40S Chemistry

Acid – Base Equilibrium

Topic Overview

- Electrolytes vs. Non-Electrolytes
- Defining Acids & Bases
- Strength of Acids & Bases
- Dissociation Constants
- pH
- Buffers
- Neutralization
- Titration *

* Lab scheduled

40S Chemistry

Electrolytes vs. Non-Electrolytes

Lesson Overview

- Electrolytes
- Dissociation
- Theory of Ionization
- Weak and Strong Electrolytes
- Dissociation Equations

Outcomes

- Define electrolyte and non-electrolyte.
- Identify strong and weak electrolytes.
- Distinguish between pure ionic and molecular compounds and their solutions.
- Using molecular diagrams, describe the dissolution of pure ionic and molecular compounds in water.
- Write balanced equations for the dissolution of pure ionic and molecular compounds in water.

Electrolytes

- Both acids and bases are electrolytes.

  Electrolytes are compounds that conduct an electric current in aqueous solution or in the molten state.

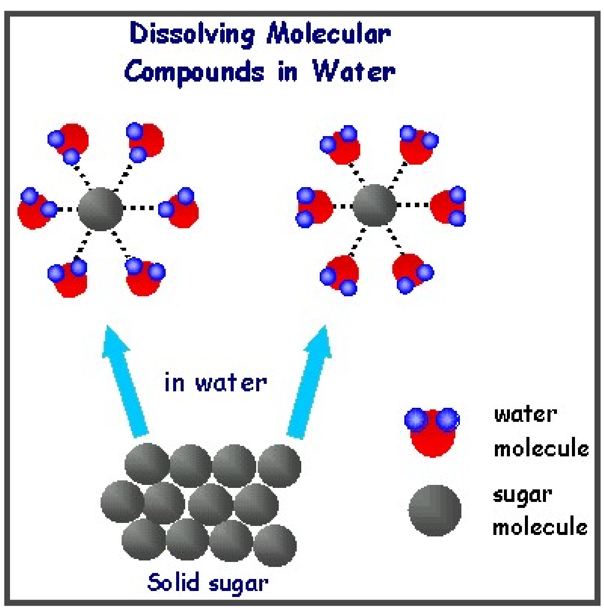
- Acids and bases conduct electricity. This is an indication that ions are present in acidic and basic solutions.

- All ionic compounds are electrolytes.
- Most covalent compounds are not.
- Strong electrolytes have high concentrations of ions while weak electrolytes have low ion concentrations.

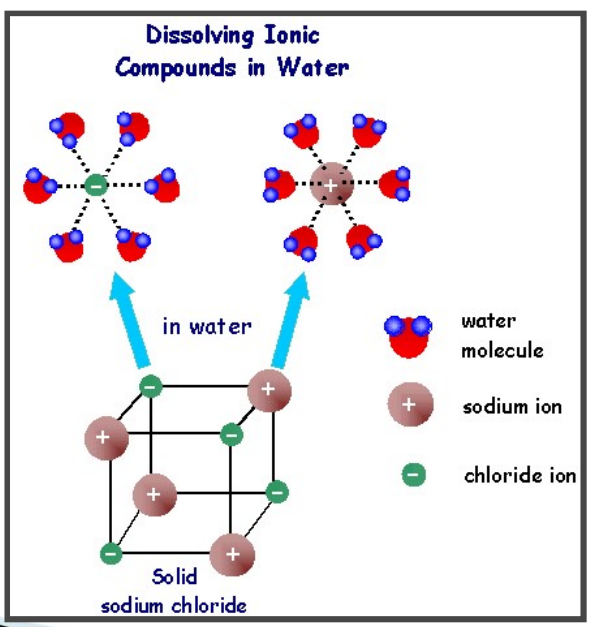
Dissociation

- Pure water is a non-electrolyte because it does not contain anything to carry an electric current.
- Charged particles or ions must be present in the solution for a solution to conduct an electric current
- The more ions present in a solution the greater the conductivity, up to a certain limit.

- The conductivity is due to the movement of the ions in solution.
- This explains the difference between the dissolving of sugar and sodium chloride in water.
- Sugar dissolves as whole molecules. These molecules are uncharged so there is nothing to carry the charge . . .



- Sodium chloride is an ionic compound.
- When it dissolves, it breaks up into ions, or dissociates.
- These ions are free to move and carry an electric current through the solution . . .



Theory of Ionization

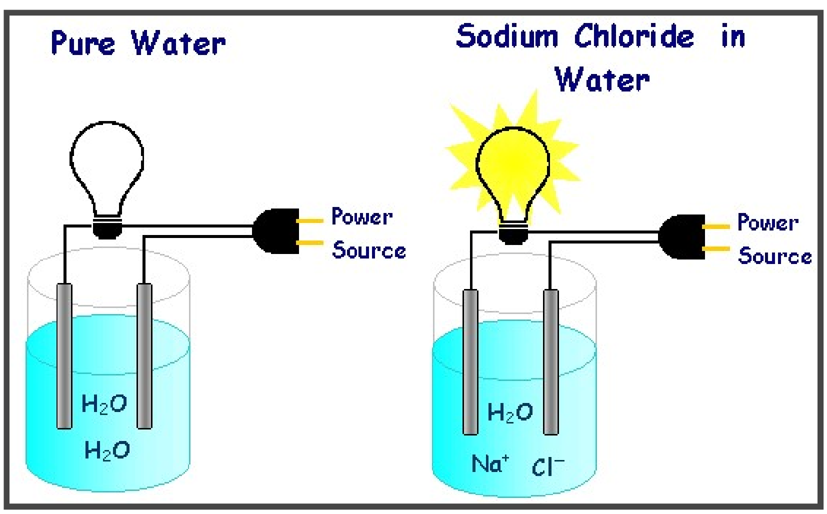
- In the early 1880's, a Swedish chemist named Svante Arrhenius (1859 – 1927) was studying the conductivity of solutions.
- Very little was known about the structure of atoms and molecules.
- He developed the idea that electrolytes break up into charged particles in water.

- The modern Theory of Electrolytes or Theory of Ionization includes the following principles:

1. Electrolytes in solution exist in the form of ions.
2. An ion is an atom or group of atoms which has an electric charge.
3. In electrolyte solutions, the total positive ionic charge must equal the total negative ionic charge.

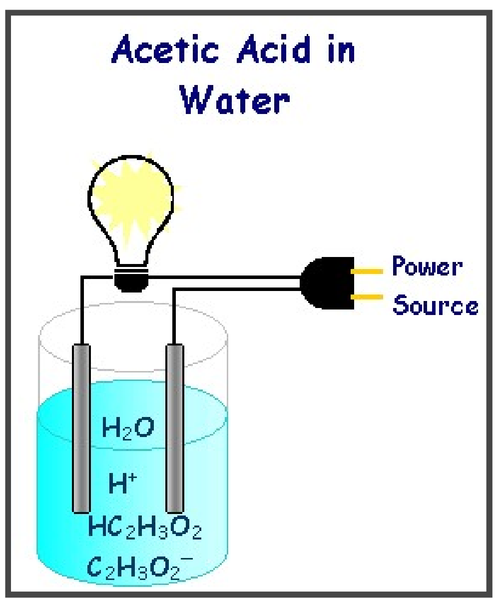
Weak and Strong Electrolytes

- Electrolytes are classified as strong and weak.
- A strong electrolyte is one which produces many ions in solution.
- Ionic compounds are strong electrolytes, since they will dissociate 100% in water. Ex: Sodium chloride



- A weak electrolyte is a substance which only partially dissociates.
- These are usually polar covalent molecules which partially ionize.
- Weakly conduct an electric current.

  Ex: acetic acid



Dissociation Equations

- The dissociation or ionization of a substance in water can be shown using chemical equations.
- Strong electrolytes, such as sodium chloride, are shown below:

  $NaCl(s) \rightarrow Na+(aq) + Cl^{-}(aq)$
- The (s) after the sodium chloride represents solid. The (aq) represents aqueous, which means dissolved in water.

- A non electrolyte such as table sugar or sucrose ($C_{11}H_{22}O_{11}$) dissociates according to the following equation:

  $C_{11}H_{22}O_{11}\ (s) \rightarrow C_{11}H_{22}O_{11}\ (aq)$
- The sugar does not separate into charged particles, but dissolves as whole molecules.

Assignment

- Complete the worksheet: *Electrolytes and Non-electrolytes*.

40S Chemistry

Defining Acids and Bases

Outcomes

- Identify substances that result in acidic or basic solutions when dissolved in water.
- Define and compare Arrhenius and Bronsted-Lowry acid/base theories.
- Describe the limitations of the Arrhenius theory for acids and bases.
- Write the formula for a conjugate acid-base pair.
- Write balanced chemical equations for reactions of acids and bases with water.

Lesson Overview

- Properties of Acidic Solutions
- Properties of Basic Solutions
- Arrhenius Definition of Acids & Bases
- Problems With the Arrhenius Definition
- Brønsted-Lowry Definition of Acids & Bases
- Conjugate Acids & Bases
- Hydrolysis
- Advantages of the Brønsted-Lowry Theory

Properties of Acidic Solutions

- taste sour
- conduct electricity
- have no characteristic feel
- turns blue litmus paper red
- react with metals to produce H_2
- react with carbonate compounds to produce CO_2

Properties of Bases

- taste bitter
- conduct electricity
- feel slippery
- turns red litmus paper blue
- do not react with metals
- do not react with carbonate compounds

Arrhenius Definition

- Acids produce hydrogen ions (H^+) in aqueous solution.
- Bases produce hydroxide ions (OH^-) when dissolved in water.
- Limits to aqueous solutions.
- Only one kind of base.
- NH_3 ammonia could not be an Arrhenius base.

Write the formulas for the following acids:

- √ acetic – $HC_2H_3O_{2\,(aq)}$
- √ carbonic – $H_2CO_{3\,(aq)}$
- √ hydrochloric – $HCl_{(aq)}$
- √ nitric – $HNO_{3\,(aq)}$
- √ nitrous – $HNO_{2\,(aq)}$
- √ sulfuric – $H_2SO_{4\,(aq)}$
- √ sulfurous – $H_2SO_{3\,(aq)}$
- √ phosphoric – $H_3PO_{4\,(aq)}$
- √ hypochlorous –$HClO_{(aq)}$

Write the formulas for the following bases:

√ ammonia	NH_3
√ sodium hydroxide	NaOH
√ magnesium hydroxide	$Mg(OH)_2$
√ calcium hydroxide	$Ca(OH)_2$
√ aluminum hydroxide	$Al(OH)_3$
√ potassium hydroxide	KOH

Monoprotic acids

- Any acid that contains one ionizable hydrogen.
- HNO_3, for example.
- Select and record the monoprotic acids from your list of acids.

Diprotic acids

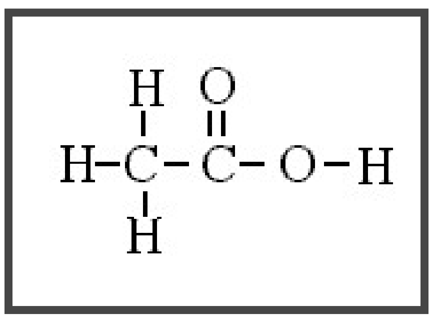
- Any acid that contains two ionizable hydrogens.
- H_2SO_4, for example.
- Select and record the diprotic acids from your list of acids.

Triprotic acids

- Any acid that contains three ionizable hydrogens.
- H_3PO_4, for example.
- Select and record the triprotic acids from your list of acids.

- Only the hydrogens in very polar bonds are ionizable. This occurs when H is joined to a very electronegative element.
- H+ ions are stabilized by solvation.

What about $HC_2H_3O_2$? Is it mono, di, or triprotic?



Acetic acid is monoprotic. Only the hydrogen bonded to the oxygen is ionnizable

Problems With the Arrhenius Definition

- There are many substances which are acidic or basic but do not have a hydrogen ion or a hydroxide ion.

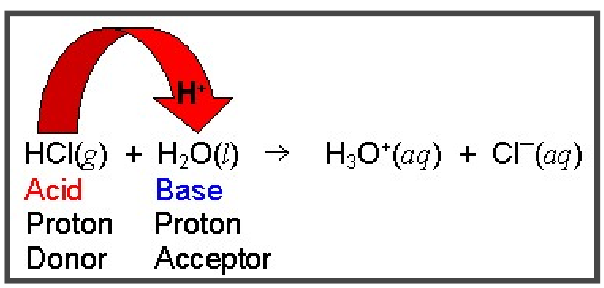
For example:

- baking soda (sodium bicarbonate, $NaHCO_3$) in water turns litmus blue, but has no apparent hydroxide ion.

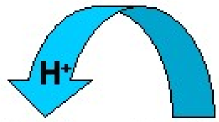
- metal ions, such as iron (III) and aluminum, turn litmus red, but have no hydrogen ions.
- The Arrhenius definition does not account for the acidity and basicity of these examples.

Bronsted-Lowry Definitions

- An acid is a proton (H^+) donor and a base is a proton acceptor.
- Acids and bases always come in pairs.
- HCl is an acid.
- When it dissolves in water it gives its proton to water.
- Water is a base and makes an hydronium ion



- Hydrogen chloride reacts with water by donating a proton.
- Water acts as the base, accepting the proton. The result is the H_3O+ ion called the hydronium ion.



$NH_3(g) + H_2O(l) \rightarrow NH_4^+(aq) + OH^-(aq)$

Base	Acid
Proton Acceptor	Proton Donor

- Ammonia's properties as a base are better explained by the Brønsted-Lowry definition.
- Ammonia accepts a proton from water, making ammonia a base and water the acid.
- The result is the ammonium ion and the hydroxide ion.

Bronsted-Lowry Acids

- Because of the donation and acceptance of H+ ions, Bronsted-Lowry acids and bases exist in conjugate acid-base pairs.
- Conjugate acid-base pairs are two substances that are related by the loss or gain of a single hydrogen ion.
- For example . . .

Consider the reaction:

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

acid base

In this reaction, HCl donates an H^+ to the H_2O. That makes HCl a Bronsted-Lowry acid (H^+ donor).

H_2O is a Bronsted-Lowry base (H^+ acceptor).

Consider the reaction:

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

acid base Conjugate acid

In this reaction, H_3O^+ is referred to as a conjugate acid.

A conjugate acid is the particle formed when a base gains a hydrogen ion.

Consider the reaction:

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

acid base Conjugate acid Conjugate base

In this reaction, Cl^- is referred to as a conjugate base.

A conjugate base is the particle that remains when an acid has donated a hydrogen ion.

Consider the reaction:

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

acid base Conjugate acid Conjugate base

- Each acid has a conjugate base.
- Each base has a conjugate acid.

Consider the reaction:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

base acid

- In this reaction, H_2O donates an H^+ to the NH_3. That makes H_2O a Bronsted-Lowry acid (H^+ donor).
- NH_3 is a Bronsted-Lowry base (H^+ acceptor).

Consider the reaction:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

base acid Conjugate acid

- In this reaction, NH_4^+ is referred to as a conjugate acid.
- A conjugate acid is the particle formed when a base gains a hydrogen ion.

Consider the reaction:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

base acid Conjugate acid Conjugate base

- In this reaction, OH^- is referred to as a conjugate base.
- A conjugate base is the particle that remains when an acid has donated a hydrogen ion.

Consider the reaction:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

base acid Conjugate acid Conjugate base

- Each acid has a conjugate base.
- Each base has a conjugate acid.

Conjugate Acid-Base Pairs

- Follow the general equation

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

Acid + Base ⇄ Conjugate acid + Conjugate Base

- This is an equilibrium.

Advantages of The Brønsted-Lowry Theory

- The Arrhenius Theory of acids and bases was limited to aqueous solutions.
- The Brønsted-Lowry Theory expands the definition of an acid and base to a proton donor or acceptor.
- This means the acid or base can be in any state.

Hydrolysis

- The Brønsted–Lowry definition can explain why a solution of sodium hydrogen carbonate (baking soda) is basic and can react with either acidic or basic solutions.
- In water, $NaHCO_3$ produces two ions:

$$NaHCO_3(s) \rightleftharpoons Na^+(aq) + HCO_3^-(aq)$$

- Sodium chloride in water produces a neutral solution, so it is not the cause of the basic solution.
- This means the hydrogen carbonate ion produces a basic solution.

$$HCO_3^-(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq) + OH^-(aq)$$

Base Acid

- The hydrogen carbonate ion accepts a proton from water, making it a base.
- In this example, water undergoes hydrolysis.
- hydrolysis means the breaking or splitting of water.

- Another example of hydrogen carbonate ion acting as a base would be . . .

$HCO_3^-(aq) + H_3O^+(l) \rightleftharpoons H_2CO_3(aq) + H_2O(l)$

Base Acid

- The hydrogen carbonate accepts a proton from the hydronium ion and becomes carbonic acid in solution.
- The hydronium ion is converted to water.

- Hydrogen carbonate ions can also act as an acid.

  $HCO_3^-(aq) + OH^-(aq) \rightleftharpoons CO_3^{2-}(aq) + H_2O(l)$

  Acid Base

- The hydrogen carbonate donates a proton to the hydroxide ion.

Practice Questions

1. Write equations for the ionization of HNO_3 and Na_2CO_3 in water.
   a) Identify the $[H^+]$ donor and $[H^+]$ acceptor in each case.
   b) Label the conjugate acid-base pairs.

Answers

1a) $HNO_3 \rightarrow H^+ (aq) + NO_3^- (aq)$

$HNO_3 + H_2O (l) \rightarrow H_3O^+ (aq) + NO_3^- (aq)$

acid base conjugate acid conjugate base

1b) $Na_2CO_3 \rightleftharpoons 2Na^+ (aq) + CO_3^{2-} (aq)$

$CO_3^{2-} (aq) + H_2O (l) \rightleftharpoons HCO_3^- (aq) + OH^- (aq)$

base acid conjugate acid conjugate base

- The Brønsted-Lowry Theory is able to explain why substances such as the hydrogen carbonate ion can act as an acid and a base.

Remember:

All Arrhenius acids are Brønsted-Lowry acids, but

NOT all Brønsted-Lowry acids are Arrhenius acids.

All Arrhenius bases are Brønsted-Lowry bases, but

NOT all Brønsted-Lowry bases are Arrhenius bases.

Lesson Summary

- Arrhenius defined acids as substances which release H+ ions in water and bases as substances which release OH^- ions in water.
- A disadvantage of the Arrhenius theory is that it does not explain why sodium hydrogen carbonate is basic but contains no hydroxide ion. It also is limited to substances dissolved in water.

- Bronsted and Lowry defined acids as proton donors and bases as proton acceptors.
- The Bronsted and Lowry definition allows for substances to behave as BOTH acids and bases, these are called amphoteric.
- According to the Bronsted-Lowry Theory, the general form for an acid-base reaction is:

acid + base ⇔ conjugate base + conjugate acid

- The conjugate acid of a base is what remains after the base accepts a proton.
- The conjugate base is what remains after an acid donates a proton.
- An acid and its conjugate base or a base and its conjugate acid are known as conjugate pairs.

Homework Assignment

- Complete the worksheet: *Conjugate Acid-Base Assignment*.

40S Chemistry

Strength of Acids and Bases

Introduction

- Solutions are described in terms of concentration, in moles per Litre.
- This does not adequately describe the reactivity of acids and bases.
- Another way to describe acids and bases is in terms of their strength, or how they react with water.

Outcomes

- Identify strong and weak acids by their reactions with water.
- Predict the products of an acid-base reaction and the position of the equilibrium.

Lesson Overview

- Strong/Weak Acids
- Strong/Weak Bases
- Acid Strength Table
- Using the Acid Strength Chart
- Acid Base Reactions

Strong Acids

- Strong acids are substances that easily donate protons.
- They must easily dissociate or ionize in water.
- The bond holding the proton to the rest of the molecule is not very strong.

- Strong acids are usually strong electrolytes.
- Strong acids are considered to COMPLETELY dissociate in water.
- For example,

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

- When hydrogen chloride dissolves in water to form hydrochloric acid, virtually no molecular HCl remains.

- We say the HCl completely ionizes, even though it is a molecular substance (that is, covalently bound).
- A single arrow indicates the complete dissociation.

Strong Acid Table

- Memorize the names and chemical formulas of these acids!

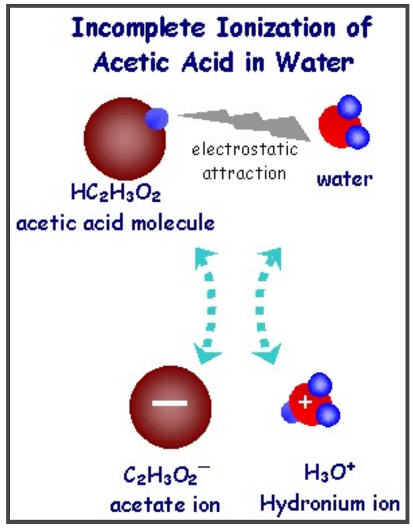
$HClO_4$	perchloric acid
HCl	hydrochloric acid
H_2SO_4	sulfuric acid
HNO_3	nitric acid
HBr	hydrobromic acid
HI	hydroiodic acid

Weak Acids

- are poor or weak electrolytes.
- do not conduct an electric current in water very well.
- are poor electrolytes because they ionize incompletely.
- have a high affinity for their proton(s). The hydrogens are bound quite strongly to the rest of the molecule, so it takes a very strong base to "rip" the proton from a weak acid.

Acetic Acid Example

- Only a small number of aqueous acetic acid molecules react with water to form an acetate ion and a hydronium ion.
- The rest exist as acetic acid molecules.



- This establishes an equilibrium between the ions and the molecules of acetic acid.
- Show this using the double arrows:

$HC_2H_3O_2$ (aq) + H_2O(l) $\rightleftharpoons$ H_3O+(aq) + $C_2H_3O_2^-$(aq)

Weak Acid Table

- Names and formulas could be memorized!

H_2CO_3	carbonic Acid
HNO_2	nitrous acid
HF	hydroflouric acid

Strong Bases

- are substances have a very high affinity for protons, H+ ions.
- are strong electrolytes.
- usually dissociate into hydroxide ions, OH^-, or oxide ions, O^{2-}. These two ions have a strong "hunger" for protons.
- They will even tear away some of the most tightly bound protons.

- The most common strong bases are sodium hydroxide, NaOH, and potassium hydroxide, KOH.

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

$$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$$

- Memorize these strong bases!

$Mg(OH)_2$	Magnesium hydroxide
$Ca(OH)_2$	Calcium hydroxide
CaO	Calcium oxide (lime)

Weak Bases

- Are poor proton acceptors.
- do not have a large affinity for protons.
- do not dissociate completely.
- form an equilibrium between the molecule and the ions formed.
- ammonia is a weak base. Its dissociation should be memorized.

$NH_3(g) + H_2O(l) \rightleftarrows NH_4+(aq) + OH^-(aq)$

Weak Bases

- Don't Memorize these weak bases

H_3NH_2	methylamine (used in production of medicines)
$C_6H_5NH_2$	aniline (used in production of synthetic dyes)
C_5H_5N	pyridine (used as a solvent and in the maufacture of food flavourings and medicines)

Acid Strength Table

- The strength of the acid decreases going down the table on the left side of the reaction arrow.
- According to the Brønsted-Lowry Theory, each acid has a conjugate base. The stronger the acid, the weaker its conjugate base. That is, the lower its attraction for protons.

- The weaker the acid the stronger its conjugate base. Notice the hydroxide ion actually has a proton to donate and can act as an acid. This is not likely, but possible. If this occurs, its conjugate base is the oxide ion which is a strong base!
- On the right side of the reaction arrow, the strong bases are at the bottom, and strength of the bases decrease as you move up the table.

Strong Acids — Strong proton donors

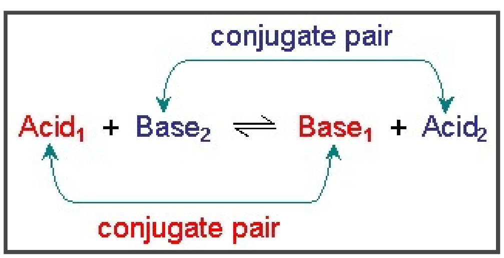
Weak Bases — Very low proton affinity

Perchloric acid	$HClO_4 + H_2O \rightarrow H_3O^+ + ClO_4^-$
Hydriodic acid	$HI + H_2O \rightarrow H_3O^+ + I^-$
Hydrobromic acid	$HBr + H_2O \rightarrow H_3O^+ + Br^-$
Hydrochloric acid	$HCl + H_2O \rightarrow H_3O^+ + Cl^-$
Nitric acid	$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$
Sulfuric acid	$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$
Hydrogen sulfate ion	$HSO_4^- + H_2O \rightarrow H_3O^+ + SO_4^{2-}$
Phosphoric acid	$H_3PO_4 + H_2O \rightarrow H_3O^+ + H_2PO_4^-$
Hydrogen telluride	$H_2Te + H_2O \rightarrow H_3O^+ + HTe^-$
Hydrofluoric acid	$HF + H_2O \rightarrow H_3O^+ + F^-$
Nitrous acid	$HNO_2 + H_2O \rightarrow H_3O^+ + NO_2^-$
Hydrogen selenide	$H_2Se + H_2O \rightarrow H_3O^+ + HSe^-$
Hydrogen sulfite ion	$HSO_3^- + H_2O \rightarrow H_3O^+ + SO_3^{2-}$
Ammonium ion	$NH_4^+ + H_2O \rightarrow H_3O^+ + NH_3$
Hydrogen carbonate ion	$HCO_3^- + H_2O \rightarrow H_3O^+ + CO_3^{2-}$
Hydrogen telluride ion	$HTe^- + H_2O \rightarrow H_3O^+ + Te^{2-}$
Monohydrogen phosphate	$HPO_4^{2-} + H_2O \rightarrow H_3O^+ + PO_4^{3-}$
Hydrogen sulfide ion	$HS^- + H_2O \rightarrow H_3O^+ + S^{2-}$
Water	$H_2O + H_2O \rightarrow H_3O^+ + OH^-$
Hydroxide ion	$OH^- + H_2O \rightarrow H_3O^+ + O^{2-}$

Weak Acids — Weak proton donors

Strong Bases — Strong proton acceptors

Acid-Base Reactions



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Example 2

- Complete the reaction below, indicating the acids and bases. which direction is favoured and why?

$$HCO_3^- + PO_4^{3-} \Leftrightarrow$$

Solution

- Identify the proton donor: HCO_3^-
- Complete the reaction by donating the hydrogen carbonate ion's proton, then label the acids and bases.

$$\underset{\text{acid1}}{HCO_3^-} + \underset{\text{base2}}{PO_4^{3-}} \Leftrightarrow \underset{\text{base1}}{CO_3^{2-}} + \underset{\text{acid2}}{HPO_4^{2-}}$$

- The forward reaction is favoured since the hydrogen carbonate ion is the stronger of the two acids. It will donate its proton.

Lesson Summary

- Strong acids and bases are strong electrolytes.
- Weak acids and bases are weak electrolytes.
- Weak acids and bases partially ionize in water.
- The stronger the acid, the weaker its conjugate base.

- The stronger the base, the weaker its conjugate acid.
- When acids and bases react, the stronger acid donates its proton and the stronger base accepts the proton.

Homework

- Complete the worksheet: *Strengths of Acids and Bases*.
- Complete the *Strengths of Acids and Bases Test*.

40S Chemistry

Ionization of Water

Introduction

- We have learned adding solutes to water can cause water to hydrolyze or become ionized.
- Water also "self-ionizes", resulting in an equilibrium between ionized and molecular water . . .

Outcomes

- Describe the equilibrium between hydronium and hydroxide ions in water and the effect of changing concentration of these ions using Le Chatelier's Principle.
- Define the ion product of water (Kw).
- Describe the realtive concentrations of hydronium and hydoxide ions in acidic, basic and neutral solutions.
- Given the concentration of hydronium or hydroxide ions, calculate the other.

Lesson Overview

- Ionization of Water
- Ion Product of Water
- Acidic and Basic Solutions
- Calculating Hydroxide Ion Concentration
- Calculating Hydronium Ion Concentration

Ionization of Water

- Water is amphoteric – it can act as both an acid and a base.

$$HA + H_2O(l) \rightleftarrows H_3O^+ (aq) + A^-(aq)$$

or

$$B + H_2O(l) \rightleftarrows BH^+(aq) + OH^-(aq)$$

- pure water dissociates into ions, or ionizes, slightly. This process self-ionization or autoionization.

- The equation for self-ionization is written as:

$$H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$$

or

$$H_2O(l) \rightleftharpoons H^+(aq) + OH^-(aq)$$

- An equilibrium established between hydronium and hydroxide ions.

Ion Product of Water

- If an equilibrium is established between hydronium ions, hydroxide ions and water molecules, an equilibrium law can be written:

$$K_a = \frac{[H_3O^+][OH^-]}{[H_2O][H_2O]}$$

$$K_a \cdot [H_2O][H_2O] = [H_3O^+][OH^-]$$

- Since water is a liquid, the product of Ka and water results in the ion product for water, Kw. The equilibrium law for water becomes:

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

- At 25°C, the concentration of the hydronium and hydroxide ions are equal at 1.0×10^{-7} mol/L. The value of Kw is constant at 1.0×10^{-14}

Acidic and Basic Solutions

- If the ionization of water occurs by the equation . . .

$$H_2O(l) + H_2O(l) \Leftrightarrow H_3O^+(aq) + OH^-(aq)$$

- The effect of dissolving an acid or base on hydronium and hydroxide ion concentrations can be predicted by using Le Chatelier's Principle.

Adding Acid

$H_2O(l) + H_2O(l) \Leftrightarrow H_3O^+(aq) + OH^-(aq)$

- When an acid is dissolved in water, the acid produces a large amount of H_3O^+ ions.
- If the $[H_3O^+]$ increases, the equilibrium will shift left to use up some of the added H_3O^+ and maintain Kw at 1.0×10^{-14}.
- Since equilibrium shifts left, the $[OH^-]$ is reduced.
- Adding a strong acid to water increases the $[H_3O^+]$ and reduces the $[OH^-]$.

Adding Base

$H_2O(l) + H_2O(l) \Leftrightarrow H_3O^+(aq) + OH^-(aq)$

- When a base is dissolved in water, the $[OH^-]$ increases.
- According to Le Chatelier's Principle, the equilibrium shifts left to use up some of the added $[OH^-]$ and maintain Kw at 1.0×10^{-14}.
- Since equilibrium shifts left, the $[H_3O^+]$ is reduced.
- Adding a strong base to water increases the $[OH^-]$ and reduces the $[H_3O^+]$.

Calculating Hydroxide Ion Concentration

Example 1

If 2.5 moles of HCl is dissolved in 5.0 L of water, what is the concentration of the hydroxide ions? Assume the volume remains unchanged.

Solution

If 2.5 moles is dissolved in 5.0 L, the concentration of the HCl would be

$$C = \frac{n}{V} = \frac{2.5 \text{ moles}}{5.0 \text{ L}} = 0.50 \text{ mol/L HCl}$$

Since HCl is a strong acid, the $[H_3O^+] = 0.50$ mol/L.

- Use the equilibrium law for water, Kw, to find the concentration of the hydroxide ion:

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

$$[OH^-] = \frac{K_w}{[H_3O^+]} = \frac{1.0 \times 10^{-14}}{0.50} = 2.0 \times 10^{-14} \text{ mol/L } OH^-$$

Calculating Hydronium Ion Concentration

Example 2

O.40 g of NaOH is dissolved in water to make a solution with a volume of 1.0 L. What is the hydronium ion concentration in this solution?

Solution

Calculate the number of moles of NaOH:

NaOH = 22.99g/mol+16.00g/mol+1.01g/mol=40.0 g/mol

$$n = \frac{m}{M} = \frac{0.40\ g}{40.0\ g/mol} = 0.010 \text{ moles of NaOH}$$

Since NaOH is a strong base, $[OH^-]$ = [NaOH] = 0.010 mol/L

- Substitute into the equilibrium law and solve for hydronium ion concentration:

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

$$[H_3O^+] = \frac{K_w}{[OH^-]} = \frac{1.0 \times 10^{-14}}{0.010} = 1.0 \times 10^{-12} \text{ mol/L } H_3O^+$$

Lesson Summary

- Water undergoes self-ionization into H_3O^+ and OH^- ions.
- The equilibrium law for water is Kw = $[H_3O^+][OH^-]$
- At 25°C, Kw = 1.0×10^{-14}
- Adding a strong acid to water decreases $[OH^-]$.
- Adding a strong base to water decreases $[H_3O^+]$.

Homework Assignment

- Complete the worksheet: *K_w Problems*.

40S Chemistry

Dissociation Constants

Introduction

- The strength of acids and bases is determined by their ability to donate or accept protons. If weak acids and bases establish an equilibrium in water, equilibrium laws can be written to describe their ionization.
- Another way to describe the degree of ionization of an acid or base is in terms of % of the original acid or base that dissociates in water.

Outcomes

- Write the equilibrium expression for an acid or base from a balanced chemical equation.
- Calculate the acid dissociation constant/base dissociation constant given the solution concentration and the hydronium/hydroxide ion concentration.
- Calculate the percent dissociation of a weak acid or base solution.
- Calculate the acid/ base dissociation constant given the percent dissociation.

Lesson Overview

- Acid Dissociation Constant
- Base Dissociation Constant
- Calculating the Dissociation Constant
- Calculating Concentration of Dissociated Species
- Percent Dissociation

Acid Dissociation Constant

- Strong acids dissociate completely, and do not establish an equilibrium.
- Weak acids establish equilibrium.
- The Acid Dissociation Constant, Ka, is the equilibrium constant for the reaction in which a weak acid is in equilibrium with its conjugate base in aqueous solution.
- The larger the Ka, the stronger the acid.

- The smaller the Ka, the less product, so the weaker the acid.
- K_a is the ratio of the concentration of the dissociated form of an acid to the undissociated form. For any weak acid, K_a follows the form . . .

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- $[A^-]$ refers to the concentration of the anion of the acid (conjugate base).
- [HA] refers to the concentration of the acid remaining after dissociation.

- K_a reflects the fraction of an acid that is in the ionized form.
- A small K_a indicates a small dissociation of the acid in the solution – a weaker acid.
- A large K_a indicates a large dissociation of the acid in the solution – a stronger acid.
- Polyprotic acids lose their H one at a time. They have a different K_a for each ionizable H.

Base Dissociation Constant

- The Base Dissociation Constant, K_b, is the equilibrium constant for the reaction in which a weak base is in equilibrium with its conjugate acid in aqueous solution.

- For any weak base, K_b follows the form . . .

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

- $[BH^+]$ refers to the concentration of the conjugate acid.
- [B] refers to the concentration of the base remaining after dissociation.

- K_b reflects the fraction of a base that is in the ionized form.
- A small K_b indicates a small dissociation of the base in the solution.
- A large K_b indicates a large dissociation of the base in the solution.

Calculating the Dissociation Constant

Example 1

- A 0.10 mol/L solution of acetic acid is only partly ionized. If at equilibrium, the hydronium ion concentration is 1.3×10^{-3} mol/L, what is the acid dissociation constant, Ka?

Solution:

Step 1

Write out the dissociation equation:

$$HC_2H_3O_2\ (aq) + H_2O(l) \Leftrightarrow H_3O^+(aq) + C_2H_3O_2^-(aq)$$

The molar ratio of hydronium to acetate is 1:1

Step 2

$$[H_3O^+] = [C_2H_3O_2^-] = 1.3 \times 10^{-3}\ \text{mol/L}$$

Step 3
Set up an "ICE" table

$HC_2H_3O_2$	$+ H_2O$	$\Leftrightarrow H_3O +$	$C_2H_3O_2^-$
0.10	-	0	0

Insert initial concentrations

Step 3
Set up an "ICE" table

$HC_2H_3O_2$	$+ H_2O$	$\Leftrightarrow H_3O +$	$C_2H_3O_2^-$
0.10	-	0	0
$- 1.3 \times 10^{-3}$	-	$+ 1.3 \times 10^{-3}$	$+ 1.3 \times 10^{-3}$

Both hydronium and the acetate ion have the same concentrations. The acetic acid will be reduced by 1.3×10^{-3} mol/L since that is converted to product.

Step 3
Set up an "ICE" table

$HC_2H_3O_2$	$+ H_2O$	$\Leftrightarrow H_3O^+$ +	$C_2H_3O_2^-$
0.10	-	0	0
$- 1.3 \times 10^{-3}$	-	$+ 1.3 \times 10^{-3}$	$+ 1.3 \times 10^{-3}$
0.0987	-	$+ 1.3 \times 10^{-3}$	$+ 1.3 \times 10^{-3}$

Step 4
Substitute into equilibrium law

$$K_a = \frac{[H_3O^+][C_2H_3O_2^-]}{[HC_2H_3O_2]}$$

$$= \frac{(1.3 \times 10^{-3})(1.3 \times 10^{-3})}{0.0987}$$

- $Ka = 1.7 \times 10^{-5}$
- There are no units for the Ka value.

Calculating the Concentration of Dissociated Species

- If we know the acid or base dissociation constant, we can calculate all the species in the solution.
- This requires the use of the "ICE" table.

Example 2

HA is a weak acid with a Ka of 7.3 x 10^{-8}. What is the concentration of all species (HA, H_3O^+ and A^-) if the initial concentration of HA is 0.50 mol/L?

Solution:
Step1
Start by writing out the dissociation equation and an "ICE" table.

HA	+ H_2O	⇔ H_3O +	A^-
0.50	---	0	0

Insert initial concentrations

Solution:
Step1
Start by writing out the dissociation equation and an "ICE" table.

HA	+ H_2O	⇔ H_3O +	A^-
0.50	---	0	0
- x	---	+ x	+ x

We don't know what the equilibrium concentrations of the hydronium and A ions are. We will represent their equilibrium concentrations with "x".

Solution:
Step1
Start by writing out the dissociation equation and an "ICE" table.

HA	+ H_2O	⇔ H_3O +	A^-
0.50	---	0	0
- x	---	+ x	+ x
0.50 - x	---	x	x

Step 2
Write the equilibrium law and substitute the values form the ICE table.

$$7.3 \times 10^{-8} \approx \frac{x^2}{0.50}$$

$$(7.3 \times 10^{-8})(0.50) = \left(\frac{x^2}{\cancel{0.50}}\right)(\cancel{0.50})$$

$$\sqrt{3.65 \times 10^{-8}} = \sqrt{x^2}$$

$$1.9 \times 10^{-4} = x$$

If we rearrange the equation, it would result in a quadratic equation. However, since Ka is small, we can assume that x is negligible compared to 0.50. We can disregard the x and simplify our calculation.

Step 2
Write the equilibrium law and substitute the values form the ICE table.

$$7.3 \times 10^{-8} \approx \frac{x^2}{0.50}$$

$$\left(7.3 \times 10^{-8}\right)(0.50) = \left(\frac{x^2}{\cancel{0.50}}\right)(\cancel{0.50})$$

$$\sqrt{3.65 \times 10^{-8}} = \sqrt{x^2}$$

$$1.9 \times 10^{-4} = x$$

$[H_3O^+] = [A^-] = x = 1.9 \times 10^{-4}$ mol/L

The concentration of the hydronium and A ions will be equal to x.

[HA] = 0.50 - x

= 0.50 - 1.9 x 10^{-4}

= 0.49981 mol/L

[HA] = 0.50 mol/L

The concentration of the HA is equal to 0.50 - x, however when we subtract x from 0.50 mol/L, we get a value which is about equal to 0.50 mol/L.
This means our assumption that x is negligible is valid!

Percent Dissociation

- Another way to describe the amount of dissociation is by percent dissociation.
- For percent dissociation, the hydronium ion, or hydroxide ion concentrations are used in the calculation . . .

- The general form of the equation is . . .

$$\text{Percent dissociation} = \frac{\text{concentration of dissociated species}}{\text{original concentration of acid or base}} \times 100\%$$

- For the acid HA:

$$\text{Percent dissociation} = \frac{[H_3O^+]}{[HA]} \times 100\%$$

- For the base BOH:

$$\text{Percent dissociation} = \frac{[OH^-]}{[BOH]} \times 100\%$$

Where [HA] and [BOH] are the initial concentrations of the acid and base.

Calculating Percent Dissociation

Example 3

Calculate the percent dissociation of a 0.100 mol/L solution of formic acid (HCH_2O_2) if the hydronium ion concentration is 4.21 x 10^{-3} mol/L.

Solution:

Substitute into the appropriate equation . . .

$$\text{Percent dissociation} = \frac{[H_3O^+]}{[HCH_2O_2]} \times 100\%$$

$$= \frac{4.21 \times 10^{-3}}{0.100} \times 100\%$$

$$= 4.21\% \text{ dissociated}$$

Calculating Dissociation Constant Given Percent Dissociation

Example 4

Calculate the K_b of the hydrogen phosphate ion (HPO_4^{2-}) if a 0.25 mol/L solution of sodium hydrogen phosphate is dissociated is 0.080%.

Solution:

Start by writing out the dissociation equation:

$$HPO_4^{2-} + H_2O \Leftrightarrow H_2PO_4^{-} + OH^{-}$$

In order to calculate K_b, we need to find the concentration of the hydroxide and the dihydrogen phosphate ions . . .

$$\text{Percent dissociation} = \frac{[OH^-]}{[HPO_4^{2-}]} \times 100\%$$

$$[OH^-] = \frac{(\text{Percent dissociation})([HPO_4^{2-}])}{100\%}$$

$$= \frac{(0.080\%)(0.25\ \text{mol/L})}{100\%}$$

$$= 2.0 \times 10^{-4}\ \text{mol/L}$$

$[OH^-] = [H_2PO4^-] = 2.0 \times 10^{-4}$ mol/L

Write the equilibrium law and substitute the values:

$$K_b = \frac{[H_2PO_4^-][OH^-]}{[HPO_4^{2-}]}$$

$$= \frac{(2.0 \times 10^{-4})(2.0 \times 10^{-4})}{0.25}$$

$$K_b = 1.6 \times 10^{-7}$$

Lesson Summary

- The acid dissociation constant, Ka, for the weak acid, HA, is

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

- The base dissociation constant, Kb, for the weak base, BOH, is (Example: KOH)

$$K_b = \frac{[B^+][OH^-]}{[BOH]}$$

- The base dissociation constant, Kb, for the weak base, B, is (Example: NH_4^+)

$$K_b = \frac{[BH^+][OH^-]}{[B]}$$

- The smaller the Ka or Kb, the weaker the acid or base.

- The percent dissociation also describes the amount of acid/base dissociated. The percent dissociated is calculated by . . .

$$\% \text{ dissociation} = \frac{[\text{ionized}]}{\text{initial [acid or base]}} \times 100\%$$

Homework Assignment

- Complete the worksheet: *Acid-Base Calculations*.
- Complete the *Acid and Base Calculations Test*.

40S Chemistry

pH

Introduction

We hear about pH so much. pH in:

- soil
- blood
- shampoo
- swimming pools

What does it really mean?

How do we measure pH?

Outcomes

- Formulate a definition of pH and pOH.
- Given any one of the values pH, pOH, $[H_3O^+]$ or $[OH^-]$, calculate the remaining values.
- Calculate the acid/base dissociation constant, given the pH or pOH and the concentration of a weak acid solution.

- Calculate the pH or pOH of a solution of a monoprotic acid or base, given its concentration and its dissociation constant
- Calculate the pH or pOH, given the concentration of a strong acid or base.
- Describe how an acid-base indicator works in terms of the colour shifts and Le Chatelier's Principle.

Lesson Overview

- Defining pH
- Logarithms
- The pH Scale
- pH Calculations
- pOH Calculations
- Calculating pH, Given Ka
- The pH Meter
- Indicators

Defining pH

- pH describes the concentration of the hydronium or hydrogen ions in solution.
- pH is defined as the negative logarithm of the hydronium ion concentration, or

  $pH = -\log[H_3O^+]$ or $pH = -\log[H^+]$
- pOH is the negative logarithm of the hydroxide ion concentration, or

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

Logarithms

- Logarithms are a way of working with very large and very small numbers.
- A number's base 10 or common logarithm is its exponent when 10 is the base. For example,

$$\log(10\,000) = 4 \qquad 10\,000 = 10^{4}$$

- The logs of other numbers are not whole numbers, but the process is the same:

$$\log(250) = 2.3979 \qquad 250 = 10^{2.3979}$$

- A number greater than one represents a positive log and a number less than one results in a negative log:

$\log(0.001) = -3$ $\quad 0.001 = 10^{-3}$

$\log(0.000\ 25) = -3.602$ $\quad 0.000\ 25 = 10^{-3.602}$

- Find the log of each of the following values:
  a) 1.3×10^{-5}
  b) 7.2×10^{-11}
  c) 0.0054

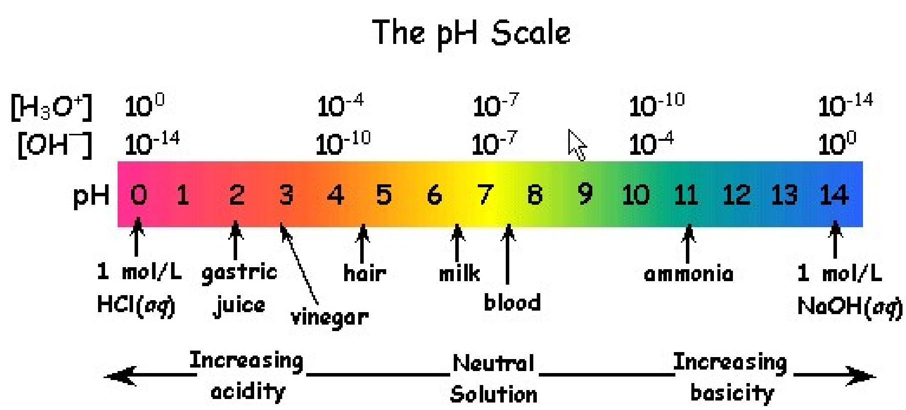
a) $\log(1.3 \times 10^{-5}) = -4.89$

b) $\log(7.2 \times 10^{-11}) = -10.14$

c) $\log(0.0054) = -2.26$

The pH Scale

- Values for pH in most solutions range from 0.0 to 14.0.
- Pure water is considered to be neutral, or a pH of 7.0.
- The lower the pH, the more acidic the solution.
- The higher the pH the more alkaline or basic the solution.



pH Calculations

Example 1

Calculate the pH of an HCl solution whose concentration is 5.0 x 10^{-6} mol/L.

Solution:

Since HCl is a strong acid, it ionizes 100%. Therefore the concentration of the HCl is equal to the hydronium ion concentration.

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

pH = $-\log(5.0 \times 10^{-6})$

pH = $-(-5.30) = 5.30$

Round the log to two decimal places.

Example 2

The pH of a solution is 3.25. Calculate the hydronium ion concentration in the solution.

Solution:

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

$[H_3O^+] = 10^{-3.25}$

$[H_3O^+] = 5.6 \times 10^{-4}$ mol/L

pOH

Recall, the ion product for water:

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

If we take the negative log of each term, we get

$$-\log(k_W) = pH + pOH$$

According to the rules of logs, when multiplying terms is equivalent to adding their logs.

If we calculate the negative log k_W,

$$-\log(kW) = -\log(1.0 \times 10^{-14}) = 14.00$$

Therefore . . .

$$pH + pOH = 14.00$$

pOH Calculations

Example 3

The pH of a solution is 10.30, what is the hydroxide ion concentration?

Solution

pH + pOH = 14.00

pOH = 14.00 – pH

pOH = 14.00 – 10.30 = 3.70

$[OH^-] = 10^{-pOH}$

$[OH^-] = 10^{-3.70}$

$[OH^-] = 2.0 \times 10^{-4}$ mol/L

Example 4

What is the pH of a 5.0×10^{-5} mol/L $Mg(OH)_2$ solution?

Solution

Mg(OH)2 is a strong base – 100% dissociation

$$Mg(OH)_2(s) \Rightarrow Mg^{2+}(aq) + 2\ OH^{-}(aq)$$

$[OH^{-}] = 2 \times [Mg(OH)_2] = 1.0 \times 10^{-4}$ mol/L

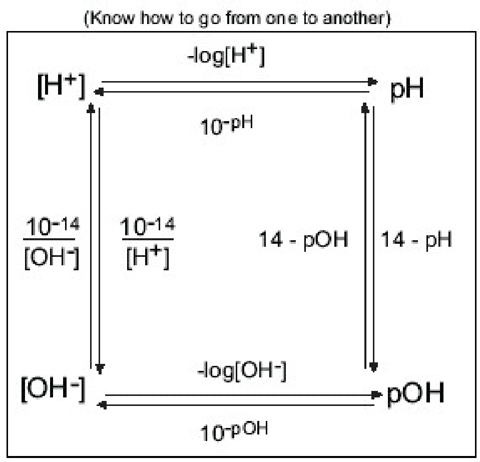
Calculate the pOH, then subtract from 14 to get the pH.

$pOH = -\log[OH^{-}] = -\log(1.0 \times 10^{-4}) = 4.00$

$14.00 = pH + pOH$

$pH = 14.00 - 4.00 = 10.00$

The pH Square



Homework Assignment

Complete the worksheets:

- *Finding pH from [$H_3$0+]* Sets 1 and 2
- *Finding pOH from [$H_3$0+]* Sets 1 and 2
- *pH Calculations* # 1 – 5

Calculating pH, Given Ka

- The pH of a weak acid (or base) solution can be calculated, given the percent dissociation or K_a (K_b)and the acid concentration.

Example 5

Calculate the pH of a 0.10 mol/L hydrogen sulfide solution. (Ka=1.0 x 10^{-7})

Solution

- We need [H_3O^+] to calculate the pH.
- Use the Ka value to calculate [H_3O^+].
- First, set up an "ICE" table.

H_2S	+ H_2O	⇔ H_3O +	HS^-
0.1	-	0	0
- x	-	+ x	+ x
0.1 - x	-	x	x

- Next, substitute these values into the Ka expression.

$$K_a = \frac{[H_3O^+][HS^-]}{[H_2S]}$$

$$1.0\times10^{-7} = \frac{x^2}{0.10 - x}$$

Assume x is negligible, Ka is very small.

$$1.0\times10^{-7} \approx \frac{x^2}{0.10}$$

$$\left(1.0\times10^{-7}\right)(0.10) = \left(\frac{x^2}{0.10}\right)(0.10)$$

$$\sqrt{1.0\times10^{-8}} = \sqrt{x^2}$$

$$1.0\times10^{-4} = x$$

$[H_3O^+] = x = 1.0 \times 10^{-4}$ mol/L

pH = –log $[H_3O^+]$ = –log(1.0×10^{-4})

pH = 4.00

NOTE:

When determining pH, we only examine the first donated proton. This is because any successive protons do not have a significant effect on pH.

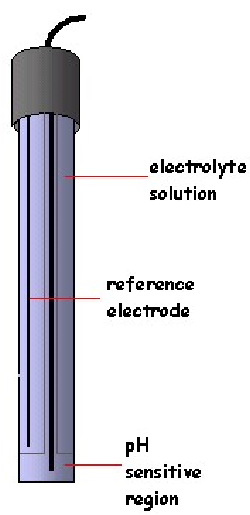
Measuring pH

There are two ways to measure pH:

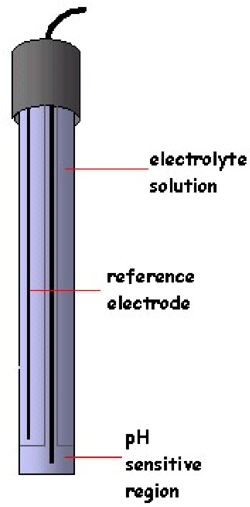
1. pH Meters
2. Indicators

pH Meters

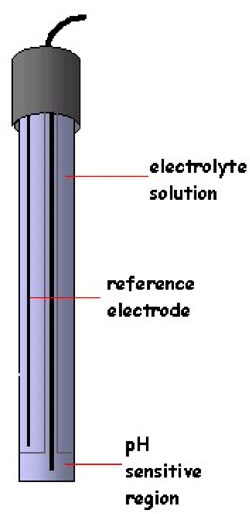
- The most important part of the pH meter is the pH probe.
- The probe consists of two parts: the hydrogen sensitive bulb at the end of the probe and the reference electrode.
- The H^+ ion concentration inside the probe is compared to the H^+ ion concentration outside the probe.



- The probe acts like a battery whose voltage changes as the pH of the surrounding solution changes.
- A small electric current is sent to a meter proportional to the hydrogen ion concentration in the solution surrounding the probe. For example, at pH 7.00 the probe produces zero volts.



- Before it is used, a pH probe must be calibrated. That is, it is inserted in a standard solution, whose pH is known.
- The meter is adjusted to read the correct pH.



Indicators

- An acid-base indicator is a weak acid or base that undergoes a colour change when they gain or lose hydrogen ions.
- The first indicators were dyes from natural sources, such as red cabbage juice, flower petal extracts, teas and fruit extracts.

- The ionization of the indicator establishes an equilibrium:

$$HIn(aq) + H_2O(l) \Leftrightarrow H_3O^+(aq) + In^-(aq)$$

colour 1 (HIn) — colour 2 (In^-)

Adding an acid increases [H_3O^+].

- According to Le Chatelier's Principle, increasing [H_3O^+] causes a shift in the equilibrium to the left, in order to use up some of the added product.
- This decreases the [In^-(*aq*)] and increases the [HIn(*aq*)]. As HIn increases, its colour will increase. As In^- decreases, its colour decreases. The solution becomes colour 1.

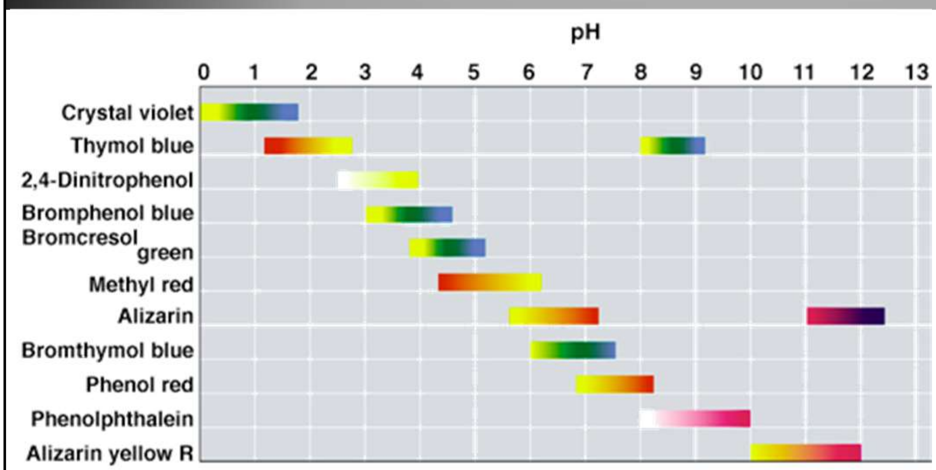
Adding a base, decreases the [H_3O^+].

- The OH^- ions from the base reacts with the H_3O^+ ions from the indicator to form H_2O.
- According to Le Chatelier's Principle, the equilibrium will shift right to replace the lost product.
- This increases the [In^-(*aq*)] and decreases the [HIn(*aq*)]. As HIn decreases, its colour will decrease. As In^- increases, its colour increases. The solution becomes colour 2.

Common Indicators

- Indicators are a convenient and inexpensive way to measure pH.
- The colour of each indicator can be coordinated with the pH and can be compared to a standard to determine the pH of the solution.
- Using an indicator is not as accurate as a pH meter.
- Indicators are also affected by the colour of the solution.

pH Range of Acid-Base Indicators



Lesson Summary

- The pH scale goes from 0 – 14. pH units describe the acidity of a solution.
- $pH = -\log[H_3O^+]$ and $[H_3O^+] = 10^{-pH}$
- $pOH = -\log[OH^-]$ and $[OH^-] = 10^{-pOH}$
- $pOH + pH = pK_a = 14$
- pH can be measured with a pH meter or pH indicators.

Homework Assignment

40S Chemistry

Neutralization

Introduction

- We have all heard or read the advertisements that claim a certain antacid will "neutralize seven-times its weight in stomach acid". How does neutralization occur? How can we make a prediction like the advertisers?
- In this lesson we will be reviewing the acid-base reaction known as neutralization.

Outcomes

- Write balanced equations for neutralization reactions.
- Calculate the concentration, volume or mass of an unknown acid or base from the mass or concentration and volume of a known acid and base required for neutralization.

Lesson Overview

- Defining Neutralization
- Balanced Neutralization Equations
- Calculating an Unknown Concentration
- Calculating Volume Needed to Neutralize

Defining Neutralization

- A neutralization reaction is one whose products are a salt and water.
- A salt is defined as a compound composed of the negative ion of an acid and the positive ion of the base
- In general, the reaction for a neutralization reaction is given by

acid + base → a salt + water

Practice

Identify the products and write balanced equations for these neutralization reactions:

1. HNO_3 + KOH → $KNO_3 + H_2O$
2. HCl + $Mg(OH)_2$ → $MgCl_2 + 2H_2O$
3. 2 H_2SO_4 + 2 NH_4OH → $(NH_4)_2SO_4 + 2H_2O$

- When hydrochloric acid and sodium hydroxide are mixed the following reaction occurs:

  $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
- This is known as the neutralization reaction. We can describe this reaction by showing the ions involved, or the ionic equation:

$$H^+(aq) + Cl^-(aq) + Na^+(aq) + OH^-(aq) \rightarrow Na^+(aq) + Cl^-(aq) + H_2O(l)$$

- Sodium ions and chloride ions are found on both sides of the arrow. We call them spectator ions since they remain unreacted through the reaction.
- The net ionic equation shows the reaction that occurs, omitting the spectator ions. The net ionic equation for this reaction omits spectator ions.

$$H^+(aq) + \cancel{Cl^-}(aq) + \cancel{Na^+}(aq) + OH^-(aq) \rightarrow \cancel{Na^+}(aq) + \cancel{Cl^-}(aq) + H_2O(l)$$

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Homework Assignment

- Complete the assignment *Neutralization Reactions*.

Balanced Neutralization Equations

- Not all neutralization reactions occur between monoprotic acids and monohydroxic bases.
- In these cases the reaction must be balanced . . .

Example 1

- Write the neutralization, the ionic and net ionic equations for the reaction between sulfuric acid and sodium hydroxide.

Solution

- For sulfuric acid, H_2SO_4, assume both protons are transferred.
- Sodium hydroxide is a strong base and has a high affinity for protons.
- The hydroxide ions will remove all protons from the sulfate ion.
- The neutralization reaction would then be:

$H_2SO_4(aq) + NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

- This is not balanced.
- Two hydroxide ions are needed to neutralize sulfuric acid's two protons.
- As a result, two water molecules are produced, along with the salt, sodium sulfate.
- The balanced equation is

The ionic equation is:

/ / / /

- Notice, the net ionic equation for ANY neutralization reaction is:

$$H^{+}(aq) + OH^{-}(aq) \rightarrow H_2O(l)$$

Homework Assignment

- Complete the assignment *Neutralization and Salts*.

Calculating an Unknown Concentration

- The equations of neutralization, permit the use of stoichiometric calculations to determine the amount of acid or base needed for neutralization.
- At neutralization,

  moles of H^+ ions = moles of OH^- ions

Example 2

- Calculate the concentration of hydrochloric acid, if 25.00 mL is just neutralized by 40.00 ml of a 0.150 mol/L sodium hydroxide solution.

Solution

Step 1 : Write the balanced neutralization reaction.

$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$

Step 2 : Calculate the number of moles of sodium hydroxide.

$C = \frac{n}{V}$

$n = C \times V$

$n_{NaOH} = (0.150 \text{ mol/L})(0.04000 \text{ L}) = 0.00600$ moles NaOH

Step 3 : Calculate the moles of hydrochloric acid used.

Since the molar ratio is one-to-one, there are 0.00600 moles of HCl in the sample

Step 4 : Calculate the concentration of the HCl.

$$C = \frac{n}{V} = \frac{0.00600\ \text{mol}}{0.02500\ \text{L}} = 0.240\ \text{mol/L}$$

Calculating Volume Needed to Neutralize

Example 3

- What volume of a 0.250 mol/L solution of carbonic acid is needed to neutralize 30.0 g of sodium hydroxide?

Solution

Step 1 : Write the balanced neutralization reaction.

$H_2CO_3(aq) + 2\ NaOH(aq) \rightarrow Na_2CO_3(aq) + 2\ H_2O(l)$

Step 2 : Determine the number of moles of base.

Find the number of moles of the reactant . . .

NaOH = 22.99g + 16.00g +1.01g = 40.00 g/mol

$$n = \frac{\text{mass}}{\text{Molar mass}} = \frac{m}{M} = \frac{30.0\ g}{40.0\ g/mol} = 0.750 \text{ moles NaOH}$$

Step 3 : Calculate moles of H_2CO_3 needed to neutralize the base.

Using the molar ratio . . .

$$\text{moles } H_2CO_3 = (0.750 \text{ moles NaOH})\left(\frac{1 \text{ mole } H_2CO_3}{2 \text{ moles NaOH}}\right)$$
$$= 0.375 \text{ moles } H_2CO_3$$

Step 4 : Calculate the volume of H_2CO_3 required.

$$V = \frac{n}{C} = \frac{0.375 \text{ moles}}{0.250 \text{ mol/L}} = 1.5 \text{ L}$$

Lesson Summary

- A neutralization reaction is a reaction between an acid and a base that produces a salt and water.
- A salt is formed from the negative ion of the acid and the positive ion of the base.
- Neutralization occurs when the moles of hydrogen ions and moles of hydroxide ions are equal.

Homework Assignment

- Complete questions 1 – 8 on the *Neutralization Worksheet*.

40S Chemistry

Buffers

Introduction

- During exercise, muscles "burn" fats and carbohydrates.
- CO_2 is a waste product released into the blood stream until it is expired out the lungs.
- Carbon dioxide dissolved in water forms H_2CO_3. Our bodies must keep the blood pH within a narrow range.
- How does our body keep its pH constant? Our blood has a built-in buffer.

Outcomes

- Define an acid-base buffer system.
- Classify a given buffer system as either acidic or basic.
- Describe how a buffer minimizes a change in the pH when either a base or an acid is added.
- Write a balanced chemical equation for the equilibrium reaction in a given buffer solution.
- Describe how buffer solutions are prepared.
- Describe the importance of buffer systems in the body.

Lesson Overview

- Buffered Solutions
- Composition of Buffers
- Acidic Buffers
- Basic Buffers
- Carbonate Buffer

Buffered Solutions . . .

- solutions that resist a change in pH when a small amount of acid (hydronium ions) or base (hydroxide ions) are added.
- contain BOTH a species which will neutralize added hydronium ions AND a species which can neutralize added hydroxide ions.

Composition of Buffers

There are four ways to make a buffer:

1. Add a conjugate base to a weak acid.

- Mix a weak acid-base conjugate pair, like acetic acid and acetate ions ($HC_2H_3O_2$ & $C_2H_3O_2^-$).
- The weak base is the salt of the conjugate base. For example, to make a buffer with acetate ions you would probably add sodium acetate and acetic acid to water.

The following equilibrium occurs:

$$HC_2H_3O_2 + H_2O \Leftrightarrow C_2H_3O_2^- + H_3O^+$$

In general,

$$HA + H_2O \Leftrightarrow A^- + H_3O^+$$

Buffers made of a weak acid and its conjugate base are called acidic buffers, as their pH's are below 7.

2. Add a conjugate acid to a weak base.

- To make a buffer of this type, the conjugate acid is the salt of the weak base. For example, ammonia and ammonium chloride.
- Buffers made of a weak base and its conjugate acid are usually referred to as basic buffers, as their pH's are usually above 7.

The following equilibrium occurs:

$$NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$$

The general equation for this type of buffer system is

$$B + H_2O \Leftrightarrow BH^+ + OH^-$$

3. Add a strong acid to a weak base.

- Adding HCl to a weak base, such as an ammonia solution, the following reaction occurs

$$HCl + NH_3 \rightarrow NH_4Cl + H_2O$$

- If the neutralization is incomplete, both NH_3 and NH_4^+ will be present in solution.

$$NH_3 + H_2O \Leftrightarrow NH_4^+ + OH^-$$

4. Add a strong base to weak acid.

- Adding NaOH to acetic acid, $HC_2H_3O_2$, the following reaction occurs:

$$NaOH + HC_2H_3O_2 \rightarrow NaC_2H_3O_2 + H_2O$$

- If the neutralization is incomplete, both the acetic acid and the acetate ion will be in solution.

$$HC_2H_3O_2 + H_2O \Leftrightarrow C_2H_3O_2^- + H_3O^+$$

Acidic Buffers

- A buffer system contains a weak acid, HA, and its conjugate base, A^-. The equilibrium equation would be

$$HA + H_2O \Leftrightarrow A^- + H_3O^+$$

Adding a Strong Acid . . .

to H_2O increases $[H_3O^{+]}$. In a buffer, the conjugate base
reacts with the acid as shown . . .

$$A^- + H_3O^+ \rightarrow HA + H_2O$$

The H_3O^+ ions:

- do not accumulate.
- are absorbed by the conjugate base
- Are converted into H_2O. The conjugate base, A^-, is converted into the weak acid, HA.
- The $[H_3O^{+]}$ and $[OH^-]$ do not significantly change. pH remains unchanged.

Adding a Strong Base . . .

to H_2O, the $[OH^-]$ increases considerably. In a buffer
system, the acid in the mixture, HA, reacts with the
OH^- as shown. . .

$$HA + OH^- \rightarrow A^- + H_2O$$

OH^- ions:

- do not accumulate.
- are neutralized by the acid and replaced by the conjugate base, A^-.
- The $[H_3O^+]$ and $[OH^-]$ do not significantly change. pH remains unchanged.

Basic Buffers

A basic buffer contains a weak base, B, and its conjugate acid, BH+. The equilibrium reaction is . . .

$$B + H_2O \Leftrightarrow BH^+ + OH^-$$

Adding a Strong Acid . . .

to H_2O increases the $[H_3O^+]$. In a basic buffer, the weak
base reacts with the H_3O^+ as shown . . .

$$B + H_3O^+ \rightarrow BH^+ + H_2O$$

The H_3O^+:

- do not accumulate.
- are neutralized by the weak base and converted into H_2O. The weak base, B, is converted into its conjugate acid, BH^+.
- The $[H_3O^+]$and $[OH^-]$ do not significantly change. pH remains unchanged.

Adding a Strong Base . . .

to water, the OH^- ion concentration increases considerably. In a buffer system, the acid in the mixture,
BH^+, reacts with the OH^- as shown . . .

$$BH^+ + OH^- \rightarrow B + H_2O$$

- The OH^- from the base are neutralized by the weak base's conjugate acid and replaced with the weak base, B.
- The $[H_3O^+]$ and $[OH^-]$ do not change significantly. pH remains unchanged.

Carbonate Buffer

- In the blood, the body produces its own buffer.
- Cellular respiration releases CO_2 into the bloodstream.
- This CO_2 is dissolved in the blood and is converted into carbonic acid, H_2CO_3.
- The body's cells cannot tolerate large pH changes.
- The body produces the bicarbonate (hydrogen carbonate) ion , HCO_3^-. The following equilibrium reaction occurs in the blood plasma:

$$H_2CO_3 + H_2O \rightarrow HCO_3^- + H_3O^+$$

- As the body's cells respire, especially during exercise, the $[H_2CO_3]$ can reach very high levels.
- pH does not change significantly.
- The H_3O^+ produced by the H_2CO_3 are neutralized by the HCO_3^- as shown . . .

$$HCO_3^- + H_3O^+ \rightarrow H_2CO_3 + H_2O$$

- If 1 mL of 10 mol/L HCl were added to 1 L of blood plasma, the plasma's pH would change from 7.4 to 7.2.
- If the same HCl were added to water, the pH would change from 7.0 to 2.0!
- If the blood becomes alkaline, the buffer system can maintain a healthy pH. The reaction proceeds as shown . . .

$$H_2CO_3 + OH^- \rightarrow HCO_3^- + H_2O$$

Lesson Summary

- A buffered solution resists changes in pH when small amounts of acids or bases are added.
- A buffer can absorb or neutralize both hydronium ions and hydroxide ions.
- Buffers are usually made of weak acid-base conjugate pairs.
- An acidic buffer is made of a weak acid and its salt or conjugate base.
- A basic buffer is made of a weak base and its salt or conjugate acid.

40S Chemistry

Titration Procedure

The Titration Process

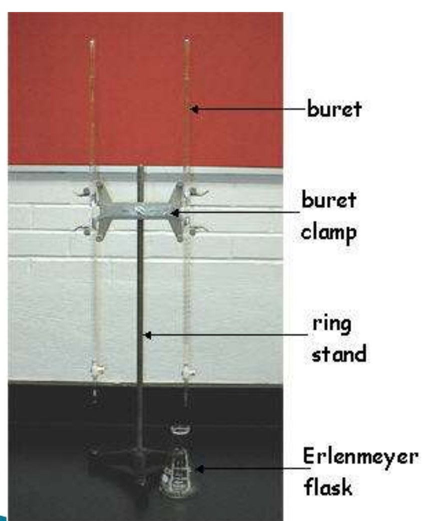
A titration is a carefully controlled neutralization reaction.

To perform a titration

- a standard solution is required.
- A standard solution is a solution of a strong acid or base of known concentration.
- To determine the concentration of an acid, a basic standard solution is required, and vice versa for an unknown base.
- The standard solution is usually added to a sample of unknown concentration, until neutralization has occurred.

- Titration is a process chemists use to determine the molarity of an acid or base by using an acid-base reaction where one reactant is of known molarity.
- In this process, the solution whose molarity is known is referred to as a standard solution.

- When an acid or base is just neutralized, the moles of hydronium ions from the acid and moles of hydroxide ions from the base are equal.
- The point at which the amount of standard acid or base solution added just neutralizes the unknown sample is called the equivalence point.



Apparatus Required:

- a buret
- solution of known *M* / solution of unknown *M*
- Erlenmeyer flask
- appropriate indicator
- wash bottle with distilled water

1. To fill a buret, close the stopcock at the bottom and use a funnel. Lift up on the funnel slightly, to allow the solution to flow in freely.

- Wear all appropriate protective equipment!

2. Before titrating, condition the buret with titrant solution.

- Make sure the buret is flowing freely.
- To condition glassware, rinse it so that all surfaces are coated with solution, then drain.
- Repeat two or three times.

3. Check the tip of the buret for an air bubble.

- To remove an air bubble, whack the side of the buret tip while solution is flowing.
- If an air bubble is present during a titration, volume readings may be in error.

4. Rinse the tip of the buret with water from a wash bottle and dry it carefully.

- After a minute, check for solution on the tip to see if your buret is leaking.
- The tip should be clean and dry before you take an initial volume reading.

5. When your buret is conditioned and filled, with no air bubbles or leaks, record the initial volume reading.

- A buret reading card with a black rectangle can help you to take a more accurate reading.
- Read the bottom of the meniscus.

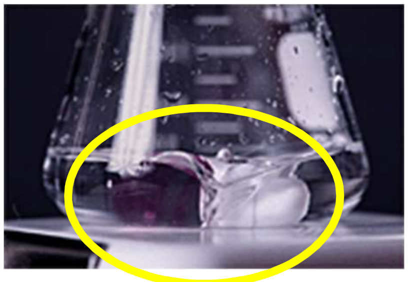
Time to begin the titration . . .

- A buret is used to deliver the second reactant to the flask and an indicator is used to detect the endpoint of the reaction.

6. Prepare the solution to be analyzed by placing it in a clean Erlenmeyer flask or beaker.

7. Add the appropriate indicator for the titration you are performing to the solution to be analyzed.

- Common indicators include Bromothymol blue, Methyl red (or orange), and Phenolphthalein.
- Phenolphthalein is used in this illustration.



8. Use the buret to deliver a stream of titrant to within a couple of mL of your expected endpoint.

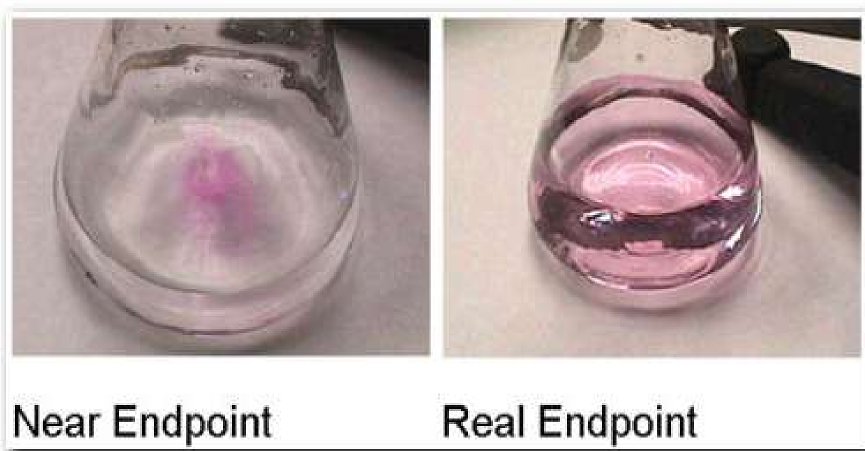
- You will see the indicator change color when the titrant hits the solution in the flask.
- The color change disappears upon stirring.

9. Approach the endpoint more slowly and watch the color of your flask carefully.

10. As you approach the endpoint, you may need to add a partial drop of titrant.

- You can do this with a rapid spin of a teflon stopcock.
- An alternate method involves partially opening the stopcock and rinsing the partial drop into the flask with a wash bottle.

- Make sure you know what the endpoint should look like. For phenolphthalein, the endpoint is the first permanent pale pink.
- If you think you might have reached the endpoint, you can record the volume reading and add another partial dro



The endpoint of the titration is signaled when a permanent color change is observed (longer than 30 seconds).
A correct endpoint is shown on the left. An overshot endpoint would be darker in color.

- If the flask looks like this, . . .
- you have gone too far!

11. When you have reached the endpoint, read the final volume in the buret and record it in your notebook.

- Subtract the initial volume to determine the amount of titrant delivered. Use this, the concentration of the titrant, and the stoichiometry of the titration reaction to calculate the number of moles of reactant in your analyte solution.
- Dispose of the titrant and analyte solutions as directed by your instructor.

Determining Concentration

Example 1

- A titration was performed using a standard solution of 0.100 mol/L NaOH into and unknown HCl solution. The following data was collected:

Determine the [ ] of the acid.

	ACID	BASE
Final Volume Reading	14.45 mL	12.57 mL
Initial Volume Reading	0.62 mL	1.13 mL

Solution:

- Determine the volume of acid and base used by subtracting the initial from the final readings:

	ACID	BASE
Final Volume Reading	14.45 mL	12.57 mL
Initial Volume Reading	0.62 mL	1.13 mL
Volume used	13.83 mL	11.44 mL

The neutralization reaction is:

$HCl(aq) + NaOH(aq) \qquad NaCl(aq) + H_2O(l)$

- At the endpoint,

  moles H^+ = moles OH^-
- Calculate the moles of base used:

  moles NaOH = C x V

  = (0.100 mol/L)(0.01383 L)

  = 0.001383 moles

- Since the acid:base stoichiometry is 1:1, the sample of acid contains 0.001383 moles.
- Determine the concentration:

$$C = \frac{n}{V} = \frac{0.001383 \text{ moles}}{0.01144 \text{ L}} = \mathbf{0.121 mol/L\ HCl}$$

- Therefore the concentration of the acid is 0.121 mol/L.

Determining the Mass of an Unknown Sample

- Since we can determine the number of moles of an acid or base, we can also determine the mass or molar mass of an unknown solid sample . . .

Example 2

- A student receives a solid sample of sulfamic acid, NH_2SO_3H, (molar mass is 97.09 g/mol), a monoprotic acid, and dissolves the sample in enough water to make 100.0 mL of solution.
- The students takes a 12.00 mL aliquot and titrates with 0.0985 mol/L sodium hydroxide. If 13.38 mL of the base is needed to reach the endpoint, what is the mass of the sample of acid?

Solution

Step 1: Determine moles of base used.

moles base = CV

= (0.0985 mol/L)(0.01338 L)

= 0.001318 moles base

Step 2: Determine moles of acid in the aliquot.

Since the acid is monoprotic, the acid:base stoichiometry is 1:1.

This means the moles of acid in the aliquot = moles of base used.

moles of acid in the aliquot = 0.001318 moles.

Step 3: Determine moles of acid in the sample.

The concentration of the 100.0 mL sample will be the same as the aliquot:

$$C = \frac{n}{V} = \frac{0.001318 \text{ moles}}{0.01200 \text{ L}} = 0.1098 \text{ mol/L acid}$$

The total moles in the 100.0 mL sample is:

moles = CV

= (0.1098 mol/L)(0.1000L)

= 0.01098 moles acid

Step 4: Determine mass of the sample of acid.

mass = nM

= (0.01098 moles)(97.09 g/mol)

= 1.07 g

The mass of the solid sample of sulfamic acid was 1.07 g.