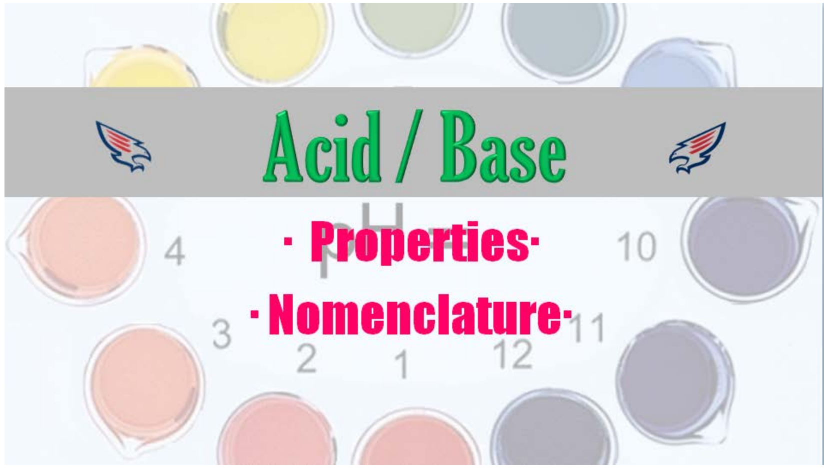


Acids and Bases

Brackets, [ ], mean *concentration* or...

MOLARITY (M)

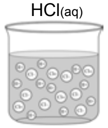
So, "$[H^+] = 1M$" means...

The *concentration of hydrogen ions* is 1 M.

Acid

A compound that produces **H^+** ions when dissolved in water.

HCl(aq)



Examples!

Vinegar - **Acetic** acid

Lemon Juice - **Citric** acid

Sour Candy - **Malic** acid (and others)

Milk - **Lactic** acid

Acid Properties

- **Sour** taste
- Reacts with (corrodes) **metals** to form **H_2** gas

Click here

$$Zn(s) + HCl(aq) \longrightarrow ZnCl_2(aq) + H_2(g)$$

- Turns blue litmus paper **red**

- **Aqueous** solutions of acids are **electrolytes.**
- React with BASES to form **salt** and **water**.

Click here

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

Acid Properties

Acids react with (corrode) metals to produce hydrogen gas and an ionic compound.

back to properties

Acid Properties

When an acid is mixed with a base, water is produced.

back to properties

Naming Acids

Binary Acids - ***H*** bonded to ***one nonmetal*** (HF)

- Begin with _hydro_, use the _root_ of the anion name, add the suffix _-ic_.

    1. HBr - _hydrobromic_ acid

Ternary Acids - ***H*** bonded to ***a polyatomic ion*** (H_2SO_4)**(a.k.a oxo-acids or oxyacids)**

- Begin with _polyatomic_ ion without the _suffix_.
- Add suffix _-ic_ if polyatomic ion ended in _-ate_.
- Add suffix _-ous_ if polyatomic ion ended in _-ite_.

    1. HNO_3 - _Nitric_ acid
    2. $HClO_2$ - _Chlorous_ acid

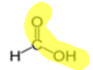
Naming Acids

***Note:** when the acid anion contains sulfur or phosphorus...

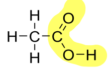
- the roots are sulfur- and phosphor-, respectively,
- not sulf- and phosph-.
- So H_2SO_4 is sulfuric acid, not sulfic acid.
- And H_3PO_4 is phosphoric acid, not phosphic acid.

Naming Acids

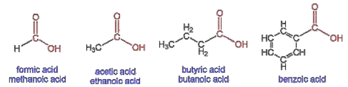
Organic acids - on carbon chains. Called carboxylic (carb - oxyl - ic) acids - contain a carboxyl group - carbon doubly bound to oxygen and singly bound to hydroxide.



formic acid



acetic acid



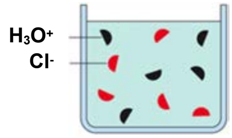
Formula	Common Name	Source	IUPAC Name	Melting Point	Boiling Point
HCO_2H	formic acid	ants (L. formica)	methanoic acid	8.4 °C	101 °C
CH_3CO_2H	acetic acid	vinegar (L. acetum)	ethanoic acid	16.6 °C	118 °C
$CH_3CH_2CO_2H$	propionic acid	milk (Gk. protus prion)	propanoic acid	-20.8 °C	141 °C
$CH_3(CH_2)_2CO_2H$	butyric acid	butter (L. butyrum)	butanoic acid	-5.5 °C	164 °C
$CH_3(CH_2)_3CO_2H$	valeric acid	valerian root	pentanoic acid	-34.5 °C	186 °C
$CH_3(CH_2)_4CO_2H$	caproic acid	goats (L. caper)	hexanoic acid	-4.0 °C	205 °C
$CH_3(CH_2)_5CO_2H$	enanthic acid	vines (Gk. oenanthe)	heptanoic acid	-7.5 °C	223 °C
$CH_3(CH_2)_6CO_2H$	caprylic acid	goats (L. caper)	octanoic acid	16.3 °C	239 °C
$CH_3(CH_2)_7CO_2H$	pelargonic acid	pelargonium (an herb)	nonanoic acid	12.0 °C	253 °C
$CH_3(CH_2)_8CO_2H$	capric acid	goats (L. caper)	decanoic acid	31.0 °C	219 °C

Acids Strength

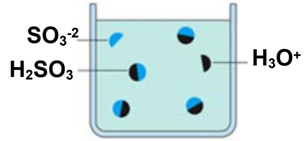
STRONG Acids - **completely** ionize, or ***dissociate***, in water. All the acid breaks into ions. To dissociate is to form IONS!

WEAK Acids - do not ionize completely in solution. Some acids stays intact instead of ionizing.

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



$H_2SO_3 + H_2O \leftrightarrow H_3O^+ + SO_3^{-2}$



Acids Strength

Binary Acid Strength

- There are **3 STRONG Binary** acids
  - > **HCl, HBr, HI**
- ALL other binary acids are *weak*: H_2S, H_2Se, HF

Ternary Acid Strength

- There are **4 STRONG Ternary** acids
  - > **H_2SO_4, HNO_3, $HClO_3$, $HClO_4$**
- ALL other ternary acids are *weak*: H_3PO_4, H_2SO_3, HNO_2, etc...

Acids Strength

KNOW the
STRONG SEVEN

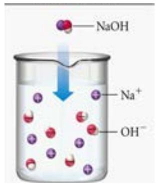
H_2SO_4 HNO_3 $HClO_3$ $HClO_4$

HCl HBr HI

Acid Base 2016 OTHS

Bases

A compound that produces **OH⁻**,
hydroxide ions when dissolved in water.



Ex.

Milk of Magnesia- **Magnesium** hydroxide

Drain Cleaner - **Sodium** hydroxide

Window Cleaner - **Ammonia, NH_3**

NH_3 in water makes NH_4^+ and OH^- !!

Base Properties

- **Bitter** taste
- Feels **slippery** to the touch.
- Turns red litmus paper **blue**

- **Aqueous** solutions of bases are **electrolytes.**
- React with ACIDS to form **salt** and **water**.

$$NaOH + HCl \longrightarrow NaCl + H_2O$$

Naming Bases

Use the same rules as for **polyatomic** ions (name the cation, then name the anion).

1. NH_3 - known informally as AMMONIA
2. NaOH - ________________
3. $Ca(OH)_2$ - ________________
4. KOH - ________________
5. $Mg(OH)_2$ - ________________

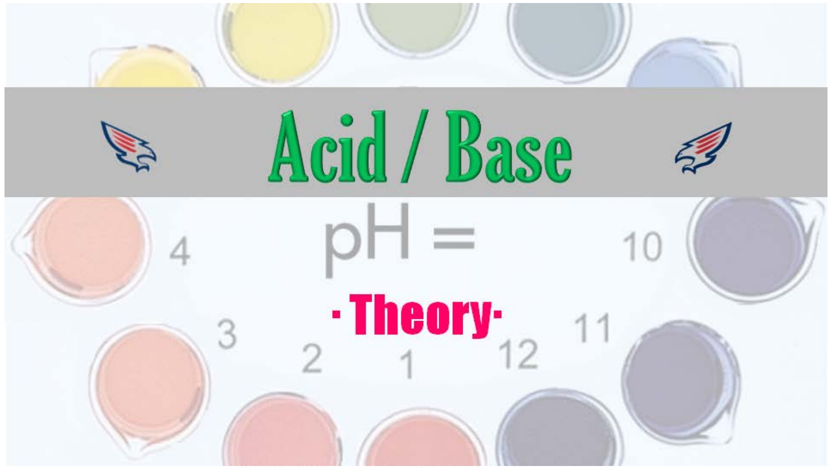
Base Strength

STRONG bases - **completely** ______ in water.
All bases break into ions.

- All hydroxides with groups **1** and **2** metals (except **Be**) are **STRONG bases**.
- LiOH, NaOH, KOH, RbOH, CsOH
- $Mg(OH)_2$, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, $Ra(OH)_2$

WEAK bases - **ionize** only **slightly**.
Some of the base stays intact instead of ionizing.

- All bases NOT LISTED ABOVE AS STRONG like $Al(OH)_3$, NH_3



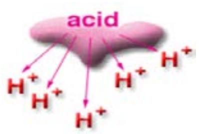
The Hydronium ion

Hydrogen ions are not really "____________" in an acidic solution. Water molecules strip the hydrogen from the anion forming the ____________________ ion.This happens when an acid is dissolved in H_2O. H_2O and H^+ combine to form H_3O^+.

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

$[H_3O^+]$ and $[H^+]$ are interchangeable, you will see both used throughout chemistry.

Arrhenius Theory



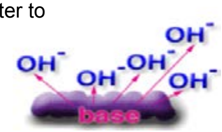
An Arrhenius ACID ionizes in aqueous solution to produce ____________________.

$$HCl + H_2O \rightarrow \mathbf{H_3O^+} + Cl^-$$

An Arrhenius BASE contains hydroxide ions and ionizes in water to produce ____________________.

$$NaOH + H_2O \rightarrow Na^+ + \mathbf{OH^-}$$

$$NH_3 + H_2O \rightarrow NH_4^+ + \mathbf{OH^-}$$



Arrhenius Theory

Practice:

Classify each of the following as an Arrhenius acid or Arrhenius base.

$Ca(OH)_2$	**Arrhenius BASE**
HBr	**Arrhenius ACID**
H_2SO_4	**Arrhenius ACID**
LiOH	**Arrhenius BASE**

Arrhenius ACID

Arrhenius BASE

Acid Base 2016 OTHS

Arrhenius Theory

Disadvantages of Arrhenius Theory

- Can only be applied to reactions that occur in **water**. And acid-base reaction can occur in gas form.
- Some bases, such as ammonia, do not contain hydroxide ions. They are formed when they contact water.

The Hydrogen Ion

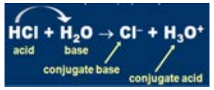
A hydrogen ion (H^+) is a PROTON!

A hydrogen atom contains a single proton and an electron, so when it loses an _______electron_______ all that remains is a PROTON.

Hence, hydrogen ions (H^+) are often referred to as _______protons_______.

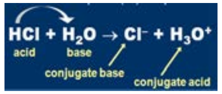
Bronsted-Lowry Theory

In **Bronsted-Lowry** theory, substances are defined by their ACTIONS. Its all about the exchange of _______protons_______.



- The Bronsted-Lowry **acid** is the **reactant** that **DONATES** a **proton (H^+)**
- The Bronsted-Lowry **base** is the **reactant** that **ACCEPTS** a **proton (H^+)**

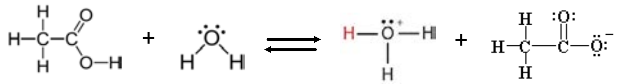
Bronsted-Lowry Theory



- The Bronsted-Lowry **Conjugate acid** is the **product** that forms after the B-L _____BASE_____ accepts an H+
- The Bronsted-Lowry **Conjugate base** is the **product** that forms after the B-L _____ACID_____ has donated an H+

Identify Bronsted Lowry A & B!

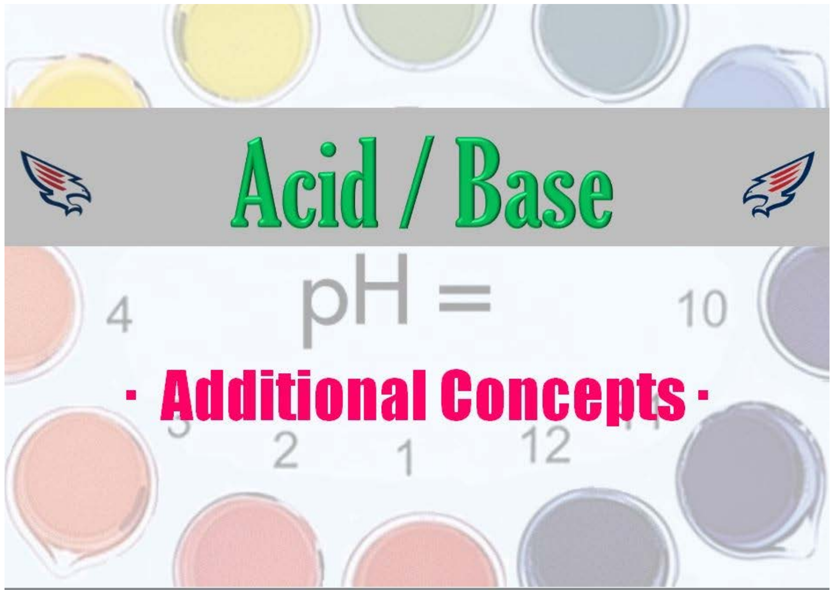
Use this mental process to build *2 conjugate pairs* between reactants and products.



1. *Which reactant loses an H+?* This is the _______ Its partner in the products (missing an H+) is the **conjugate** ________. **Connect with an arrow.**
2. *Which reactant gains an H+?* This is the _______ Its partner in the products (with its new H+) is the **conjugate** ______. **Connect with an arrow.**
3. Remember: Conjugates are always ______________!

Bronsted Lowry Conjugates

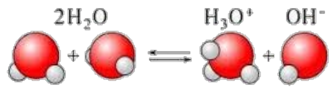
- A strong acid makes a ____________ conjugate base
- A weak acid makes a ____________ conjugate base



Amphoteric Substances

Amphoteric Substances –act as an ______________ in some reactions (accept H+) and as a ________________ in others (donate H+).

Examples: H_2O, NH_3



How Protic Is It?

★ Acids can be defined by how many H^+ ions they can **donate.**

Type	# of H^+ to give	Example
•	•	•
•	•	•
•	•	•

How Basic Is It?

★ Bases can be defined by how many H^+ ions they can **accept.**

Type	# of H^+ to accept	Example
•	•	•
•	•	•
•	•	•

Strength vs. Concentration

Compare solutions by type AND *relative* concentration

0.1M HCl vs. 0.1M HBr

Both strong, same concentration

0.01M HCl vs. 0.1M HBr

Both strong, HBr more concentrated

0.1M HCl vs. 0.1M $HC_2H_3O_2$

HCl strong, same concentration

0.1M $H_3C_6H_5O_7$ vs. 0.01M H_3PO_4

Both weak, $H_3C_6H_5O_7$ more concentrated

Acidic and Basic Salts

1) Acidic salts – formed when a strong acid and weak base react

Ex: $AlCl_3$

- Parent acid – HCl (strong)
- Parent base – $Al(OH)_3$ (weak)
- Solutions of acidic salts have a pH >5 and <7. When placed in water, form an acidic system.

Acidic and Basic Salts

2) Basic salts – formed when a weak acid and strong base react

Ex: $LiC_2H_3O_2$

- Parent acid - $HC_2H_3O_2$ (weak)
- Parent base – LiOH (strong)
- Solutions of basic salts have a pH >7 and <9

Acidic and Basic Salts

3) Neutral salts – formed when a strong acid and strong base react

Ex: LiCl

- Parent acid – HCl (strong)
- Parent base – LiOH (strong)
- Solutions of neutral salts have a pH of 7

Acidic and Basic Salts

Salt	**Predicted pH**	**Parent Acid**	**Parent Base**
$NaNO_3$			
$Fe_2(SO_4)_3$			
$Ca_3(PO_3)_2$			

Self-Ionization of Water

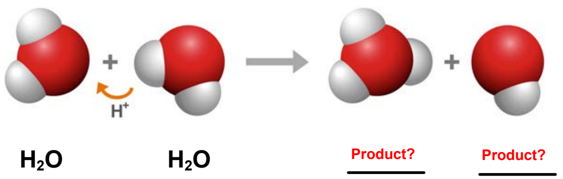
Lets Label!

B-L Acid

B-L Base

Conjugate Base

Conjugate Acid



Self-Ionization of Water

In any sample of ***water***, small but equal amounts of H^+ and OH^- ions will form, creating ***conjugate pairs***. This is called the **self -**_________ of water. About 1:2,000,000,000 water molecules does this.

This also happens in pure ammonia (NH_3) and other pure, polar substances.

$$NH_3 + NH_3 \rightarrow$$

Self-Ionization of Water

Equal concentrations of $[H^+]$ and $[OH^-]$ are present at 1×10^{-7} M at room temperature.

- *This is NEUTRAL;*
- *the total concentration of ions in any aqueous solution is:* ***1×10^{-14} M^2.***
- *(Multiply the molarities of $[H^+]$ and $[OH^-]$ together)*

This is **K_w,** the **ion product constant** for **w**ater
(K means constant)

For pure water $K_w = [H^+] \times [OH^-] = (1 \times 10^{-7} M)^2 = 1 \times 10^{-14}$ M^2

Self-Ionization of Water

So... $[H^+] \times [OH^-] = 1 \times 10^{-14}$ M^2

As $[H^+] \uparrow$, $[OH^-]$ must $\downarrow$

As $[OH^-] \uparrow$, $[H^+]$ must $\downarrow$

This is an **inverse** relationship between hydrogen and hydroxide ions. When $[H^+]$ dominates you have an acid, with greater $[OH^-]$, a base.

Calculating [H+] and [OH-]

*Using the **ion product constant for water (Kw)**, we can solve for the acid or base counterpart of a known value because we are working with an aqueous solution!*

$K_w = [H^+] \times [OH^-]$

Mathematically, the formula can be manipulated to show:

$$[H^+] = \frac{1\,x\,10^{-14}\ M^2}{[OH^-]} \quad \textbf{OR...} \quad [OH^-] = \frac{1\,x\,10^{-14}\ M^2}{[H^+]}$$

Ex: $[OH^-]=2.5 \times 10^{-5}$ M: Calculate $[H^+]$ $[H+] = 1x10^{-14}/2.5x10^{-5} = 4.0x10^{-10}$ M

Ex: $[H^+]=8.90 \times 10^{-2}$ M: Calculate $[OH^-]$. $[OH-] = 1x10^{-14}/8.90x10^{-2} = 1.12x10^{-13}$ M

Calculating [H+] and [OH-]

In an acid...

*$[H^+] > [OH^-]$ so… 1×10^0 > **$[H^+]$** > 1×10^{-7} M*

In a base ...

*$[OH^-] > [H^+]$ so… 1×10^0 > **$[OH^-]$** > 1×10^{-7} M*

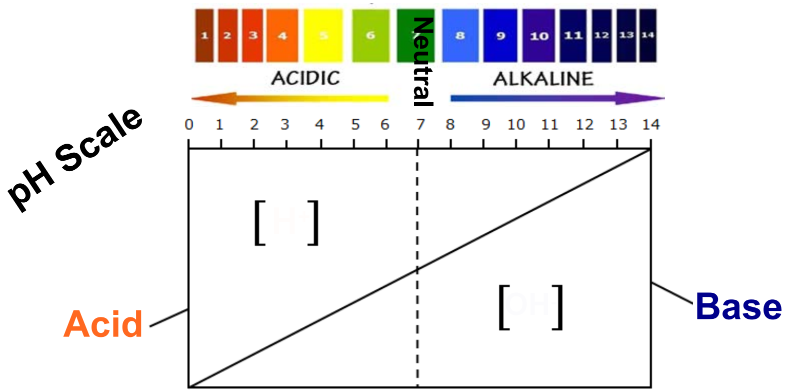
Acidic or basic?

A. $[H_3O^+] = 1 \times 10^{-3}$ ___acidic___

B. $[H_3O^+] = 1 \times 10^{-11}$ ___basic___

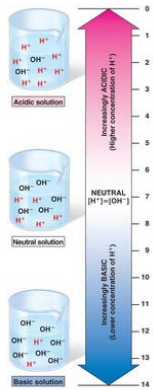
C. $[OH^-] = 1 \times 10^{-4}$ ___basic___

Finding pH with Logarithms



pH Scale

- pH range is from **0** to **14**.
- pH 0-7 is acidic, 7 is neutral, 7-14 is basic
- Using a calculator, we use the [H+] and the "**log**" key to generate pH.
- pH is a base 10 **log**arithm
- **we do this to work on the whole number scale**

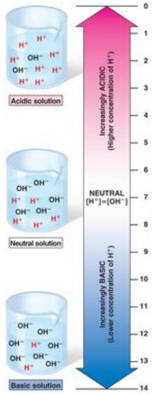


pH Scale

- A ***logarithm*** is the _______________ to which 10 must be raised to create a certain number. In this case, the *power* is pH or pOH.
- ***log*** $\boldsymbol{10^{Y}} = \boldsymbol{Y}$
  > *ex:* log 10^{1} = 1> *ex:* log 10^{-5} = –5
- pH is a _____________ log, because we want pH to be positive.

pH formula

IF $[H_3O^+] = 1 \times 10^{-8}$, then pH = ______



pH Scale

Ex1: Find pH if $[H_3O^+] = 1.0 \times 10^{-4}$ M

Ex2: Find pH if $[H_3O^+] = 1 \times 10^{-13}$ M

Ex3: Find pH if $[H_3O^+] = 1 \times 10^{-5}$ M

Ex4: Find pH if $[OH^-] = 1 \times 10^{-5}$ M

If the concentration is not a whole number *exponent,* we must use the formula.

pH Scale

pH = -log[H⁺]

To convert from [H⁺] to pH with most any calculator

IF $[H^+] = 3.09 \times 10^{-10}$M

Press (–) log 3.09 EE (–) 10 *pH is 9.510*

Sometimes the log key is log_{10} *Use (–) or +/- not "minus"*

SigFigs and pH

The # of sigfigs in the coefficient of the concentration…

[H⁺] = <u>3.09</u> x 10^{-10}M

Is the number of digits to keep AFTER THE DECIMAL in the pH.

pH = –log (<u>3.09</u> x 10^{-10}) = 9.<u>510</u>

Ex: Find pH if $[H_3O^+] = 5.0 \times 10^{-6}$M

Ex6: Find pH if $[H_3O^+] = 2.5 \times 10^{-9}$M

pOH Scale

- We can examine the base ion concentration by calculating pOH in the same manner as pH
- ***pOH is the mirror image of pH. Perfectly opposite.***
- pOH range is also from ______ to _______ and fluctuates inversely with pH
- **pOH 0 to 7 is basic**
- **pOH = 7 is neutral**
- **pOH 7 to 14 is acidic**

pOH Scale

Ex. What is the pOH if $[OH^-] = 1 \times 10^{-4}$

Ex. What is the pOH if $[OH^-] = 1 \times 10^{-3}$

Ex. What is the pOH if $[H_3O^+] = 1 \times 10^{-7}$

Ex. What is the pOH if $[H_3O^+] = 1.34 \times 10^{-8}$

pH and pOH

- ***Recall:*** **$[H^+] \times [OH^-] = 1 \times 10^{-14}$ M**
- Because pH and pOH are also inversely related…

pH + pOH = 14

Ex11. pH = 6 ; pOH = ___

Ex12. pH = 4 ; pOH = ___

Ex13. pOH = 3 ; pH = ___

Ex14. pOH = 11 ; pH = ___

Find $[H_3O^+]$ and $[OH^-]$

Given pH or pOH, we can determine _________ by using the inverse of the log! To find $[H^+]$, set 10 to the power of the –pH. Same idea for pOH

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

To convert FROM pH to $[H^+]$ with any calculator

Example: **pH = 9.510;** to solve, find **$[H_3O^+]=10^{-9.51}$**

Push “10^x” then (–) 9.510 then enter. This gives **3.09×10^{-10}**.

Find [H_3O^+] and [OH^-]

[H_3O^+] = 10^{-pH} **[OH^-] = 10^{-pOH}**

Practice:

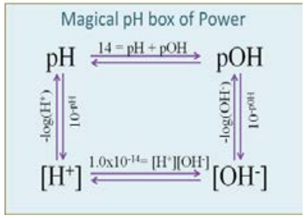
If the pH of a solution is 4.92, what is the [H_3O^+]?

1.2×10^{-5}M
______

If the pOH of a solution is 9.29, what is the [OH^-]?

5.1×10^{-10}M
______

Magical Box of pH Power



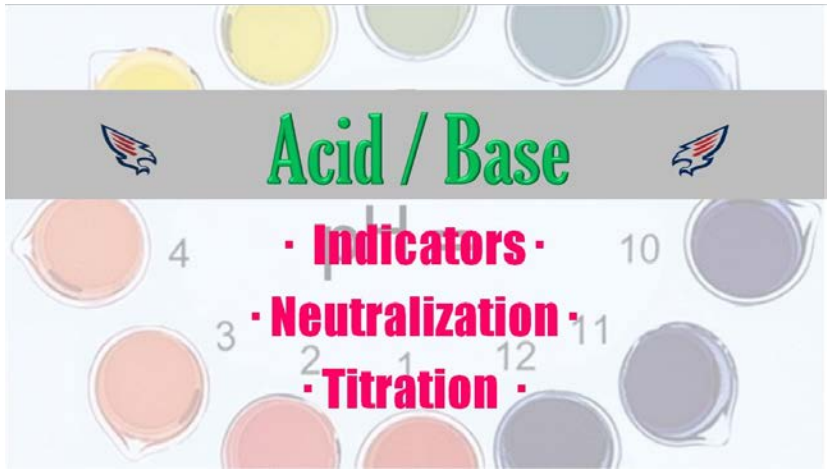
Practice

pH = -log[H+] pOH = -log[OH+]

[H_3O^+] = 10^{-pH} [OH^-] = 10^{-pOH}

[H^+][OH^-] = $1x10^{14}$ pH + pOH = 14

pH	pOH	[H_3O+]	[OH–]	Acid/Base Neutral
4.9	•	•	•	•
•	6.8	•	•	•
•	•	1.39×10^{-5}	•	•
•	•	•	9.85×10^{-11}	•



Indicators

- Indicators are substances that change **color** in solutions of different pH.
- Indicators are usually **weak** acids.
- They are one color in the acid form and a different color in the base form.
- $H\mathit{In} \leftrightharpoons H^+ + \mathit{In}^-$
- Indicators help determine approximate **pH**
- Limitations
  - > color is subjective
  - > ranges are narrow - sigfigs?
  - > colors can fade

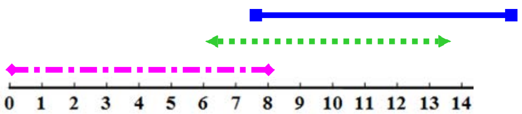
Indicators

Indicator	Color in acid	Transition Color	Color in base	Transition pH range
Bromocresol green	Yellow/green <3.8	None 3.8-5.4	Blue >5.4	3.8-5.4
Phenolphthalein	Colorless <8.2	Pink 8.2-10.2	Red >10.2	8.2-10.2
Bromothymol blue (BTB)	Yellow <6.0	Green 6.0-7.6	Blue >7.6	6.0-7.6
Methyl orange	Red <3.1	Orange 3.1-4.4	Yellow >4.4	3.1-4.4
Methyl red	Red <4.4	Orange 4.4-6.2	Yellow >6.2	4.4-6.2
Phenol red	Yellow <6.8	Orange 6.8-8.4	Red >8.4	6.8-8.4
Litmus	Blue turns red	n/a	Red turns blue	n/a
Universal (BTB + Phenolphthalein+ Methyl Red)	2-red, 3-red/orange 4-orange, 5-yellow/ orange 6-yellow	7-green	8-blue/green 9-blue.gray 10-violet	all

Indicators

Ex: Identify the pH of the substances that turn the following colors in the presence of the listed indicators. Hint: Use a number line!

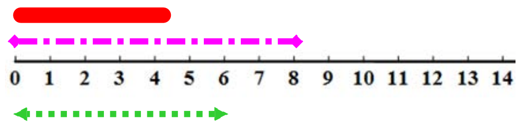
Indicator	Substance A
Phenolphthalein	Colorless
Methyl Red	Yellow
Bromothymol blue	Blue
pH	



Indicators

Ex: Identify the pH of the substances that turn the following colors in the presence of the listed indicators. Hint: Use a number line!

Indicator	Substance B
Phenolphthalein	Colorless
Methyl Red	Red
Bromothymol blue	Yellow
pH	



Neutralization

When acids are mixed with bases in equal quantities of hydronium and hydroxide ions, neutralization takes place.

ACID + BASE → WATER + a SALT

2 HBr + Ca(OH)2 → $\mathbf{H_2O + CaBr_2}$

Neutralization

Practice: Write the balanced chemical equation for these reactions

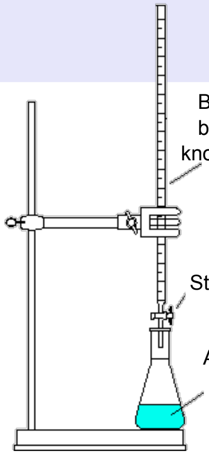
1. Nitric acid (HNO_3) and potassium hydroxide (KOH)

$1\ HNO_3 + 1\ KOH \rightarrow 1\ H_2O + 1\ KNO_3$

2. Sulfuric acid and magnesium hydroxide

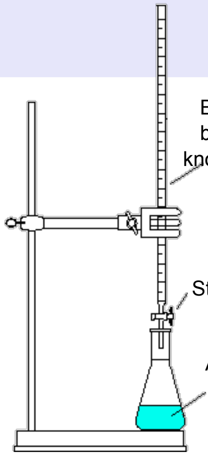
Titration

a process in which a **neutralization** reaction is used to determine the **molarity** of an unknown solution.



Titration

A known solution, the **titrant,** is dripped carefully into an unknown solution, the **analyte,** containing an indicator. When the new solution is at equivalent concentrations of [H+] and [OH-], the analyte concentration can be calculated.

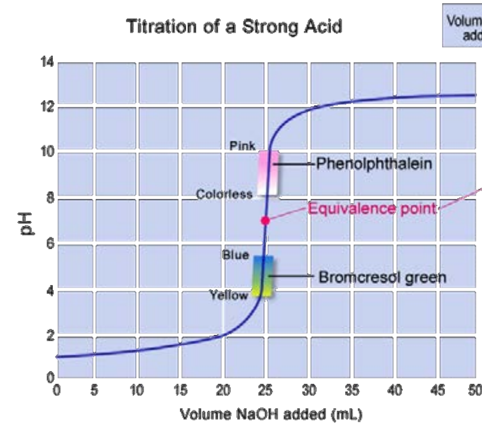


Titration

- **<u>Equivalence point</u>** – point of **neutrality** in a titration. For an acid/base titration: This is when #moles of [H+] = #moles of [OH-].
- **<u>Endpoint</u>** – point at which an indicator used in a titration changes color
- Choose an indicator that will give color change when you reach equivalence, so…
  > You want the endpoint to be as close to equivalence as possible!

Titration

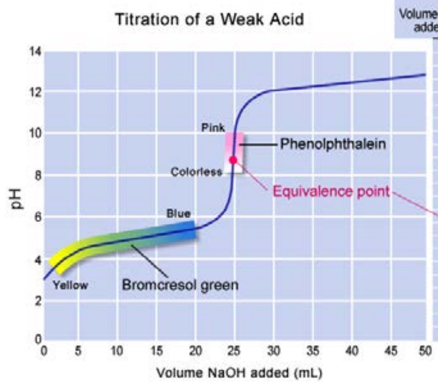
strong acid titrated with strong base



Volume NaOH added (mL)	pH
0	1.00
5	1.18
10	1.37
15	1.60
20	1.95
21	2.06
22	2.20
23	2.38
24	2.69
25	7.00
26	11.29
27	11.59
28	11.75
29	11.87
30	11.96
35	12.22
40	12.36
45	12.46
50	12.52

Titration

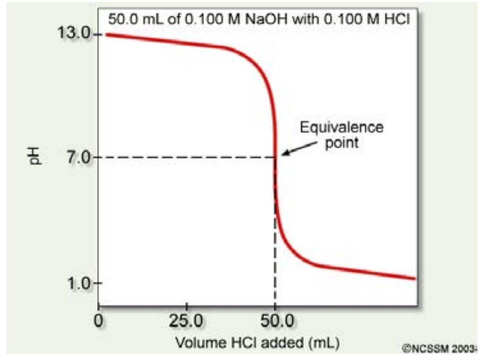
weak acid titrated with strong base



Volume NaOH added (mL)	pH
0	2.92
1	3.47
2	3.79
3	3.98
4	4.13
5	4.25
10	4.67
15	5.03
20	5.45
21	5.57
22	5.72
23	5.91
24	6.23
25	8.78
26	11.29
27	11.59
28	11.75
29	11.87
30	11.96
35	12.22
40	12.36
45	12.46
50	12.52

Titration

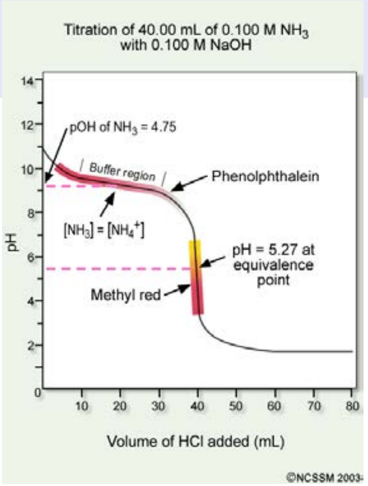
strong base titrated with strong acid



Titration

weak base titrated

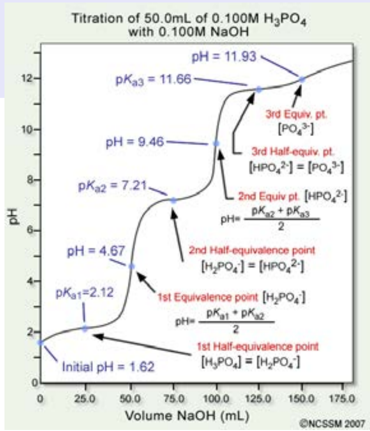
with strong acid



Titration

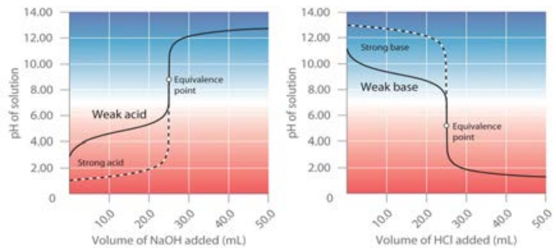
polyprotic weak acid

titrated with strong base



Titration

Compare Titration curves



(a) Weak acid titrated with strong base

(b) Weak base titrated with strong acid

Titration

To solve a titration (molarity of the unknown analyte):

1. you must work from a **BALANCED** chemical equation
   > *the mole ratio of acid to base is critical*
2. Write known quantities below reactants
3. Use stoich, starting with the **volume of titrant (in L),** convert to moles of titrant using the known molarity.
4. Then convert moles of titrant to moles of analyte using the mole ratio.
5. Solve for concentration of the analyte by dividing by volume of analyte (in L) used in the titration.

OR

Titration

Example

46.4mL of unknown molarity HCl are added to 25.0 mL of 1.00 M KOH to reach equivalence in a titration. What is the molarity of the acid?

Pull

Titration

Practice

1) What is the molarity of nitric acid (HNO_3) if 15.0 mL of the solution is completely neutralized by 38.5 mL of 0.150 M NaOH?

Pull

What is the pH of the titrant?

What type of salt is formed?

What is the estimated pH of the salt solution formed?

Titration

Practice

2) A 25.0 mL solution of sulfurous acid (H_2SO_3) of unknown molarity is completely neutralized by 18 mL of 1.0 M NaOH.

Pull

What is the pH of the titrant?

What type of salt is formed?

What is the estimated pH of the salt solution formed?

End

Working with weak acids & bases

Weak acids dissociate incompletely. Equilibrium constants for the dissociation of weak acids are called K_a values (K_b for weak bases). This is the degree to which those acids IONIZE. Strong acids do not have a K_a value because they ionize completely!

For a generic weak acid:

$HA \rightleftarrows H^+ + A^-$

A means anion!

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

A strong acid would have zero [HA], which messes up the math a tad.

example

$HCN \rightleftarrows H^+ + CN^-$

$$K_a = \frac{[H^+][CN^-]}{[HCN]}$$

Working with weak acids & bases

**Each H in a diprotic or triprotic acid has a separate K_a value. The same is true for bases (dibasic, tribasic).*

Example: K_{a1}: $H_2CO_3 \rightleftarrows H^+ + HCO_3^-$

K_{a2}: $HCO_3^- \rightleftarrows H^+ + CO_3^{-2}$

Equilibrium constants for weak bases are called ***K_b values.***

$NH_{3(aq)} + H_2O_{(l)} \rightleftarrows NH_4^+{}_{(aq)} + OH^-{}_{(aq)}$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

The larger the K_b value, the stronger the base

Math with weak acids & bases

1. Write the balanced chemical equation
2. Set up a reaction diagram (RICE diagram)
3. Set up K_a or K_b expression
4. Substitute values into K_a or K_b expression
5. Solve K_a or K_b expression for X. Use the **5% rule**
6. Calculate pH from H^+ concentration.

RICE diagram!

Reaction	$HA \rightleftarrows$	H^+ +	A^-
Initial			
Change			
Equilibrium			

Math with weak acids & bases

Ex: Calculate the pH of a 0.10 M solution of acetic acid. The K_a for acetic acid is 1.8×10^{-5}.

RXN $HC_2H_3O_2 \rightleftarrows H^+ + C_2H_3O_2^-$

RICE

Reaction	$HC_2H_3O_2$ ⇄	H^+ +	$C_2H_3O_2^-$
Initial	0.10M	0	0
Change	-x	+x	+x
Equilibrium	0.10-x M	x M	x M

Ka

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

$$K_a = \frac{[x][x]}{[0.10-x]} \rightarrow K_a = \frac{[x][x]}{[0.10]}$$

5% rule – x value far smaller than 5% of 0.10 so we can assume [0.10-x]≈[0.10]

Math with weak acids & bases

Ex: Calculate the pH of a 0.10 M solution of acetic acid. The K_a for acetic acid is 1.8×10^{-5}.

$$1.8 \, x \, 10^{-5} = \frac{x^2}{0.10} \rightarrow 1.8 \, x \, 10^{-6} = x^2$$

$$\sqrt{1.8 \times 10^{-6}} = x$$

$$x = 0.001342 = [H^+]$$

$$-\log(0.001342) = pH$$

$$pH = 2.87$$

Math with weak acids & bases

Practice

1) Calculate the pH of a 0.25 M solution of HCN. (K_a is 6.2×10^{-10})

RXN $HCN \rightleftarrows CN^- + H^+$

Pull

K_a $K_a = \frac{[CN^-][H^+]}{[HCN]}$

RICE

Reaction	HCN ⇄	H_3O^+ +	CN^-
Initial	0.25M	0	0
Change	-x	+x	+x
Equilibrium	0.25-x	x	x

Percent Dissociation

We can also calculate the percent dissociation of the acid...

$$\% \, dissociation = \frac{amount \; dissociated \; (\frac{mol}{L})}{initial \; concentration \; (\frac{mol}{L})} \, x \, 100 \quad OR \ldots \frac{[H_3O^+]}{[HA]} \, x \, 100$$

For practice 1) $\% \, dissociation = \frac{1.2 \, x \, 10^{-5}}{0.25} \, x \, 100 =$ 0.0048%

Weak base with % dissociation!

1) Find the [OH-] of a 1.0 M solution of methylamine. (K_b = 4.38 x 10^{-4}) Then calculate the percent dissociation of the base. Methylamine is CH_3NH_2.

K_b $K_b = \frac{[CH_3NH_3^+][OH^-]}{[CH_3NH_2]}$

Pull

RXN $CH_3NH_2 \leftrightarrows CH_3NH_3^+ + OH^-$

RICE

Reaction	CH_3NH_2 ⇄	OH^- +	$CH_3NH_3^+$
Initial	1.0M	0	0
Change	-x	+x	+x
Equilibrium	1.0-x	x	x

End

Titration

Compare Titration curves