Unit 12: Acid and Bases

Chapter 19

- This tutorial is designed to help students understand scientific measurements.
- Objectives for this unit appear on the next slide.
  - Each objective is linked to its description.
  - Select the number at the front of the slide to go directly to its description.
- Throughout the tutorial, key words will be defined.
  - Select the word to see its definition.

Objectives

34 Identify properties of acids and bases

35 Define acids and bases by theories, include conjugate base pairs

36 Explain the difference between monoprotic, diprotic, and triprotic acids and relate to strength of acids

37 Define and calculate pH, pOH, and hydronium and hydroxide ion concentrations

38 Determine K_w, K_a, and K_b and use the values to predict strength

39 Define titration, neutralization, and indicators and perform and calculate titrations

34 Properties of Acids and Bases

- Several solutions have either acidic or basic properties.
- There is a general misconception that acids are all dangerous and bases are not.
  - This is not necessarily the case.
- Each category has distinct properties that will separate it from the other.

Properties

- Acids
  - Sour taste
  - 0-6.9 on pH scale
  - Turns litmus paper red
  - Common Examples
    - Pop
    - Citric acid
    - Battery acid
    - Vinegar
- Bases
  - Slippery
  - 7.1-14 on pH scale
  - Turns litmus paper blue
  - Common Examples
    - Lye
    - Soaps
    - Drano
    - Sodium hydroxide

Nomenclature of Acids

- Acid nomenclature depends on the anion.
  - If a single element, then:
    - Start with hydro-
    - Base of the anion (chlor for chlorine)
    - Finish with –ic acid

Hydro-_______-ic acid

HCl = hydrochloric acid

  - If a polyatomic ion, then:
    - Drop ending of polyatomic ion and…
      - If –ate, add –ic acid
      - If –ite, add –ous acid

$HClO_2$ = chlorous acid

$HClO_3$ = chloric acid

Nomenclature of bases

- Bases are named exactly the same as any ionic compound.
  - Remember Roman numerals for transition metals

  - $NaOH$ = sodium hydroxide
  - $Fe(OH)_2$ = iron (II) hydroxide

35 Acid-Base Theories

- Acids
  - Arrenhius
    - Substance that produces H_3O^+
  - Bronsted-Lowery
    - Proton donors
- Bases
  - Arrenhius
    - Substance that produces OH^-
  - Bronsted-Lowery
    - Proton acceptors

Conjugate Acids/Bases

- When an acid or base is added to water to make a solution, it will dissociate.
- For strong acids/bases, the dissociation is complete.
- For weak acids/bases, the dissociation is shown by an equilibrium reaction.
- The conjugate acid/base is formed when the dissociation occurs.

Conjugate Acids/Bases

- The dissociation of any acid will produce the hydronium ion.
- The dissociation of any base will produce the hydroxide ion.

- Example: dissociation of HCl

$$HCl + H_2O \rightarrow H_3O^+ + Cl^-$$

Conjugate Base

- Example: dissociation of NaOH

$$NaOH \rightarrow OH^- + Na^+$$

Conjugate Acid

36 Mono-, Di-, or Triprotic

- Acid strength is measured by the number of protons that the compound can donate.
- A strong acid/base will dissociate completely.
- A weak acid/base will have an incomplete dissociation and will make an equilibrium reaction.

Mono-, Di-, or Triprotic

- Mono-, di-, or triprotic are terms that are used to describe how many protons a compound can donate.
  - A monoprotic acid can donate 1 proton.
    - Example: HCl
  - A diprotic acid can donate 2 protons.
    - Example: H_2SO_4
  - A triprotic acid can donate 3 protons.
    - Example H_3P
- However, just because an acid can donate three protons does not necessarily mean that it will.

Mono-, Di-, or Triprotic

- Let's look the dissociation of H_2SO_4 (sulfuric acid)
  - $H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$
    - This is just part of the dissociation; one proton has been donated.
    - Note that the reaction is completely dissociated so sulfuric acid is a strong acid.
  - $HSO_4^- + H_2O \leftrightarrow H_3O^+ + SO_4^{-2}$
    - The second part of the dissociation will provide the second proton.
    - Note that this part of the dissociation is an equilibrium meaning that HSO_4^- is a weak acid.

37 pH Scale

- The concentration of hydronium ions and hydroxide ions are represented by the pH scale.
- pH stands for "power of hydrogen"
  - The scale is a logarithmic scale.
  - A pH value should only be reported to one decimal place.
  - The equation for pH is

    $$-\log[H_3O^+]$$
  - The equation for pOH is

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

pH Scale

- It is possible to take the pH and reverse the process to determine concentration of hydronium ions.
  - To do this, remember that a logarithmic scale is base ten.
  - Therefore, to work in reverse use:

$$[H_3O^+] = 10^{-pH}$$

Example: $[H_3O^+] = 0.0001\ M = 10^{-4}$

$pH = 4 = -\log [H_3O^+]$

38 K_a, K_b, K_w

- Because acids and bases dissociate completely, it is necessary to be able to calculate how much they will dissociate.
- This is determined by a constant for each known as K_a for acids and K_b for bases.
  - These are setup the same as K_{eq} equations.
- Before we get to the acids and bases though, we should consider the chemical that provides a few protons itself: water.

K_w

- The dissociation of water is rare but will occur.

$$H_2O \leftrightarrow H_3O^+ + OH^-$$

- The constant for this dissociation is known as K_w or the water dissociation constant.

- K_w is 1.0×10^{-14}

Constants

- The larger the constant for acids, bases, or water, the more likely it will be to dissociate.
  - Thus a larger K_a means more protons which means a stronger acid.
- Writing equations for these constants will be the same as in Unit 15 with one additional detail.
  - Only aqueous components are used in the equation (not liquids)

Writing K equations

- Consider water's dissociation again:

$$H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + OH^-_{(aq)}$$

- The Kw for this dissociation would look like:

$$K_w = [H_3O^+][OH^-]$$

- Note because water is a liquid, it does not appear in the equation.

Writing K equations

- The same principles that apply to water will also apply to acids and bases.
- For example:

$$HBr_{(aq)} + H_2O_{(l)} \leftrightarrow H_3O^+_{(aq)} + Br^-_{(aq)}$$

$$K_a = \frac{[H_3O^+][Br^-]}{[HBr]}$$

Using K equations

- Most commonly, these K equations are used to determine the pH of a substance.
- This is done by calculating the concentration of the hydronium ion.

$$K_a = \frac{[H_3O^+][Br^-]}{[HBr]}$$

- Once that concentration is known, use the pH equation: $-\log[H_3O^+]$

K_b

- It is simple to calculate the pH for acids using a K_a equation because the hydronium ion concentration can be solved for.
- For bases though, it is the hydroxide ion that is solved for:

$$Fe(OH)_{2(aq)} \leftrightarrow Fe^{+2}_{(aq)} + 2OH^{-}_{(aq)}$$

Thus: $K_b = \dfrac{[Fe^{+2}][OH^{-}]^2}{[Fe(OH)_2]}$

K_b

- Once [OH-] is known, it is possible to calculate pOH
  - $pOH = -\log[OH^-]$
- Since most bases are aqueous solutions, we can use the water dissociation to complete the problem.
  - $K_w=[H_3O^+][OH^-]$
  - Using this formula requires the use of small numbers. An easier equation can be determined it we take the logarithmic function of each part:
  - $pK_w = pH + pOH$
    - Since $K_w = 1.4 \times 10^{-14}$, then pK_w would equal 14.

K_b Recap

- Write the dissociation equation.
- Calculate $[OH^-]$ from the K_b equation.
- Determine pOH
- Use the water dissociation to determine pH:
  - 14 = pH + pOH

39 Neutralization

- When acids and bases react, they will create a neutralization reaction.
  - This is because the pH begins to return to 7 (neutral)
- Neutralization reactions are essentially double replacement reactions in which the products are always a salt and water.
  - A salt does not refer to NaCl but rather the product of an acid/base reaction.

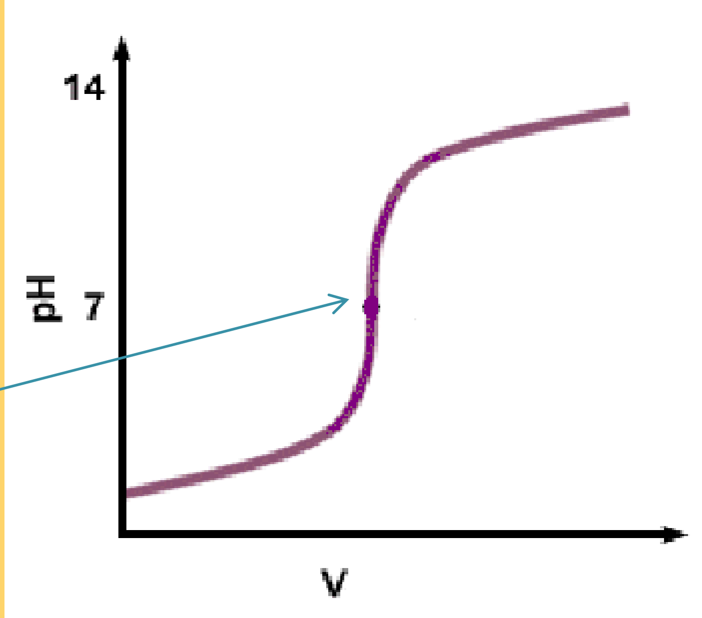
$$\text{Acid} + \text{Base} \rightarrow \text{Salt} + \text{Water}$$

Titrations

- A titration is type of experiment used to determine the concentration of an unknown.
- For acids and bases, this requires a known concentration of either the acid or the base.
- A titration looks at a titration curve to determine the equivalence point.

Equivalence Points

- The equivalence point occurs when the moles of acid equals the moles of the base.
- This will occur when the curve is the steepest.



Using a titration curve

- Once the equivalence point is found, the volume at that point can be used to calculate the concentration of the solution.
- The following equation can be used (same as the dilution equation):

$$M_aV_a=M_bV_b$$

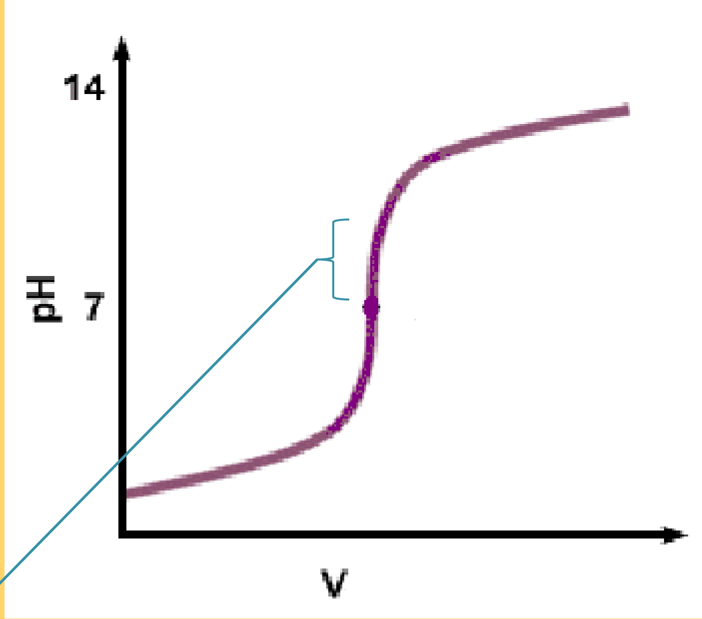
Where "a" stands for acid and "b" for base.

Indicators

- One challenge that does exist for titrations is locating the equivalence point.
- Most reactions between acids and bases occur when a clear acid solution is added to a clear base solution.
- The equivalence point can be found by measuring the pH or by using an indicator.
  - An indicator is a dye that will change color in a certain range of pH's.

Selecting an indicator

- To select an indicator, it is important to have a rough idea of where the equivalence point will occur.
- To insure you reach the equivalence point, you want to select an indicator that will change just after the equivalence point was reached.
- For the sample of the right, selecting an indicator for just above 7 would work best.
  - Phenolphthalien changes from clear to pink around a pH of 8.



- This concludes the tutorial on measurements.
- To try some practice problems, click here.
- To return to the objective page, click here.
- To exit the tutorial, hit escape.