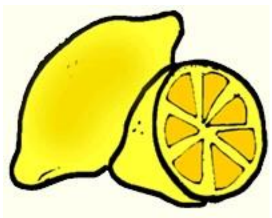
Unit 8 **(Acids and Bases)** Test Review Sheet

Reference Tables Used in This Unit- K, L, M, and T

General Properties



Acids	Bases
Similarities	
1) Are electrolytes	1) Are electrolytes
2) Change color of indicators	2) Change color of indicators
3) React with bases (in neutralization reactions) to produce water and a salt	3) React with acids (in neutralization reaction) to produce water and a salt
Differences	
4) Produce H^+ as the only positive ion in solutions	4) Produce OH^- as the only negative ion in solutions
5) Contain more H^+ than OH^- in solutions	5) Contain more OH^- than H^+ in solutions
6) When added to water, increase H^+ concentration of the water	6) When added to water, decrease H^+ ion concentration of the water
7) When added to water, decrease OH^- concentration of the water	7) When added to water, increase OH^- ion concentration of the water
8) When added to water, decrease pH	8) When added to water, increase pH
9) Have pH less than 7	9) Have pH greater than 7
10) Turn litmus red	10) Turn litmus blue
11) Have no effect on phenolphthalein (stays colorless)	11) Turn colorless phenolphthalein to pink
12) Taste sour	12) Taste bitter and feel slippery
13) React with certain metals to produce salt and hydrogen gas	

Neutral substances

1) Have pH of 7
2) Have equal amount of H^+ and OH^- ions

Acid-Base Theories

Arrhenius Theory

Arrhenius theory defines acids and bases by the ion each can produce in solutions.

Arrhenius Acids

Arrhenius acids are substances that produce H^+ (hydrogen ion, proton) in solutions.

- The H^+ produced by acids is the only positive ion in acidic solutions.
- Properties of acids (listed in the summary table) are due in parts to the properties of the H^+ ions they produce.
- The H^+ ions produced by acids usually combine with H_2O to form H_3O^+ (hydronium ion).

$$\underset{\text{hydrogen ion}}{H^+} + \underset{\text{water}}{H_2O} \rightarrow \underset{\text{hydronium ion}}{H_3O^+}$$

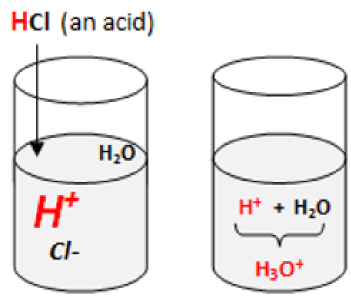


Table K: Common Acids

Formula	Name
$HCl(aq)$	hydrochloric acid
$HNO_2(aq)$	nitrous acid
$HNO_3(aq)$	nitric acid
$H_2SO_3(aq)$	sulfurous acid
$H_2SO_4(aq)$	sulfuric acid
$H_3PO_4(aq)$	phosphoric acid
$H_2CO_3(aq)$ or $CO_2(aq)$	carbonic acid
$CH_3COOH(aq)$ or $HC_2H_3O_2(aq)$	ethanoic acid (acetic acid)

Arrhenius Bases

Arrhenius bases are substances that produce OH- (hydroxide ion) in solutions.

- The OH^- produced by bases is the only negative ion in basic solutions.
- Properties of bases (listed in the summary table) are due in parts to the properties of the OH^- ions they produce.
- Most bases are ionic compounds.

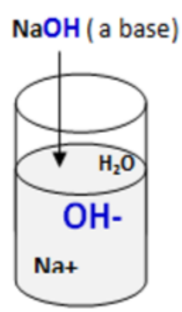


Table L: Common Bases

Formula	Name
$NaOH(aq)$	sodium hydroxide
$KOH(aq)$	potassium hydroxide
$Ca(OH)_2(aq)$	calcium hydroxide
$NH_3(aq)$	aqueous ammonia

Bronsted-Lowry (Alternate) Theory

Brönsted-Lowry Theory is an alternate theory that defines acids and bases by their ability to donate or accept a proton (H^+ or hydrogen ion) in certain reactions.

Brönsted-Lowry acids are substances that can ***donate a proton*** (H^+ or hydrogen ion) during a reaction.

Brönsted-Lowry bases are substances that can ***accept a proton*** (H^+ or hydrogen ion) during a reaction.

H_2O	+	NH_3	$\leftrightarrow$	NH_4^+	+	OH^-
acid		*base*		*conjugate acid*		*conjugate base*

H_2O is an acid because it gives up an H^+ and becomes **OH- (a base)**
NH_3 is a base because it accepts the H^+ and becomes NH_4^+ **(an acid)**

When Brönsted-Lowry acids and bases react:

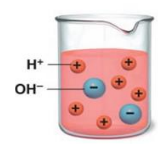
Two conjugate **acid** - **base** pairs can be determined:

*An **<u>amphiprotic</u>** substance is a substance that can act as either an alternate theory acid or base
H_2O is an example of a substance that is amphiprotic

Relative Ion Concentration of Acids and Bases

All aqueous solutions contain both H^+ and OH^- ions. A solution can be defined as acidic or basic depending on the relative concentration of H^+ and OH^- ions in the solution.

Acidic solutions contain *more* (higher concentration of) H^+ ions than OH^- ions.
The stronger the acid, the greater the H^+ ion concentration in comparison to the OH^- ion concentration.
Example: HCl*(aq)* (hydrochloric acid solution) contains more H^+ ions than OH^- ions.



Acidic solution

Neutral solutions and pure water contain *equal* amounts of H^+ and OH^-.
Example: NaCl (a neutral salt solution) contain equal concentrations of H^+ and OH^- ions.


Neutral solution

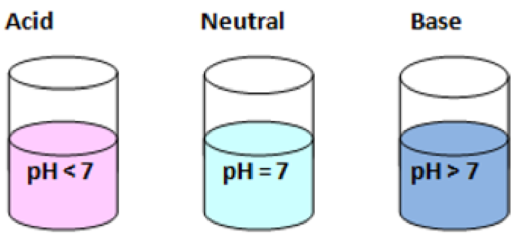
Basic solutions contain *more* (higher concentration of) *OH^-* ion than H^+ ion.
The stronger the base, the greater the OH^- ion concentration in comparison to H^+ ion concentration.
Examples: NaOH*(aq)* (sodium hydroxide solution) contains more OH^- ions than H^+.


Basic solution

The pH Scale and H^+ ion concentration

pH is a measure of the hydrogen ion (H^+) or hydronium ion (H_3O^+) concentration of a solution. A pH scale ranges in value from 0 - 14. Acids and bases can be defined by their pH values.

- **Acids** are substances with pH values *less* than 7.
- **Neutral** substances have pH *equal* to 7.
- **Bases** are substances with pH values *greater* than 7.



The strength of an acid or a base can be determined by its pH value:

Strong acids have very low pH values.

Strong bases have very high pH values.

A typical ***pH scale*** is shown below. Some common substances are indicated below the scale to show their approximate pH values.

Acid		Neutral		Base	
Strong	Weak		Weak		Strong

1-----------------------------------7-------------------------------14

H_2O

HCl   Acetic acid NH_3 NaOH

Mathematically, **pH** is defined as the –log of H^+ ion concentration of a solution.

pH = -log [H^+]   or   [H^+] = 10^{-pH}

pH is, therefore, a measure of how much (concentration of) **H^+** or H_3O^+ ion are in a solution.

A calculator is not required to calculate pH if the H^+ concentration is given as 1×10^{-x} M. The **x** being a value (between 1 to 14) of a negative exponent.

When the [H^+] or [H_3O^+] of a solution is given in the form of 1×10^{-x} M, the **pH** value of the solution = **x**

Examples are given below.

[H_3O^+]	*pH Value*	*Type of Solution*
1.0×10^{-2} M	2	Acidic
1.0×10^{-8} M	8	Basic

As you know a solution with a pH of 3 is more acidic than a solution with a pH of 4. Since it is the concentration of H^+ ions that determines the pH of the solution, a solution with a pH of 3 has a greater concentration of H^+ ions than a solution with a pH of 4.

- The lower the pH, the more H^+ ions there are in the solution
- As [H^+] in a solution increases, pH of the solution decreases

How much more (or fewer) H^+ is in one solution in comparison to another solution can be determined if the pH values of the two solutions are known. Because of the mathematical relationship between H^+ concentration and pH (shown above):

Difference in H^+ of two solutions = $10^{\text{(difference in pH)}}$

That means:

1 value difference in pH = 10 times (fold) difference in [H^+]

Acid-Base Indicators

Acid–Base indicators are substances that change color in the presence of an acid or a base.

Acids and bases can be defined by changes they cause on indicators. Two common acid-base indicators are litmus paper and phenolphthalein.

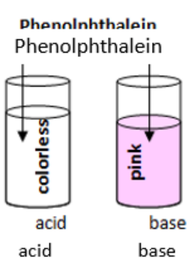
Phenolphthalein is substance that dissolves in water to produce a colorless indicator solution.

Acids are substances that have *no effect* on phenolphthalein

- Phenolphthalein stays colorless in the presence of an acid.

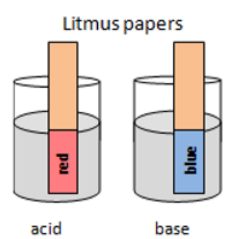
Bases are substances that change colorless phenolphthalein to *pink*.

- Phenolphthalein is a good indicator to test for the presence of a base



Litmus papers come in varieties of colors. When wet with an acidic or a basic solution, a litmus will change its color.

- *Acids* are substances that will change litmus to *red.*
- *Bases* are substances that will change litmus to *blue.*



Other common indicators are listed on Reference Table M below.

Table M
Common Acid–Base Indicators

Indicator	Approximate pH Range for Color Change	Color Change
methyl orange	3.1–4.4	red to yellow
bromthymol blue	6.0–7.6	yellow to blue
phenolphthalein	8–9	colorless to pink
litmus	4.5–8.3	red to blue
bromcresol green	3.8–5.4	yellow to blue
thymol blue	8.0–9.6	yellow to blue

Reading Table M:

Methyl orange will be:
Red in pH below 3.1
Yellow in pH above 4.4

Thymol blue will be:
Yellow in pH below 8.0
Blue in pH above 9.6

You must know how to use the various indicator color changes to predict relative acidity/ basicity of a substance.

Neutralization Reactions

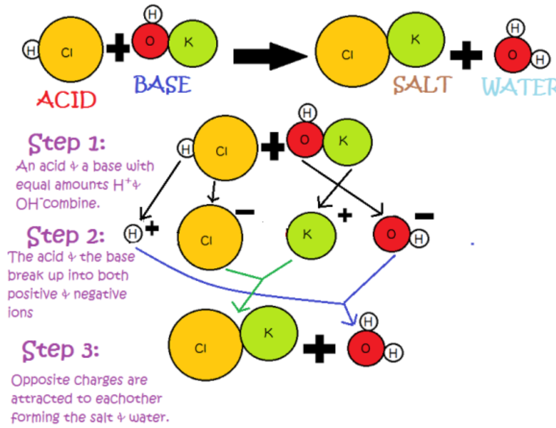
Neutralization is a reaction between an acid and a base to produce *water* and a *salt.*
During neutralization reactions:

- Equal moles of H^+ (of the acid) and OH^- (of the base) combine to neutralize each other
- Water and salt are produced
- The salt formed depends on the acid and the base that reacted

Below, general equation, example equation, and net ionic equation showing a neutralization reaction are given:

	Reactants				*Products*		
General equation:	Acid	+	Base	→	Water	+	Salt
Example reaction equation:	HCl (hydrochloric acid)	+	$NaOH$ (Sodium hydroxide)	→	H_2O	+	$NaCl$ (sodium chloride)
Formation of water in neutralization	H^+	+	OH^-	→	H_2O		

Neutralization reactions are double replacement type reactions.
The water (H_2O) is formed from the H^+ ion of the acid and the OH^- of the base.
The salt ($NaCl$) is formed from the metal of the base (Na) and the non-hydrogen part of the acid (Cl)

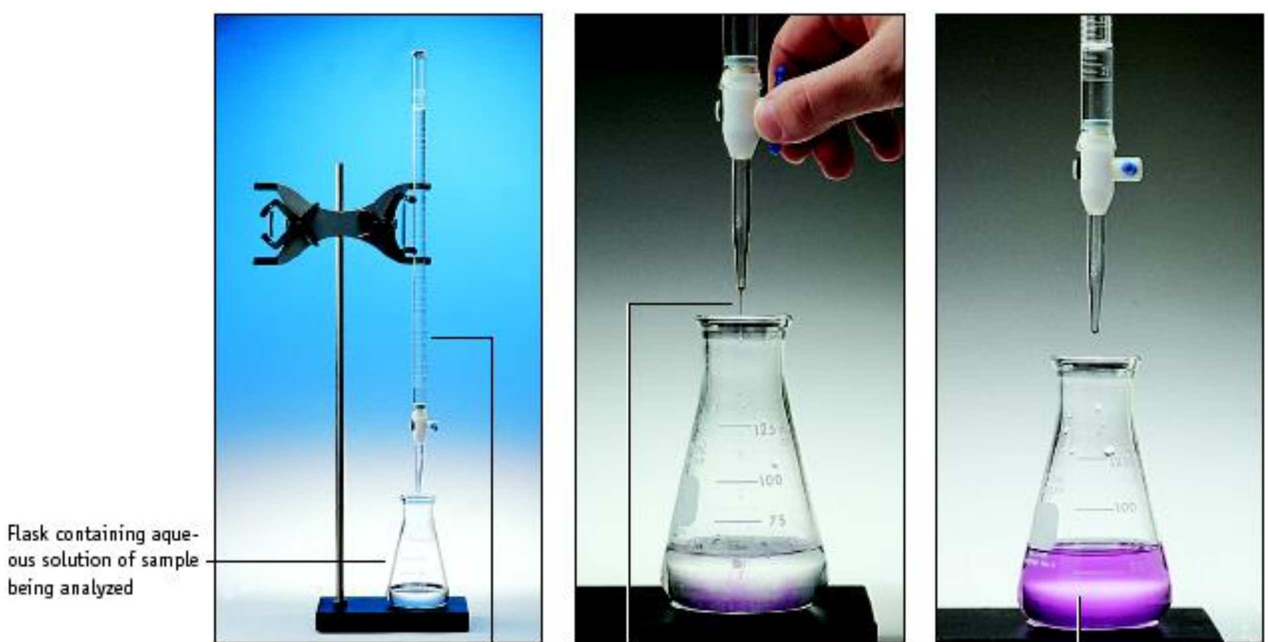


Titration

Titration is a method used to quantitatively **determine concentration** of an unknown acid or base

*Be able to describe the titration set-up, including the buret measurements

Ex: Titrating an Acid with a Base



(a) 50-mL buret containing aqueous NaOH of accurately known concentration

(b) A solution of NaOH is added slowly to the sample being analyzed. The sample is mixed.

(c) When the amount of NaOH added from the buret exactly equals the amount of H⁺ supplied by the acid being analyzed, the dye (indicator) changes color.

Titration involves a neutralization reaction between an acid and base

The "end point" is reached when the neutralization is complete and the indicator changes color

The equivalence point is when the # H^+ ions = #OH^- ions.

The acid or base of known concentration usually is called the Standard Solution

Titration Equation to Solve Problems:

$M_A \times V_A = M_B \times V_B$ **Table T Equation**	Use this equation if the mole ratio of H^+ (of acid) to OH^- (of base) is 1 : 1. ***For example:*** A titration involving HCl and NaOH has **1**H^+ : **1**OH^- ratio.
$\#H^+ \times M_A \times V_A = \#OH^- \times M_B \times V_B$	Use this equation if the mole ratio of H^+ to OH^- *is not* 1: 1. ***For examples:*** A titration involving H_2SO_4 and NaOH has **2**H^+ : **1**OH^- ratio. A titration involving HCl and $Ca(OH)_2$ has **1**H^+ : **2**OH^- ratio.

Concept Task: Be able to solve for an unknown in a neutralization problem using the titration equation.

Example 1

30 mL of 0.6 M HCl solution is neutralized with 90 mL NaOH solution. What is the concentration of the base?

$M_A \times V_A = M_B \times V_B$

$0.6 \times 30 = M_B \times 90$

$\frac{0.6 \times 30}{90} = M_B$

numerical setup

0.2 M = M_B *calculated result*

Example 2

How many milliliters of 1.5 M H_2SO_4 are needed to exactly neutralize 20 milliliters of a 1.5 M NaOH solution?

$\#OH^- \times M_B \times V_B = \#H^+ \times M_A \times V_A$

$1 \times 20 \times 1.5 = 2 \times 1.5 \times V_A$

$\frac{1 \times 20 \times 1.5}{2 \times 1.5} = V_A$

setup

10 mL = V_A *calculated result*