Unit P Acid Base (Proton) Reactions

Acid bases have to work in water – solution
Acids → donate H⁺
Base → donate OH⁻

P.2 Bronsted-Lowry Theory of Acids and Bases

An acid base reaction is a proton-transfer reaction in which the proton is transferred from the acid to the base. An acid is a proton donor, and a base is a proton acceptor (according to this theory, anything that can receive a proton is a base)
(base should be negatively charged)
- acid → proton donor, H⁺
- base → proton acceptor
- always have acid-base pairs; H – O – H + NH₃ ⇌ H – O⁻ + NH₄⁺
  (acid) (base) (base) . . (acid)
  (gave up a proton) (acceptor)
the stronger the acid, the weaker its conjugate base, see table P.1
ex: perchloric acid
HClO₄ ⇌ H⁺ + ClO₄⁻
equilibrium HClO₄ → H⁺ + ClO₄⁻
Equilibrium is going to favor the weak acid

P.8 Water Equilibrium

HOH ⇌ H⁺(aq) + OH⁻(aq)
kw = equilibrium constant for H₂O

kw = [H⁺][OH⁻] = [H⁺][OH⁻] = 1.0 x 10⁻¹⁴
[HOH]

Base [H⁺]<[OH⁻]

(1)

(7) Neutral [H⁺]=[OH⁻] under Neutral conditions; [H⁺]=[OH⁻]=1.0 x 10⁻⁷
  (water is neutral solution)
Acidic [H⁺]>[OH⁻]

Ex:  [H⁺]=1.0 x 10⁻³ > 1.0 x 10⁻⁷ = acidic, calculate [OH⁻]

\[ \text{kw} = [\text{H}^+][\text{OH}^-] \]
\[ [\text{OH}^-] = \frac{\text{kw}}{[\text{H}^+]} = \frac{1.0 \times 10^{14}}{1.0 \times 10^{-3}} = 1.0 \times 10^{11} = \text{acidic} \]

\([\text{H}^+]\) 1.0 x 10⁻³

**P.9  pH and pOH (integer values only)**

Rather than express very small [H⁺] and [OH⁻] values in negative exponentials, chemists use base-10 logarithms in the form of “p” numbers.

pH = - log[H⁺]  pOH = - log[OH⁻]

14.0 (basic) ----  7.0 (neutral) ----  0.0 (acidic)

ex: what is the [H⁺] of a pH=2.3
1. equation: pH = - log[H⁺]
2. take: - pH = log[H⁺]
3. inverse log; (-pH) = [H⁺]
4. inv log (-2.3) = 5.0 x 10⁻³ M (enter 2.3 [±] 2nd LOG = .005

ex: what is the[OH⁻]
1. equation; pH = - log[OH⁻] or 14 – 2.3 = 11.7
2. –pH = log[OH⁻]  pH = -log[OH⁻]  … inv log (11.7) = [OH⁻]

ex: what is the pH [H+] 4.2 x 10⁻¹
pH = - log[H⁺]

pH = - log [4.2 x 10⁻¹] = 2.4

ex: add 3.53g of HCl to water to form a 842ml solution, calculate resulting pH

pH = -log[H⁺]  HCl → H⁺ + Cl⁻

\[ \frac{3.5 \text{g HCl}}{1 \text{ mol HCl}} \times \frac{1 \text{ mol H⁺}}{1 \text{ mol HCl}} = 0.114 \text{ mol H⁺} \]

\[ \frac{36.45 \text{g HCl}}{1 \text{ mol HCL}} \times \frac{1 \text{ mol HCL}}{0.842 \text{ L}} \]

pH = -log[H⁺]  pH = -log[0.114] = pH = 0.943