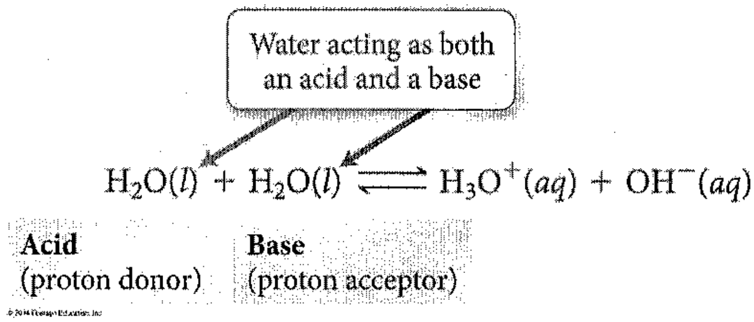
K_w: The Self-Ionization of Water

Self-Ionization of Water

- To understand whether a solution is acidic or basic, we use pure water as a neutral starting point.
- In the auto-ionization of water, a water molecule produces a hydronium ion, H_3O^+, and a hydroxide ion, OH^-.



A. The self-ionization of water produces equal (to) amounts of hydronium and hydroxide ions.

- Each ion has a concentration of 1×10^{-7} M. (in pure water at 25°C)
- Since each ion has equal concentration, we say that pure water is neutral.

B. The equilibrium constant for the self-ionization of water, K_w, is 1×10^{-14}.

$$K_w = \frac{[products]}{[reactants]} = [products] = [H^+][OH^-]$$ (reactants get crossed out because reactants are liquids)

$$K_w = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = \mathbf{1.0 \times 10^{-14}}$$ (at 25°C)

Neutral, Acidic, and Basic Solutions: A solution is then determined to be acidic or basic depending on which ion is in greater concentration.

Acidic: $[H^+]$ > $[OH^-]$ **Neutral**: $[H^+]$ ≈ $[OH^-]$ **Basic**: $[H^+]$ < $[OH^-]$

Why is K_w useful? If you know the concentration of H^+ in solution, $[H^+]$, you can calculate $[OH^-]$ (and vice versa).

Example 1: Given an aqueous solution of 2.7×10^{-4} M $Ba(OH)_2$:

a. Calculate $[OH^-]$: $2.7E{-4}\ M\ Ba(OH)_2 \times \frac{2\ OH^-}{1\ Ba(OH)_2} = \boxed{5.4 \times 10^{-4}\ M\ OH^-}$

b. Calculate $[H^+]$: $K_w = [H^+][OH^-] \Rightarrow [H^+] = \frac{K_w}{[OH^-]} = \frac{1E{-14}}{5.4E{-4}} = \boxed{1.9 \times 10^{-11}\ M\ H^+}$

c. Is the solution acidic, basic, or neutral? basic! ($[OH^-] > [H^+]$)

Example 2: An aqueous solution of HNO_3 has an $[OH^-]$ concentration of 5.1×10^{-11} M.

a. Calculate the $[H^+]$. $[H^+] = \frac{K_w}{[OH^-]} = \frac{1E{-14}}{5.1E{-11}} = \boxed{2.0 \times 10^{-4}\ M\ H^+}$

b. What is the molarity of HNO_3? $\boxed{2.0 \times 10^{-4}\ M\ HNO_3}$ Strong acid, so $[H^+] = [acid]$

c. Is the solution acidic, basic, or neutral? acidic! ($[H^+] > [OH^-]$)

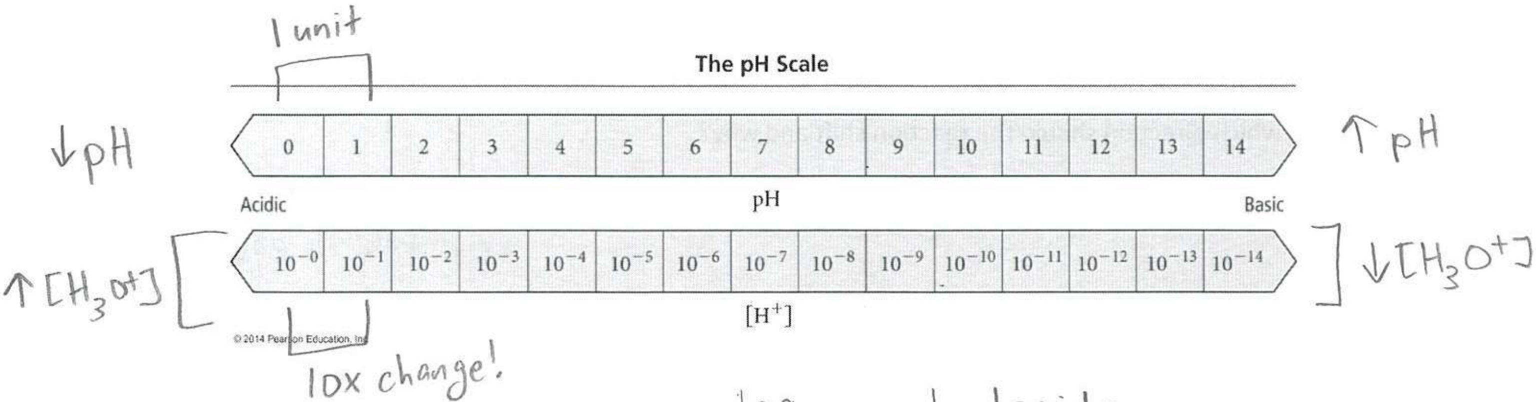
The pH Scale

pH and pOH scale: Another way to express if a solution is acid or basic is to use the pH scale.

1. The letters pH stand for the French words pouvoir hydrogène, meaning "hydrogen power"
2. The letter "p" is short for "-log" (just like "x" means multiply)
3. The pH of a solution is defined as the negative log of the hydronium ion concentration, $[H_3O^+]$.

$$pH = -\log[H^+] \quad \text{OR} \quad pH = -\log[H_3O^+]$$

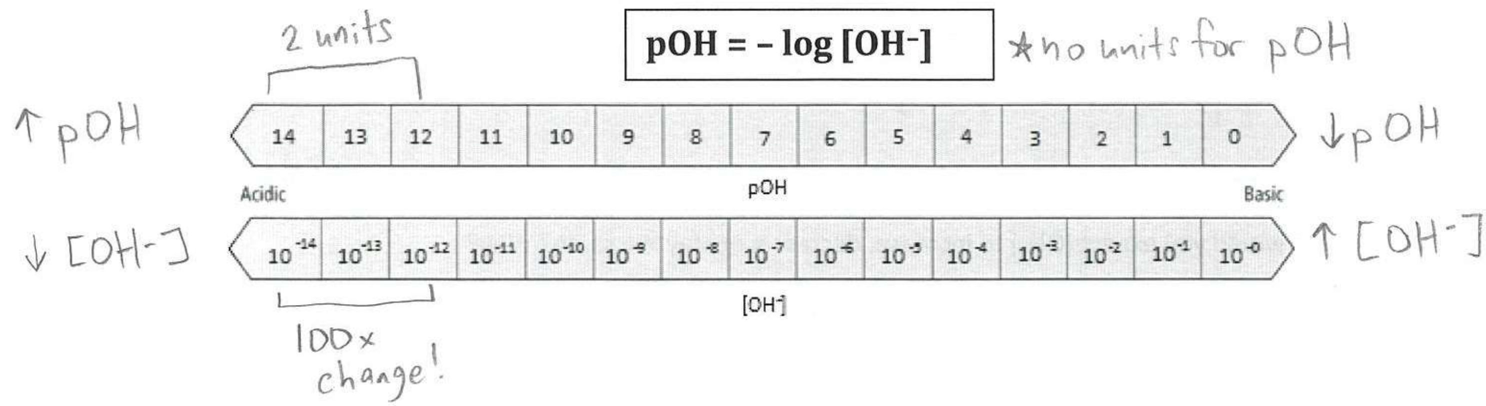
★no units for pH



4. The pOH of a solution is defined as the negative log of the hydroxide ion concentration, $[OH^-]$.

$$pOH = -\log[OH^-]$$

★no units for pOH



The best part:

$$pH + pOH = 14$$

Why? It all comes back to the expression for K_w.

$$K_w = [H_3O^+][OH^-] = (1.0 \times 10^{-7})(1.0 \times 10^{-7}) = 1.0 \times 10^{-14} \quad \text{(at 25°C)}$$

Take the log of both sides!

$$-\log[H_3O^+] + -\log[OH^-] = 14$$

$$pH + pOH = 14$$

Yay math!

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pH Calculations Concept Map

$[H^+]$ ⟵ $K_w = [H^+][OH^-] = 1E\text{-}14$ ⟶ $[OH^-]$

$[H^+]$ ↑ $[H^+] = 10^{\wedge}\text{-pH}$ | ↓ $pH = -\log[H^+]$ pH

$[OH^-]$ ↓ $pOH = -\log[OH^-]$ | ↑ $[OH^-] = 10^{\wedge}\text{-pOH}$ pOH

pH ⟵ $pH + pOH = 14$ ⟶ pOH

★ Strong acids/bases!
[strong acid] = $[H^+]$
[strong base] = $[OH^-]$
⇒ pH(SA) = −log[SA]

Significant Figures and Logarithmic Scales

Sig. figs. are the digits after **the decimal point in the log**

2 significant digits — 2 decimal places

$\log(1.0 \times 10^{-3}) = 3.00$

3 significant digits — 3 decimal places

$\log(1.00 \times 10^{-3}) = 3.000$

More Examples: $\log(2.25 \times 10^{6}) = 6.352$ $\log(0.03) = -1.5$ $10^{1.18} = 15$ $10^{1.9} = 80$

Now you try!

1. $\log(4 \times 10^{3})$ = 3.6 (1 sf)
2. $\log(6.6667 \times 10^{-8})$ = −7.17609 (5 sf)
3. $10^{0.8}$ = 6 (1 sf)
4. $10^{4.92}$ = 83,000 (2 sf)
5. Calculate $[H^+]$, $[OH^-]$, pH and pOH of a 0.25 M $Ca(OH)_2$ solution.

$[OH^-] = 0.25\,M\,Ca(OH)_2 \times \frac{2\,OH^-}{1\,Ca(OH)_2} = \boxed{0.50\,M\,OH^-}$

$[H^+] = \frac{K_w}{[OH^-]} = \frac{1E\text{-}14}{0.50} = \boxed{2.0 \times 10^{-14}\,M\,H^+}$

$p(OH) = -\log[OH^-] = -\log(0.50) = \boxed{0.30}$

$pH = -\log[H^+] = -\log(2.0E\text{-}14) = \boxed{13.70}$

6. The pH of a sample of human blood was measured to be 7.41 at 25°C. Calculate pOH, $[H^+]$ and $[OH^-]$ for the sample.

$pOH = 14 - pH = 14 - 7.41 = \boxed{6.59}$

$[OH^-] = 10^{\wedge}\text{-pOH} = 10^{-6.59} = \boxed{2.6 \times 10^{-7}\,M\,OH^-}$

$[H^+] = 10^{\wedge}\text{-pH} = 10^{-7.41} = \boxed{3.9 \times 10^{-8}\,M\,H^+}$