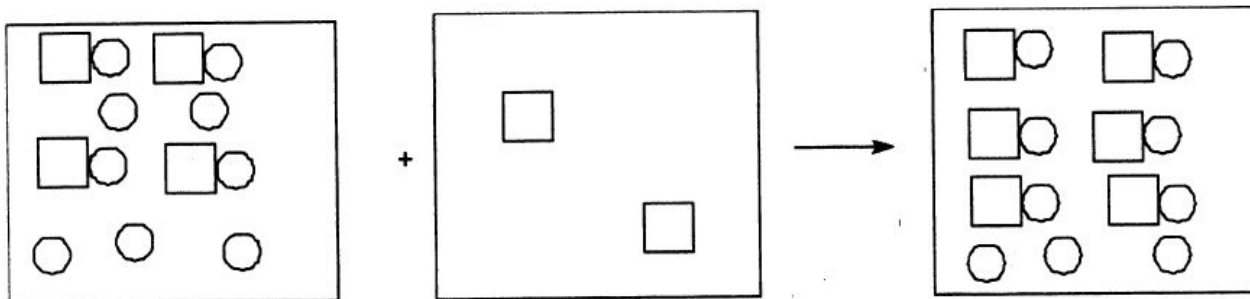
$$[\text{H}^+] = 0.00450 \text{ mol.} / 0.0677 \text{ L} = 0.0665 \text{ M}$$

$$\text{pH} = -\log [\text{H}^+] = -\log(0.0665) = 1.177$$

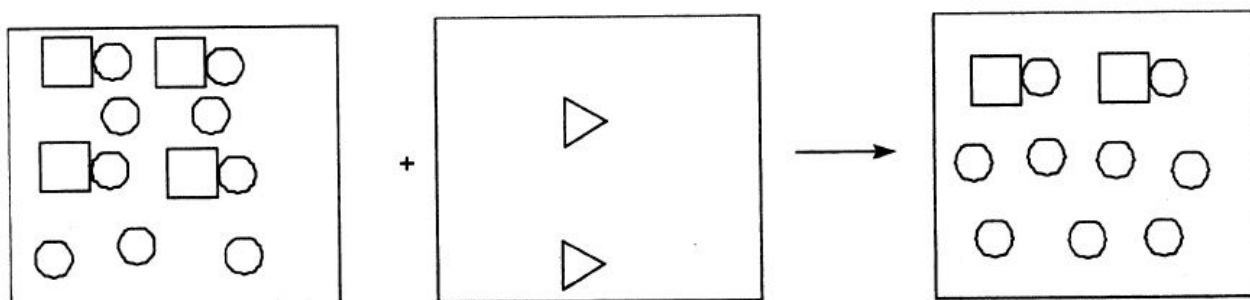
- f. At the equivalence point, the only species present are the spectator ions and water, thus the pH will be 7.

**54. Refer to Section 14.3.**

a. 2 mol. of  $H^+$  will react with 2 mol. of  $X^-$  to give 2 mol. of  $HX$ .



b. 2 mol.  $OH^-$  will react with 2 mol.  $HX$  forming 2 mol.  $X^-$  and 2 mol.  $H_2O$  (not shown).



c. 5 mol.  $OH^-$  will react with the 4 mol.  $HX$  forming 4 mol.  $X^-$  and 4 mol.  $H_2O$  (not shown).  
There will be one mol.  $OH^-$  left over, unreacted.

