(45)

c. $[OH^-] = \dfrac{1.0 \times 10^{-14}}{12} = 8.3 \times 10^{-16} M$; the solution is acidic.

d. $[OH^-] = \dfrac{1.0 \times 10^{-14}}{5.4 \times 10^{-5}} = 1.9 \times 10^{-10} M$; the solution is acidic.

45.) a) $[OH^-] = \dfrac{1.0 \times 10^{-14}}{10 \times 10^{-7}} = 10^{-7}$

Neutral

b) $[OH^-] = \dfrac{10^{-14}}{8.3 \times 10^{-16}} = 12 M$

Basic

(47.) a. Because the value of the equilibrium constant increases as the temperature increases, the reaction is endothermic. In endothermic reactions, heat is a reactant, so an increase in temperature (heat) shifts the reaction to produce more products and increases K in the process.

b. $H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$ $K_w = 5.47 \times 10^{-14} = [H^+][OH^-]$ at 50.°C
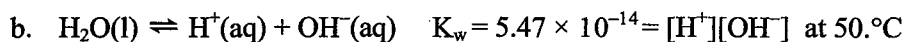
In pure water $[H^+] = [OH^-]$, so $5.47 \times 10^{-14} = [H^+]^2$, $[H^+] = 2.34 \times 10^{-7} M = [OH^-]$

(49.) $pH = -\log[H^+]$; $pOH = -\log[OH^-]$; at 25°C, $pH + pOH = 14.00$; for Exercise 45:

a. $pH = -\log[H^+] = -\log(1.0 \times 10^{-7}) = 7.00$; $pOH = 14.00 - pH = 14.00 - 7.00 = 7.00$

b. $pH = -\log(8.3 \times 10^{-16}) = 15.08$; $pOH = 14.00 - 15.08 = -1.08$

c. $pH = -\log(12) = -1.08$; $pOH = 14.00 - (-1.08) = 15.08$

d. $pH = -\log(5.4 \times 10^{-5}) = 4.27$; $pOH = 14.00 - 4.27 = 9.73$

Note that pH is less than zero when $[H^+]$ is greater than 1.0 M (an extremely acidic solution). For Exercise 46:

a. $pOH = -\log[OH^-] = -\log(1.5) = -0.18$; $pH = 14.00 - pOH = 14.00 - (-0.18) = 14.18$

b. $pOH = -\log(3.6 \times 10^{-15}) = 14.44$; $pH = 14.00 - 14.44 = -0.44$

c. $pOH = -\log(1.0 \times 10^{-7}) = 7.00$; $pH = 14.00 - 7.00 = 7.00$

d. $pOH = -\log(7.3 \times 10^{-4}) = 3.14$; $pH = 14.00 - 3.14 = 10.86$

Note that pH is greater than 14.00 when $[OH^-]$ is greater than 1.0 M (an extremely basic solution).

(51.) a. $pOH = 14.00 - 6.88 = 7.12$; $[H^+] = 10^{-6.88} = 1.3 \times 10^{-7} M$

$[OH^-] = 10^{-7.12} = 7.6 \times 10^{-8} M$; acidic

b. $[H^+] = \dfrac{1.0 \times 10^{-14}}{8.4 \times 10^{-14}} = 0.12 M$; $pH = -\log(0.12) = 0.92$

51

$pOH = 14.00 - 0.92 = 13.08$; acidic

c.　$pH = 14.00 - 3.11 = 10.89$; $[H^+] = 10^{-10.89} = 1.3 \times 10^{-11} M$

　　$[OH^-] = 10^{-3.11} = 7.8 \times 10^{-4} M$; basic

d.　$pH = -\log(1.0 \times 10^{-7}) = 7.00$; $pOH = 14.00 - 7.00 = 7.00$

　　$[OH^-] = 10^{-7.00} = 1.0 \times 10^{-7} M$; neutral

53.　$pOH = 14.0 - pH = 14.0 - 2.1 = 11.9$; $[H^+] = 10^{-pH} = 10^{-2.1} = 8 \times 10^{-3} M$ (1 sig. fig.)

$$[OH^-] = \frac{K_w}{[H^+]} = \frac{1.0 \times 10^{-14}}{8 \times 10^{-3}} = 1 \times 10^{-12} M \text{ or } [OH^-] = 10^{-pOH} = 10^{-11.9} = 1 \times 10^{-12} M$$

The sample of gastric juice is acidic because the pH is less than 7.00 at 25°C.

Solutions of Acids

55.　All the acids in this problem are strong acids that are always assumed to completely dissociate in water. The general dissociation reaction for a strong acid is $HA(aq) \rightarrow H^+(aq) + A^-(aq)$, where A^- is the conjugate base of the strong acid HA. For 0.250 M solutions of these strong acids,　0.250 M H^+ and 0.250 M A^- are present when the acids completely dissociate. The amount of H^+ donated from water will be insignificant in this problem since H_2O is a very weak acid.

a.　Major species present after dissociation $= H^+$, ClO_4^-, and H_2O;

　　$pH = -\log[H^+] = -\log(0.250) = 0.602$

b.　Major species $= H^+$, NO_3^-, and H_2O; $pH = 0.602$

57.　Strong acids are assumed to completely dissociate in water; for example; $HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$ or $HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$.

a.　A 0.10 M HCl solution gives 0.10 M H^+ and 0.10 M Cl^- because HCl completely dissociates. The amount of H^+ from H_2O will be insignificant.

　　$pH = -\log[H^+] = -\log(0.10) = 1.00$

b.　5.0 M H^+ is produced when 5.0 M $HClO_4$ completely dissociates. The amount of H^+ from H_2O will be insignificant. $pH = -\log(5.0) = -0.70$ (Negative pH values just indicate very concentrated acid solutions.)

c.　$1.0 \times 10^{-11} M$ H^+ is produced when $1.0 \times 10^{-11} M$ HI completely dissociates. If you take the negative log of 1.0×10^{-11}, this gives pH = 11.00. This is impossible! We dissolved an acid in water and got a basic pH. What we must consider in this problem is that water

46) a) $[H^+] = \dfrac{10^{-14}}{1.5 M} = 6.67 \times 10^{-15} M$ Basic

b) $[H^+] = \dfrac{10^{-14}}{3.6 \times 10^{-15} M} = 2.78 M$ Acidic

c) $[H^+] = \dfrac{10^{-14}}{1.0 \times 10^{-7} M} = 10^{-7} M$ Neutral

d) $[H^+] = \dfrac{10^{-14} M}{7.3 \times 10^{-4} M} = 1.37 \times 10^{-11} M$ Basic

48) $K_w = 2.92 \times 10^{-14} = [H^+][OH^-]$

a) $[H^+] = [OH^-] = \sqrt{2.92 \times 10^{-14}} = 1.71 \times 10^{-7} M$

b) $\underline{6.77} = -\log 1.71 \times 10^{-7} M = pH$

c $[OH^-] = 0.1 M$ $\dfrac{2.92 \times 10^{-14}}{(0.1 M)} = [H^+]$

$[H^+] = 2.92 \times 10^{-13} M$

$pH = -\log 2.92 \times 10^{-13}$

$\underline{pH = 12.5}$