

# Chapter 15

## Outline

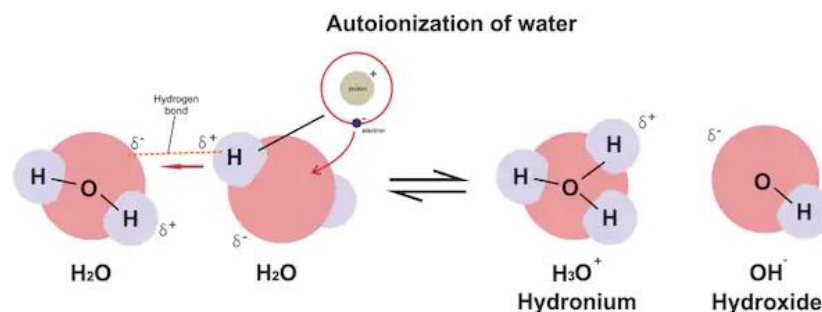
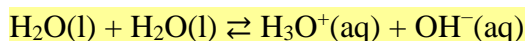
### Main Ideas

#### I. Aqueous Solutions and the Concept of pH

##### A. Hydronium Ions and Hydroxide Ions

##### 1. Self-Ionization of Water

- a. *Autoprotolysis*: proton transfer between identical species where one acts as a B-L acid, the other as a B-L base.



- b.  $\text{H}_3\text{O}^+$  ions &  $\text{OH}^-$  ions produced in 1:1 ratio during self-ionization of water  
 c. At  $25^\circ\text{C}$ ,

$$[\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol/liter}$$

- i. Brackets [ ] around a formula or number indicate concentration in mol/liter
- ii. At higher temperatures, water will ionize to a greater extent, but  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$
- iii.  $\text{H}^+$  is often used to represent  $\text{H}_3\text{O}^+$  when working with acids. They mean the same thing in acid-base chemistry. If you see  $[\text{H}^+]$ , it means the same as  $[\text{H}_3\text{O}^+]!!!$

##### 2. Ionization Constant for Water ( $K_w$ )

- a. Equilibrium constant for the ionization of pure water.

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = [1.0 \times 10^{-7}][1.0 \times 10^{-7}] = 1.0 \times 10^{-14}$$

- i.  $K_w$  is temperature dependent but considered a constant (infinite significance) at typical room temperatures.
- ii. Units are usually not shown for  $K_w$ , but would be  $\text{mol}^2/\text{liter}^2$  or  $M^2$

T ( $^\circ\text{C}$ )	$K_w$ ( $\text{mol}^2/\text{liter}^2$ )	pH
0	$0.114 \times 10^{-14}$	7.47
10	$0.293 \times 10^{-14}$	7.27
20	$0.681 \times 10^{-14}$	7.08
25	$1.008 \times 10^{-14}$	7.00
30	$1.471 \times 10^{-14}$	6.92
40	$2.916 \times 10^{-14}$	6.77
50	$5.476 \times 10^{-14}$	6.63
100	$51.3 \times 10^{-14}$	6.14

##### 3. Neutral, Acidic, and Basic Solutions

Acidic	Neutral	Basic
$[\text{H}_3\text{O}^+] > [\text{OH}^-]$	$[\text{H}_3\text{O}^+] = [\text{OH}^-]$	$[\text{H}_3\text{O}^+] < [\text{OH}^-]$

##### 4. Calculating $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$

- a. Because  $[\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{[\text{OH}^-]} \quad \& \quad [\text{OH}^-] = \frac{1.0 \times 10^{-14}}{[\text{H}_3\text{O}^+]}$$

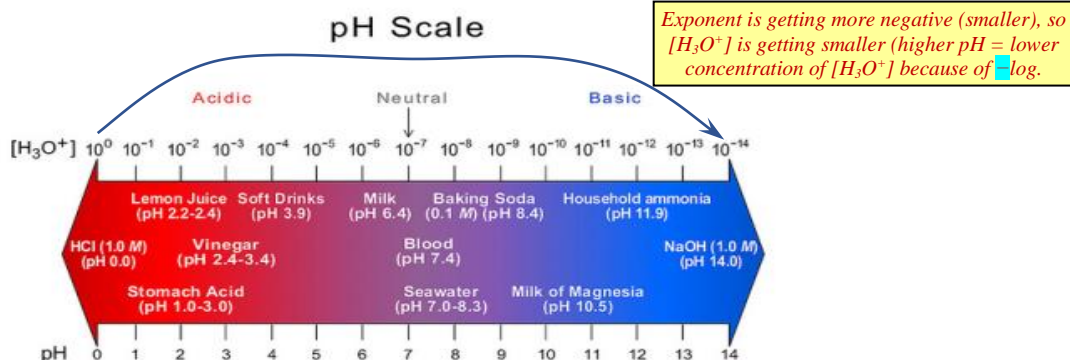
- b. Assume that strong acids and bases are completely ionized in solution:

- i.  $1.0M \text{Ba}(\text{OH})_2 = 2.0M \text{OH}^-$  ions
- ii.  $1.0M \text{H}_2\text{SO}_4 = 2.0M \text{H}_3\text{O}^+$  ions (At high concentrations,  $\text{H}_2\text{SO}_4$  would produce fewer  $\text{H}_3\text{O}^+$  ions in its second ionization step.)

# Chapter 15

## Outline

- B. pH & pOH  
1. The pH scale



Common logarithms (base 10) are very useful in representing very big numbers or very small numbers. In chemistry concentrations of hydronium ion ( $H_3O^+$ ) and hydroxide ion ( $OH^-$ ) are often represented as logarithms.

- Logarithms are basically exponents.

Examples:  $\log_{10}(10) = 1$  (because  $10^1 = 10$ ),  $\log_{10}(100) = 2$  (because  $10^2 = 100$ ),  $\log_{10}(1000) = 3$ , etc.

2. pH
- “p” is often used in science to mean the negative of the common (base 10) logarithm (i.e.,  $p(x) = -\log_{10}(x)$ )
  - pH is negative of the common logarithm of the hydrogen (really hydronium) ion concentration.

$$pH = -\log[H_3O^+]$$

- To find the  $[H_3O^+]$  from the pH, the inverse log function is used (sometimes called antilog or exponentiation). Basically, in the base 10 system it's:

$$[H_3O^+] = \text{antilog}_{10}(-pH) \quad \text{or} \quad [H_3O^+] = 10^{-pH}$$

- Use the second function (or shift +) of the “log” function on the calculator to access the antilog function. It should say “10<sup>x</sup>”.

3. pOH
- pOH is negative of the common logarithm of the hydroxide ion concentration.

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

- To find the  $[OH^-]$  from the pOH:

$$[OH^-] = \text{antilog}_{10}(-pOH) \quad \text{or} \quad [OH^-] = 10^{-pOH}$$

4. Logarithm math

- Due to the rules of logarithms (which I'm not sure you've really learned yet), the following is always true at temperatures near room temperature.

$$pH + pOH = 14.0$$

5. pH Calculations and the Strength of Acids and Bases

- Weak acids and weak bases cannot be assumed to be 100% ionized.
- $[H_3O^+]$  and  $[OH^-]$  cannot be determined from acid and base concentrations directly.
  - They be determined experimentally or calculated using equilibrium constants (*Chapter 18*)

# Chapter 15

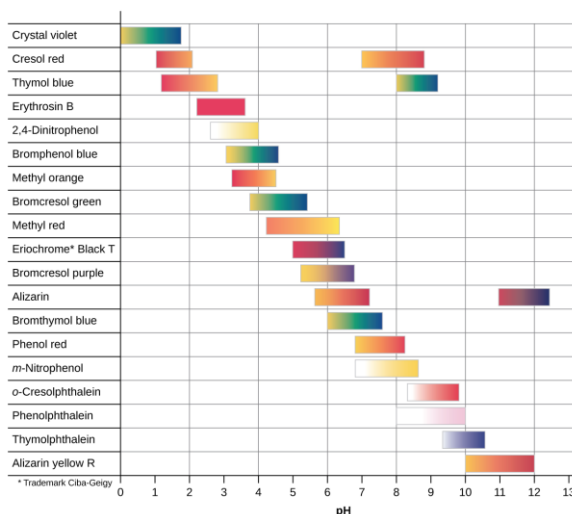
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### II. Determining pH and Titrations

#### A. Indicators and pH Meters

##### 1. Acid-Base Indicators

- Substances that change color depending on pH
- Are either weak acids or weak bases* whose equilibrium is disturbed.
- Many plants may be used as indicators
  - Red cabbage
  - Onion
  - Grape juice
  - Turmeric
- Transition Interval*
  - pH range over which an indicator color change occurs
- Choosing an Indicator
  - Indicators are useful when they change color in a pH range which includes the *equivalence point* of the reaction



##### 2. pH meter

- Measures voltage difference between two electrodes

#### B. Acid-Base Titration

##### 1. Titration

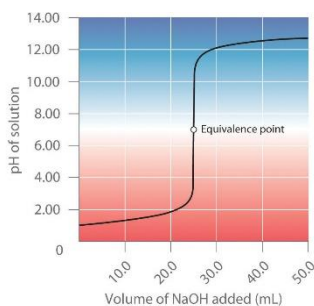
- The measured addition of a solution of a known concentration (the titrant) required to react completely with a measured amount of sol'n of unknown concentration (the analyte).

##### 2. Equivalence Point

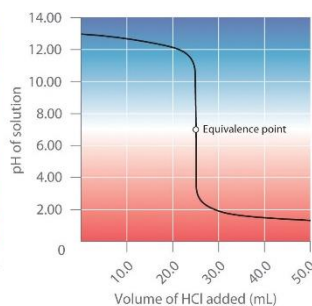
- The point at which the moles  $H^+$  from the acid = moles of  $OH^-$  from the base.

##### 3. Titration Curves

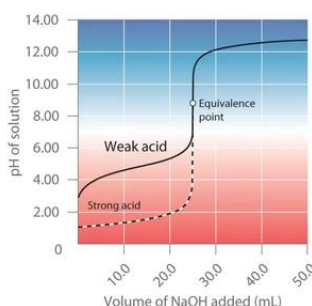
- End point* = the point in a titration when the indicator changes colors indicating the reaction is complete.
  - Not the same as equivalence point.*
    - Usually the *end point* is achieved after passing the *equivalence point*.
    - The *transition interval* of the indicator needs to overlap the pH range of the *equivalence point*.



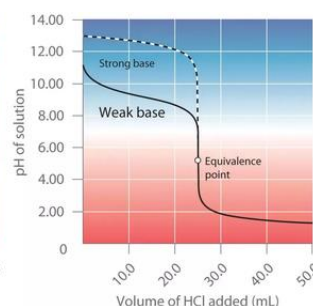
(a) Strong acid titrated with strong base



(b) Strong base titrated with strong acid



(a) Weak acid titrated with strong base



(b) Weak base titrated with strong acid

##### b. Equivalence point pH

- Strong acid – strong base titrations will reach their equivalence point at a pH of 7 (see first two graphs above)
- Strong base – weak acid titrations will have an equivalence point above 7.
  - As the weak acid is neutralized, it forms a conjugate base raising the pH of the solution.
- Weak base – strong acid titrations will have an equivalence point below 7.
  - As the weak base is neutralized, it forms a conjugate acid lowering the pH of the solution.
- Weak base – weak acid titrations are more complicated and the pH at the equivalence point will depend on the strength of the acid & base.

# Chapter 15

## Outline

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## Outline

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### C. Titration Calculations

#### 1. Standard Solution

- A solution that contains a precisely known concentration of solute, used in titration to find the concentration of the solution of unknown concentration (analyte).
  - Also referred to as the “known” solution.
- Primary Standard = a highly purified solid compound used to check the concentration of the known solution in a titration.

#### 2. Mole Calculations (Formal Method)

- Balance the equation.
  - From the balanced equation determine the mole ratio of the acid and base.
- Determine the number of moles of acid/base used in the titrant (known)
  - moles known = molarity of known  $\times$  volume of known (L)

$$\text{mol}_{(\text{known})} = M_{(\text{known})}V_{(\text{known})}$$

- Apply the mole ratio from the balanced equation to determine moles of unknown that were reacted.

$$\frac{\text{mol known}}{\text{coeff known}} \bigg| \frac{\text{coeff unknown}}{\text{coeff known}} = \text{mol unknown}$$

- Divide the moles of unknown by the volume of unknown used (in liters) to determine molarity of the unknown.

$$M_{(\text{unknown})} = \frac{\text{mol}_{(\text{unknown})}}{V_{(\text{unknown})}}$$

#### 3. Simple Titration Calculation

- If a neutralization reaction has a 1:1 mole ratio of acid to base, then:

$$\text{mol}_{(\text{known})} = M_{(\text{known})}V_{(\text{known})} \quad \text{and} \quad \text{mol}_{(\text{unknown})} = M_{(\text{unknown})}V_{(\text{unknown})}$$

since  $\text{mol}_{(\text{known})} = \text{mol}_{(\text{unknown})}$  in a 1:1 ratio,

$$M_{(\text{known})}V_{(\text{known})} = M_{(\text{unknown})}V_{(\text{unknown})}$$

or it's sometimes written  $M_A V_A = M_B V_B$

where A = acid and B = base

- Very important to remember that this only works in a 1:1 mole ratio.
- If the acid or bases produces multiple ions, then its concentration must be adjusted before using this formula. (this formula is typically shunned by the AP test.)