Acid-Base Equilibria

Section 92 – pH and pOH

92-1 What are the pH and pOH of a solution of 2.0 M HCl? Recall that HCl is a strong acid, and ionizes completely in aqueous solution.

Solution

HCl is a strong acid that undergoes complete ionization in pure water.

$[H_3O^+]$ = molarity of HCl = 2.0 *M*

pH = –log$[H_3O^+]$ = –log(2.0) = –0.30

pOH = 14.00 – pH = 14.00 – (–0.30) = 14.30

92-2 Calculate the pH and the pOH of each of the following solutions at 25 °C. (Note that these substances are strong acids, and so are completely ionized in aqueous solution).

(a) 0.200 *M* HCl
(b) 0.350 *M* HNO_3

Solution

pH = –log$[H_3O^+]$, pOH = –log[OH^-];

(a) Hydrochloric acid is a strong acid; therefore, the hydronium ion concentration is the same as the molar concentration of HCl (since there is a 1:1 ratio in the stoichiometry of the formula).
pH = – log(0.200) = –(–0.699) = 0.699;
pOH = 14.00 – pH = 14.00 – 0.699 = 13.30;

(b) Nitric acid is a strong acid; therefore, the hydronium ion concentration is the same as the molar concentration of HNO_3 (since there is a 1:1 ratio in the stoichiometry of the formula).
pH = – log(0.350) = –(-0.455) = 0.455; pOH = 14.00 – pH = 14.00 – 0.455 = 13.54.

92-3 Calculate the pH and the pOH of each of the following solutions at 25 °C. (Note that these substances are strong bases, and so are completely ionized in aqueous solution.)

(a) 0.0143 *M* NaOH
(b) 0.0031 *M* $Ca(OH)_2$

Solution

pH = –log$[H_3O^+]$, pOH = –log[OH^-];

(b) NaOH is a strong base that is completely ionized in dilute solution. The stoichiometry for the number of hydroxide ions formed per NaOH is 1:1. Therefore, since [OH^-] = 0.0143 *M*,
pOH = –log(0.0143) = –(–1.8447) = 1.845;

pH = 14.00 – pOH = 14.00 – 1.845 = 12.16;

(b) Calcium hydroxide is also a strong base, but in this case the stoichiometry is such that two hydroxides are formed when one $Ca(OH)_2$ dissociates in water. Thus, $[OH^-]$ = 2(0.0031) = 0.0062 *M*;
pOH = –log(0.0062) = –(–2.208) = 2.21;
pH = 14.00 – 2.208 = 11.792 = 11.79

92-4 Calculate the pH and the pOH of each of the following aqueous solutions at 25 °C.

(a) 0.000259 *M* $HClO_4$
(b) 0.21 *M* NaOH
(c) 0.000071 *M* $Ba(OH)_2$
(d) 2.5 *M* KOH

Solution

(a) pH = –log(0.000259) = –(–3.5867) = 3.587;
pOH = 14.0000 – 3.5867 = 10.4133 = 10.413;
Solution is acidic.

(b) pOH = –log(0.21) = –(–0.678) = 0.68;
pH = 14.000 – 0.678 = 13.322 = 13.32;
Solutions is basic.

(c) Note the formula is for $Ba(OH)_{\mathbf{2}}$. Thus $[OH^-]$ = 2(0.000071) = 0.000142 *M*;
pOH = –log(0.000142) = –(–3.848) = 3.85;
pH = 14.000 – 3.848 = 10.152 = 10.15;
Solution is basic.

(d) pOH = –log(2.5) = –(0.398) = –0.40;
pH = 14.000 – (–0.398) = 14.398 = 14.4
Solution is basic.

92-5 What is the $[H^+]$ and $[OH^-]$ in each of the following at 25 °C? Are the solutions acidic, basic or neutral?

(a) An aqueous solution at pH 5.67
(b) An aqueous solution at pH 12.67
(c) An aqueous solution at pH = 4.99
(d) An aqueous solution at pH 8.62

Solution

Recall that pH = -log$[H^+]$.

(a) pH = -log$[H^+]$
5.67 = -log$[H^+]$
log $[H^+]$ = -5.67
$[H^+] = 10^{-5.67} = 2.13 \times 10^{-6}$ M

pOH = 14.00 – 5.67 = 8.33

$[OH^-] = 10^{-8.33} = 4.68 \times 10^{-9}$ M
Solution is acidic.

(b) $pH = -\log[H^+]$
$12.7 = -\log[H^+]$
$\log [H^+] = -12.7$
$[H^+] = 10^{-12.7} = 2.1 \times 10^{-13}$ M

$pOH = 14.00 – 12.7 = 1.3$
$[OH^-] = 10^{-1.3} = 0.050$ M
Solution is basic.

(c) $pH = -\log[H^+]$
$4.99 = -\log[H^+]$
$\log [H^+] = -4.99$
$[H^+] = 10^{-4.99} = 1.02 \times 10^{-5}$ M

$pOH = 14.00 – 4.99 = 9.01$
$[OH^-] = 10^{-9.01} = 9.77 \times 10^{-10}$ M
Solution is acidic.

(d) $pH = -\log[H^+]$
$8.62 = -\log[H^+]$
$\log [H^+] = -8.62$
$[H^+] = 10^{-8.62} = 2.40 \times 10^{-9}$ M

$pOH = 14.00 – 8.62 = 5.38$
$[OH^-] = 10^{-5.38} = 4.17 \times 10^{-6}$ M
Solution is basic.

92-6 Explain why a sample of pure water at 40 °C is neutral even though $[H_3O^+] = 1.7 \times 10^{-7}$ *M*. K_w is 2.910×10^{-14} at 40 °C.

Solution

In a neutral solution $[H_3O^+] = [OH^-]$. At 40 °C, $[H_3O^+] = [OH^-] = (2.9\ 10^{-14})^{1/2} = 1.7 \times 10^{-7}$.

92-7 What are the hydronium and hydroxide ion concentrations in a solution whose pH is 6.52?

Solution

$[H_3O^+] = 10^{-6.52} = 3.0 \times 10^{-7}$ *M*; pOH = 14.00 – pH; pOH = 14.00 – 6.52 = 7.48; $[OH^-] = 10^{-7.48} = 3.3 \times 10^{-8}$ *M*

92-8 The hydronium ion concentration in a sample of rainwater is found to be 1.7×10^{-6} *M* at 25 °C. What is the concentration of hydroxide ions in the rainwater?

Solution

$[H_3O^+]$ [OH⁻] = 1.0×10^{-14}; $[1.7 \times 10^{-6}]$[OH⁻] = 1.0×10^{-14};

$$[\text{OH}^-] = \frac{1.00 \times 10^{-1}}{1.7 \times 10^{-6}} = 5.9 \times 10^{-9}\ M$$

92-9 The hydroxide ion concentration in household ammonia is 3.2×10^{-3} M at 25 °C. What is the concentration of hydronium ions in the solution?

Solution

$[H_3O^+]$ [OH⁻] = 1.0×10^{-14}; $[H_3O^+]$ $[3.2 \times 10^{-3}]$ = 1.0×10^{-14}; $[\text{OH}^-] = \frac{1.00 \times 10^{-14}}{3.2 \times 10^{-3}} = 3.1 \times 10^{-12}\ M$

92-10 The ionization constant for water (K_w) is 2.9×10^{-14} at 40 °C. Calculate $[H_3O^+]$, [OH⁻], pH, and pOH for pure water at 40 °C.

Solution

$K_w = [H_3O^+][\text{OH}^-] = 2.9 \times 10^{-14}$

$$[H_3O^+] = [\text{OH}^-] = \sqrt{2.9 \times 10^{-14}} = 1.70 \times 10^{-7} = 1.7 \times 10^{-7}\ M$$

pH = pOH = $-\log(1.70 \times 10^{-7}) = 6.769 = 6.77$

92-11 The ionization constant for water (K_w) is 9.311×10^{-14} at 60 °C. Calculate $[H_3O^+]$, [OH⁻], pH, and pOH for pure water at 60 °C.

Solution

For water, $[H_3O^+]$ = [OH⁻] = x.

$K_w = 9.311 \times 10^{-14} = x^2$

$x = 3.051 \times 10^{-7} M = [H_3O^+] = [\text{OH}^-]$

pH = $-\log 3.051 \times 10^{-7} = -(-6.5156) = 6.5156$

pOH = pH = 6.5156