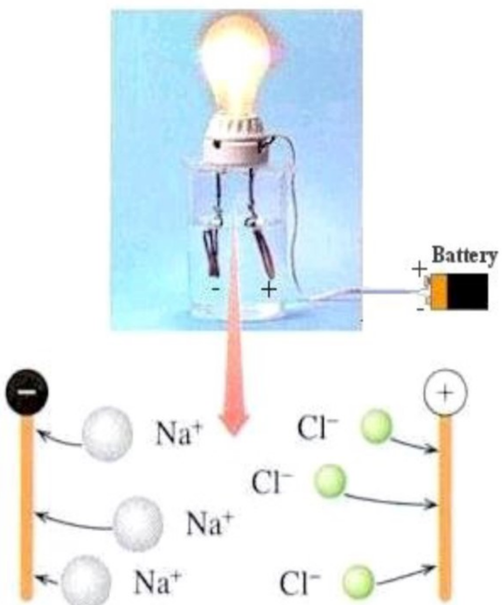
Unit 9: Acids, Bases, & Salts

Unit Vocabulary:

Amphoteric	hydronium ion
Arrhenius acid	hydroxide ion
Arrhenius base	indicator (acid/base)
Bronsted-Lowry acid	neutralization
Bronsted-Lowry base	pH scale
Electrolyte	titration



Unit Objectives:

- Compare and contrast properties of acids, bases, and salts
- Compare the Arrhenius and Bronsted-Lowry theories of acids and bases
- Explain and give examples of neutralization reactions
- Using the titration equation, determine the molarity of an unknown solution
- Understand how pH works
- Using Table M, determine the pH of a given solution

Characteristic Properties of **Acids:**

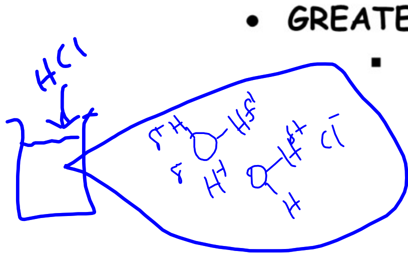
Table K
Common Acids

Formula	Name
$HCl(aq)$	hydrochloric acid
$HNO_3(aq)$	nitric 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)

- **DILUTE** acids have a **SOUR** taste
  Ex: citric acid (fruit), acetic acid (vinegar), carbonic acid (soda), boric acid used as eye-washing solution

- **CONCENTRATED** acids **BURN** skin & **EAT HOLES** in clothing

- Aqueous solutions of acids are **ELECTROLYTES** (substances that conduct electric current when dissolved in water)
  - **GREATER** concentration of **IONS** = **MORE CONDUCTIVE**
    - **STRONG** acids = **GOOD** conductors; 100% dissociation



1. HCl	4. H_2SO_4
2. HBr	5. HNO_3
3. HI	6. $HClO_4$

    - **WEAK** acids = **POOR** conductors; less than 100% dissociation
      - ✓ All other acids

❖ Acids react with **BASES** to form **NEUTRAL** solutions
- **DOUBLE REPLACEMENT** reaction
- Called a **NEUTRALIZATION** reaction
- ***Acid + Base → Salt + Water*** HOH

Ex: _1_ HCl(aq) + _1_ NaOH(aq) → _1_ NaCl(aq) (salt) + _1_ $H_2O_{(l)}$ (water)

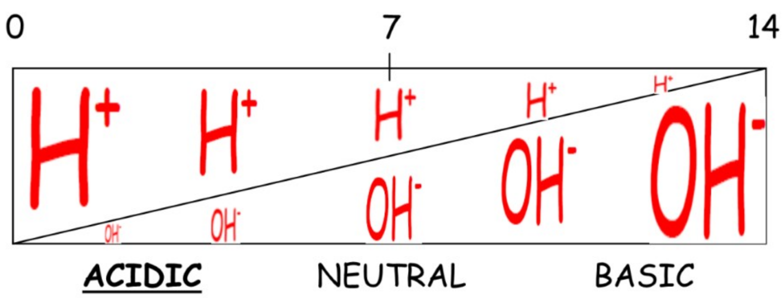
❖ Acids react with certain **METALS** to produce **HYDROGEN GAS**
- ***SINGLE REPLACEMENT*** reaction (Table J)
- ***Acid + Metal → Anion + Hydrogen Gas***

Ex: _2_ HCl(aq) + _2_ Na(s) → _2_ NaCl(aq) + ___ $H_{2(g)}$

Table M

❖ Acids cause acid-base **INDICATORS** to **CHANGE COLOR**
- Ex: **LITMUS** paper turns **RED**, **PHENOLPHTHALEIN** turns from **PINK TO COLORLESS**

❖ Acids have **pH VALUES < 7** (fall on **LOWER** end of **pH SCALE**)



❖ General formula = **HA** or **HX** (Where X = **ANION** such as **Cl**$^-$)

Characteristic Properties of **Bases:**

Table L
Common Bases

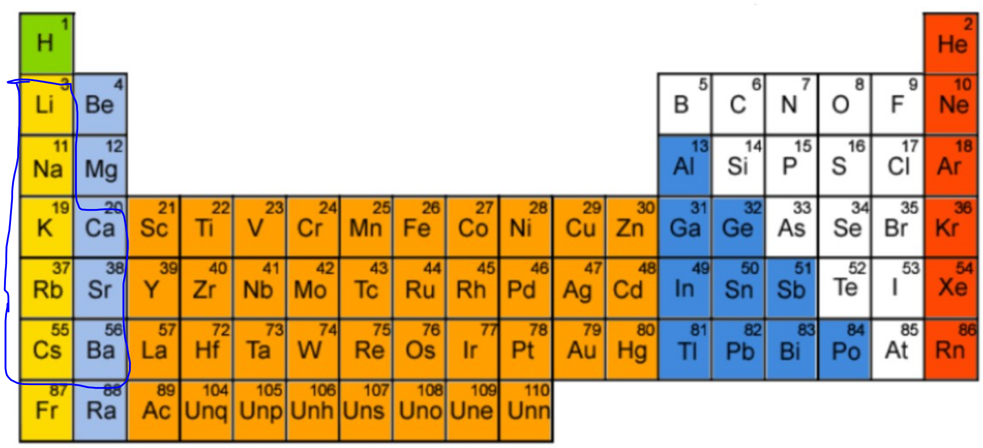
Formula	Name
$NaOH(aq)$	sodium hydroxide
$KOH(aq)$	potassium hydroxide
$Ca(OH)_2(aq)$	calcium hydroxide
$NH_3(aq)$	aqueous ammonia

- **DILUTE** bases have a **BITTER** taste
  Ex: antacids, soaps, ammonia-based cleaning products

- Bases have a **SLIPPERY** or **SOAPY** feel

- **CONCENTRATED** bases **BURN** skin & **EAT HOLES** in clothing

- Aqueous solutions of bases are **ELECTROLYTES** (substances that conduct electric current when dissolved in water)
  - **GREATER** concentration of **IONS** = **MORE CONDUCTIVE**
    - **STRONG** bases = **GOOD** conductors



- **WEAK** bases = **POOR** conductors

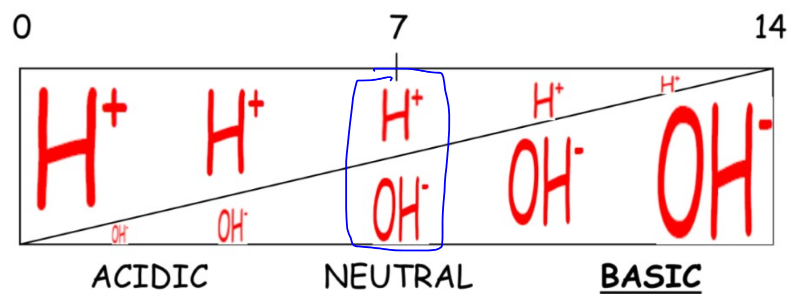
- **BASES** react with **ACIDS** to form **NEUTRAL** solutions
  - **DOUBLE REPLACEMENT** reaction
  - Called a **NEUTRALIZATION** reaction
  - ***Acid + Base → Salt + Water***

Ex: 2 $HBr(aq)$ + 1 $Ca(OH)_2(aq)$ → 1 $CaBr_2(aq)$ + 2 H_2O

salt water

- Bases cause acid-base **INDICATORS** to **CHANGE COLOR**
  - Ex: **LITMUS** paper turns **BLUE**, **PHENOLPHTHALEIN** turns from **COLORLESS TO PINK**

- Bases have **pH VALUES > 7** (fall on **HIGHER** end of **pH SCALE**)



$[H^+] = [OH^-]$

- General formula = **XOH** (Where X = **CATION** such as **Na^+**)

Characteristic Properties of **Salts:**

- Defined as **NEUTRAL IONIC SUBSTANCES** that have **POSITIVE IONS OTHER THAN HYDROGEN** and **NEGATIVE IONS OTHER THAN HYDROXIDE**; composed of a positively charged metal or polyatomic ion AND a negatively charged nonmetal or polyatomic ion (with the exception of $-OH^-$ which would then make it a base)

- Examples of salts → **LiBr, KI, $CaCl_2$, $NaNO_3$**

- Salts are **FORMED FROM NEUTRALIZATION REACTIONS** and are **NEUTRAL**

acid + base → Salt + water

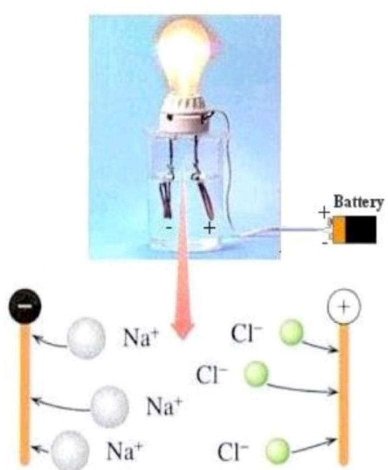
pH=7

- Exception to the rule: **NH_4Cl** is the salt of a weak acid

$NH_4Cl \rightarrow NH_4^+ + Cl^-$
NH_4^+ is a weak acid, which means it produces H^+ in solution

pH<7

Electrolyte = A substance that **DISSOLVES IN WATER** to form **MOBILE IONS** and therefore **CONDUCTS ELECTRICITY**



**ACIDS,
BASES,
& SALTS
ARE ALL ELECTROLYTES
(in SOLUTION)**

Rules for Naming Acids:

Binary Acids → acids that **START WITH H** and are attached to a **NONMETAL**; **2 TOTAL ELEMENTS** in acid's formula

1st word in name → Begin with ***hydro-*** & follow it up immediately with the name of the other element while replacing the ending with ***-ic***

2nd word in name → add the word **acid** as the second (last) word in the name

Ex: HCl — Hydrochloric acid

Ex: HI — hydroiodic acid

Ex: HBr — hydrobromic acid

Ternary Acids → acids that **START WITH H** & are attached to a **POLYATOMIC ION**; **3 TOTAL ELEMENTS** in the compound formula—**USE TABLE E!**

1st word in name: There is **NO *"HYDRO"*** in front! Same rules as the binary acids, except the ***-ate/-ite*** are replaced by ***-ic*** **and *-ous*** (respectively)

So: If the polyatomic ion ends in -ATE it gets replaced by -IC
(I ate in the café and went ic!)
If the polyatomic ion ends in -ITE it gets replaced by -OUS

2nd word in name: Add the word **acid** as the second word in the name

Ex: HNO_3 — Nitric acid

Ex: HNO_2 — Nitrous acid

Ex: $HC_2H_3O_2$ — Acetic acid

There are a couple compounds that have **SLIGHT EXCEPTIONS to the rules outlined above. Sometimes you must just go with what "rolls off the tongue" more smoothly!

Ex: H_2SO_3 — Sulfurous acid

Ex: H_2SO_4 — Sulfuric acid

Rules for Naming Bases:

*much easier than Naming Acids!

1) Name the 1[ST] **ION** in the formula **"AS IS."** This will always be the + ion and will either be a metal (ex: Na) or a polyatomic ion (ex: NH_4)
2) Name the -**OH ON THE END AS HYDROXIDE** (separate second word in the name) – Table E says that OH is called "hydroxide"
3) NOTE: it does not matter how many you have of each when it comes to the positive or negative ion (see second example below). Don't ever use prefixes!

Ex: LiOH — Lithium hydroxide

Ex: $Mg(OH)_2$ — Magnesium hydroxide

Ex: NH_4OH — Ammonium hydroxide

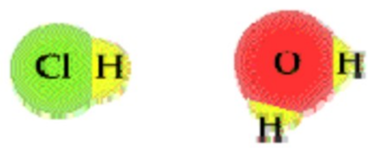
Practice Naming:

Acid Name	Acid Formula	Anion name/Symbol
Hydrobromic acid	HBr	Bromide / Br^-
Hydrosulfuric acid	H_2S	Sulfide / S^{2-}
Carbonic acid	$H_2^+ CO_3^{2-}$	Carbonate / CO_3^{2-}
Chlorous acid	$HClO_2$	Chlorite / ClO_2^-

Base Name	Formula
Calcium hydroxide	$Ca(OH)_2$
Sodium hydroxide	NaOH

Acid/Base Theories

An **Arrhenius Acid** is defined as a substance whose water (aqueous) solution contains or yields **HYDRONIUM IONS (H+ IONS)** as the **ONLY POSITIVE ION** in solution H_3O^+



Examples:

$HCl \rightarrow H^+_{(aq)} + Cl^-_{(aq)}$

$H_2SO_4 \rightarrow H^+ + H^+ + SO_4^{-2}$

- **STRONG** acids → **DISSOCIATE 100% in H_2O**

GENERAL RULE FOR ACIDS: **HYDROGEN (H)** is the **FIRST ION** seen in the formula (H is always the **POSITIVE ION**)

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** **NOT ALL SUBSTANCES THAT CONTAIN HYDROGEN ARE ACIDS.** Below is a list of hydrogen compounds that do not dissociate to yield H^+ ions.

Non-Examples of Acids:
H_2O = water (neutral & amphoteric)
CH_4 = methane (natural gas)
$C_6H_{12}O_6$ = glucose (sugar)
H_2O_2 = hydrogen peroxide
NH_3 = ammonia (a weak BASE!!!)

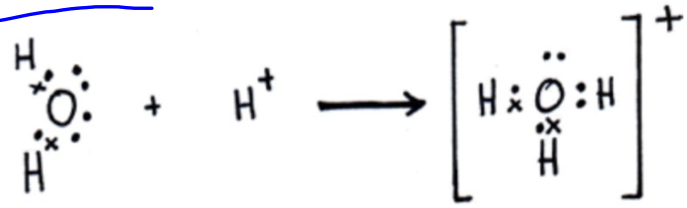
THE HYDRONIUM ION:

Let's recall what happens when a hydrogen atom becomes an ion:

(Lose e^-)

$$H \rightarrow H^+ + e^-$$

Thus, a **POSITIVE HYDROGEN ION** is essentially a **PROTON**. When in water, this **H^+ ION** is naturally **ATTRACTED** to the unshared electrons and slight negative charge of the oxygen in the water. It is believed that the hydrogen ion cannot exist as an isolated particle so what forms is called a **HYDRONIUM ION (H_3O^+)**.



According to the Arrhenius theory, the properties of acids are functions of these hydronium (hydrogen) ions. So, because of this, we say that **H^+ = H_3O^+**

DIFFERENT TYPES OF ARRHENIUS ACIDS:

Monoprotic acids (1 H^+) → acids that ionize in **1 STEP;** acids that produce **1 H^+ ION** in solution

Ex: $HCl \rightarrow$ H^+ + Cl^- (100%) strong

Diprotic acids (2 H^+) → acids that ionize in **2 STEPS**; acids that produce **2 H^+ IONS** in solution

Strong acid → Ex: $H_2SO_4 \rightarrow$ H^+ + HSO_4^- (100%)

weak acid → $HSO_4^- \rightarrow$ H^+ + SO_4^{2-} (<<<< 100%)

Triprotic acids (3 H^+) → acids that ionize in **3 STEPS**; acids that produce **3 H^+ IONS** in solution

Mildly weak acid → Ex: $H_3PO_4 \rightarrow$ H^+ + $H_2PO_4^-$ (<< 100%)

$H_2PO_4^- \rightarrow$ H^+ + HPO_4^{2-} (<<< 100%)

$HPO_4^{2-} \rightarrow$ H^+ + PO_4^{3-} (<<<<< 100%)

An **Arrhenius Base** is defined as a substance whose water (aqueous) solution contains or yields/produces **OH^- (HYDROXIDE IONS)** as the **ONLY NEGATIVE ION** when dissolved in water. Metal-OH = Base always

Examples: $NaOH(aq)$, $Ca(OH)_2(aq)$

GENERAL RULE FOR BASES: contains **–OH at the end** of the formula

One Exception to Rule: **NH_3 = Ammonia

Monohydroxy Bases → produce 1 –OH in aqueous sol$^{\underline{n}}$ (single step).

Example: KOH → K^+ + OH^-

Dihydroxy Bases → produce 2 –OH in aqueous sol$^{\underline{n}}$ (single step).

Example: $Ca(OH)_2$ → Ca^+ + $2OH^-$

Trihydroxy Bases → produce 3 –OH in aqueous sol$^{\underline{n}}$ (single step).

Example: $Al(OH)_3$ → Al^{3+} + $3OH^-$

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Exceptions

Not all compounds that contain –OH are bases. For example, **ALCOHOLS and **ORGANIC ACIDS** are **NOT BASES**. Below is a list of compounds that contain –OH but do not dissociate to yield OH^- ions.

Non-Examples of Bases:

HOH = water (neutral & amphoteric)

CH_3OH = methanol (an alcohol)

CH_3COOH = methanoic acid (an organic acid)

Alternate Acid-Base Theory (AKA Bronsted-Lowry Theory)

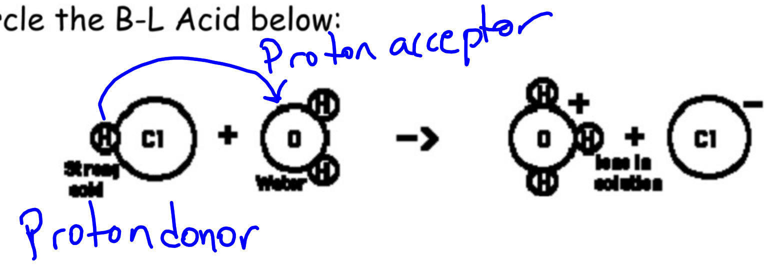
This theory explains the behavior of **WEAK ACIDS & WEAK BASES**

Definitions of the Theory:

BRONSTED-LOWRY ACIDS are the **PROTON DONORS**

- **Bronsted-Lowry Acid donates H^+ to Bronsted-Lowry Base**

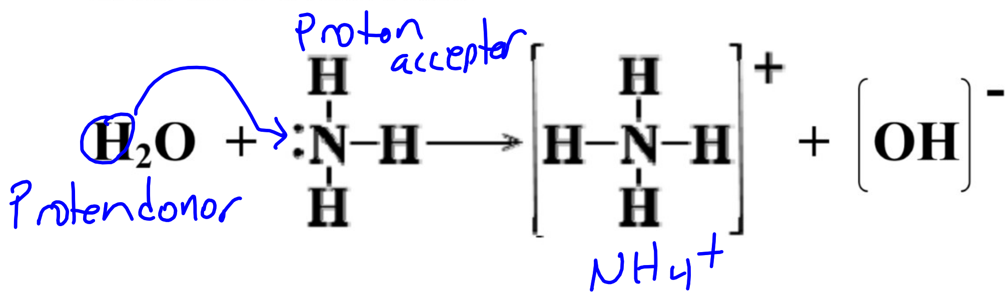
  Circle the B-L Acid below:



BRONSTED-LOWRY BASES are the **PROTON ACCEPTORS**

- **Bronsted-Lowry Base accepts H^+ from Bronsted-Lowry Acid**

  Circle the B-L Base below:



Bronsted-Lowry Conjugate Pairs:

We will use **TWO BRACKETS** connecting one side of the reaction to the other to represent the **ACID-BASE CONJUGATE PAIRS** (each member within the pair **DIFFERS** from the other **BY MERELY ONE HYDROGEN**)

A (HCl) — CB (Cl^-)

$$HCl_{(aq)} + H_2O_{(l)} \leftrightarrows H_3O^+_{(aq)} + Cl^-_{(aq)}$$

B (H_2O) — CA (H_3O^+)

KEY

A = Acid
B = Base
CA = Conjugate Acid
CB = Conjugate Base

Notice: The acid always has one more H than the base!

Use brackets and the key from the example on the previous page to indicate the conjugate pairs in the equations below.

Practice Problem 1:

$$NH_3 + HOH \rightarrow NH_4^+ + OH^-$$

(B, A, CA, CB)

Practice Problem 2:

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

(A, B, CB, CA)

Practice Problem 3:

$$HI + HOH \rightarrow I^- + H_3O^+$$

(A, B, CB, CA)

Practice Problem 4:

$$HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$$

(A, B, CA, CB)

<u>**Amphoteric/Amphiprotic**</u> = A substance that can **ACT LIKE AN ACID <u>OR</u> A BASE** (can behave as a proton donor or proton acceptor)

Example: **WATER** is amphoteric (see below)

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

(A, B, CA, CB)

*Amphoteric substances must have at least one **HYDROGEN** in their formula

Ex: Which of the following substances is (most) amphoteric?

a.) H_2SO_3

b.) ~~HBr~~

c.) HSO_3^-

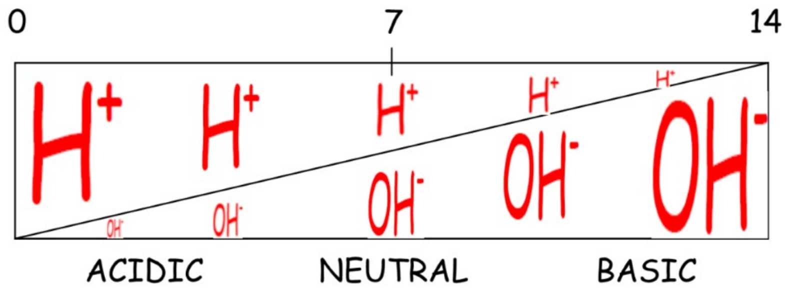
d.) ~~Br^-~~

The pH Scale - The Power of Hydrogen:

pH = direct measurement of H^+ **ION CONCENTRATION** in a solution

The **pH scale** is designed to measure **HOW ACIDIC** or **HOW BASIC** an aqueous solution is. The concentration of hydronium and hydroxide ions in a solution will determine whether a solution is acidic or basic.

The pH scale ranges from **0 TO 14** (**ACIDIC** → neutral → **BASIC**)



In an acidic substance **$[H^+]$ > $[OH^-]$** (*brackets indicate concentration*)
In a basic substance **$[H^+]$ < $[OH^-]$**
In a neutral substance **$[H^+]$ = $[OH^-]$** Example: **H_2O**

The pH scale is **LOGARITHMIC** (based on exponents of the number **10**)

1. log (10) = 10^1
2. log (100) = 10^2
3. log (1) = 10^0
4. log (.01) = 10^{-2}

The pH of a solution is the negative log of the $[H^+]$:

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

1. pH of 1.0 M $HClO_4$ = **-log (1.0) =** 0
2. pH of .5 M H_2SO_4 = **-log (.5) =** .30
3. pH of .01 M HBr = **-log (.01) =** 2.0

4. pH of 12 M HCl = **-log (12) =** -1.1

For basic solutions, find the pOH using the negative log of the $[OH^-]$, then subtract that value from 14 to get the pH:

pOH = **-log[OH^-]**

pH = **14 - pOH**

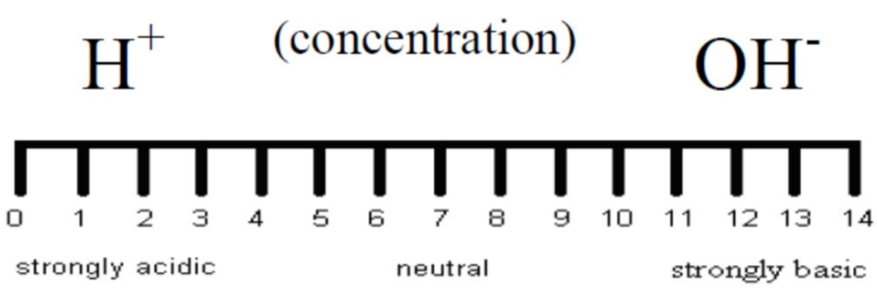
1. pH of 1.0 M NaOH = **14-(-log (1.0)) =** 14-0=14

2. pH of .5 M $Ca(OH)_2$ = **14-(-log (1.0)) =** 14-0=14

3. pH of .01 M LiOH = **14-(-log (0.01)) =** 14-2 =12

4. pH of 6.0 M KOH = **14-(-log (6.0)) =** 14 - 0.78 = 13.22

Each change of a **SINGLE pH UNIT** signifies **A TENFOLD CHANGE IN $[H^+]$ CONCENTRATION**



Ex: going from a ph of 4 to a pH of 5

Becoming more basic; $[OH^-]$ ↑ by 10X

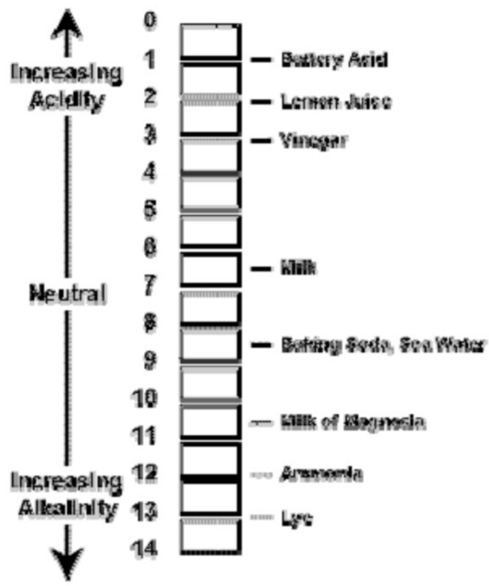
Becoming less acidic; $[H^+]$ ↓ by 10X

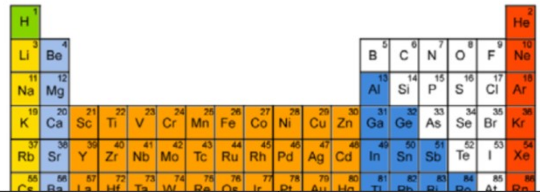
Ex: going from a pH of 13 to a pH of 10

Becoming more acidic; $[H^+]$ ↑ by 1000X

Becoming less basic; $[OH^-]$ ↓ by 1000X

*Where some common substances fall on the pH scale





Strong Acids	Strong Bases
HCl	LiOH
HBr	NaOH
HI	KOH
H_2SO_4	RbOH
HNO_3	CsOH
$HClO_4$	$Ba(OH)_2$
	$Sr(OH)_2$
	$Ca(OH)_2$

STRONG acids and bases ionize 100%:

Strong Acid (hydrochloric acid): $HCl_{(g)} \xrightarrow{H_2O} H^+_{(aq)} + Cl^-_{(aq)}$ (100% ionization)

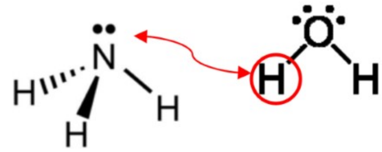
Strong Base (sodium hydroxide): $NaOH_{(s)} \xrightarrow{H_2O} Na^+_{(aq)} + OH^-_{(aq)}$ (100% ionization)

WEAK acids and bases have < 100% ionization (partial ionization)

Weak Acid (acetic acid): $HC_2H_3O_2 \xrightarrow{H_2O} H^+ + C_2H_3O_2^-$ (partial ionization)

Weak Base (ammonia): $NH_3 + H_2O \rightarrow NH_4^+ + OH^-$ (partial ionization)

Ammonia and the –OH in water battle for hydrogen. Roughly 1/10,000 times, ammonia wins, generating a SLIGHT excess of OH^- ions, making the solution BASIC.



Acid-Base Indicators (Table M):

Indicator = substance (weak acid) that **CHANGES COLOR** as a result of a **pH CHANGE**; indicators are chosen based on their **RANGE for COLOR CHANGE**

Phenolphthalein → **COLORLESS** up until a pH of 8, **LIGHT PINK** from 8 to 9, and **PINK** at a pH greater than 9

Ex: Methyl Orange → turns **RED** in a solution with a pH of less than 3.1, **ORANGE** between 3.1 and 4.4, and **YELLOW** in a solution with a pH greater than 4.4

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

*Litmus listed is liquid litmus (similar to paper litmus)

** Within the **"Approximate pH Range for Color Change"** a **MIXTURE OR BLENDING** of the two colors listed occurs. This range is therefore also called the **INTERMEDIATE COLOR REGION**.

1. bromthymol blue at a pH of 6.4 → green
2. bromcresol green at a pH of 5.0 → green
3. phenolphthalein at a pH of 9.2 → pink
4. methyl orange at a pH of 3.9 → orange

Using reference table M, complete the chart below.

Indicator	pH of sample	Color indicator will turn
methyl orange	6.0	yellow
bromothymol blue	2.0	yellow
phenolphthalein	7.0	colorless
Litmus	6.8	purple
bromocresol green	3.0	yellow
thymol blue	7.0	yellow
methyl orange	2.2	red
bromothymol blue	8.2	blue
phenolphthalein	10	pink (dark)
Litmus	3.2	red
bromocresol green	6.2	blue
thymol blue	10	blue

Neutralization Reactions:

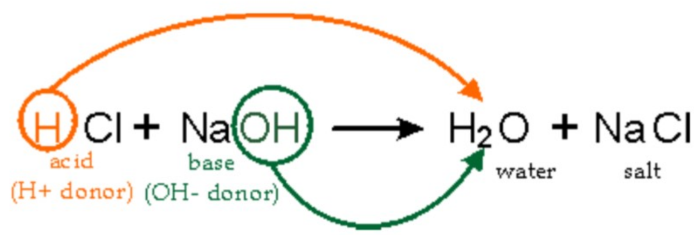
Neutral = neither acidic nor basic; $[H^+] = [OH^-]$

Neutralization occurs when an Arrhenius acid and an Arrhenius base react to form **WATER** and a **SALT**

Example: Antacid for upset stomach neutralizes the acid in stomach and makes a neutral salt to provide relief

General reaction for neutralization reactions:

ACID + BASE → salt + water



(1) The $\mathbf{H^+}$ from the acid and the $\mathbf{OH^-}$ from the base combine to form water.
(2) The **ANION** from the acid and the **CATION** from the base combine to produce a salt.

*Neutralization rxns are always **DOUBLE REPLACEMENT** reactions!
**Remember, just like any compound, the (+) and (-) charges must be balanced. Use the Criss Cross Rule to figure out the correct formula for the salt – water is always the formula H_2O!

Net ionic equation for neutralization reactions (after spectator ions are crossed out):

$$H^+(aq) + OH^-(aq) \rightarrow H_2O_{(l)}$$

Complete the following reactions. Make sure they are balanced!

1. HBr + KOH → KBr + H_2O

2. NaOH + $HC_2H_3O_2$ → $NaC_2H_3O_2$ + H_2O

3. 3KOH + H_3PO_4 → $3KPO_4$ + $3H_2O$

4. $2HNO_3$ + 2$Ca(OH)_2$ → 2$CaNO_3$ + $3H_2O$

Do we always get a completely neutral solution?

Combination	Product (Acidic/Basic/Nuetral)	Endpoint pH (<7, 7, >7)
Strong acid + Strong Base	Neutral	7
Strong acid + Weak Base	Acidic	<7
Weak acid + Strong Base	Basic	>7

(Acid-Base) Titration:

Titrations are used to **CALCULATE THE CONCENTRATION (MOLARITY) OF AN UNKNOWN SOLUTION**

- Acid of unkown molarity is reacted with a carefully measured amount of a base of known molarity to the point of **NEUTRALIZATION** (or vice versa)

In all neutralization reactions there must be a **1:1 RATIO** between the **MOLES OF H^+ IONS** and the **MOLES OF OH^- IONS**

So, in a titration, when:

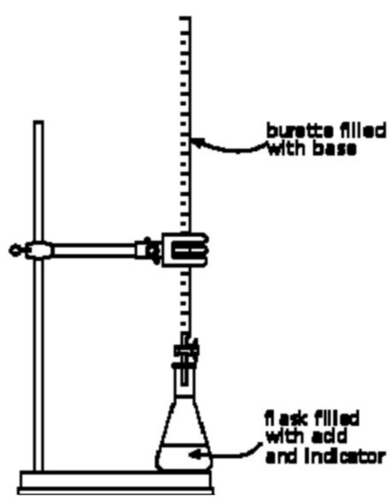
$$[H^+] = [OH^-]$$

or

MOLES OF H^+ IONS = MOLES OF OH^- IONS

that means you've reached the **EQUIVALENCE POINT** of the reaction; this is when the titration is complete. The **ENDPOINT** of a titration is the pH at which the solution changes color permanently.

What does a Titration look like?



Titration formula (Table T): $M_A V_A = M_B V_B$

Where: M_A = molarity of acid (H^+)
V_A = volume of acid
M_B = molarity of base (OH^-)
V_B = volume of base

- Use this formula when you are dealing with a titration or neutralization word problem
- Make sure that all units are in agreement when plugging into formula (so they cancel out and you get the right answer!)

Sample Problem 1: What is the concentration of a solution of HI if 0.3 L is neutralized by 0.6 L of 0.2 M solution of KOH?

$M_A = ?$
$V_A = 0.3$
$M_B = 0.2$
$V_B = 0.6$

$$\frac{M_A V_A}{V_A} = \frac{M_B V_B}{V_A} = \frac{(0.2 \times 0.6)}{0.3} = \boxed{0.4\,M}$$

Sample Problem 2: What is the concentration of a hydrochloric acid solution if 50.0 mL of a 0.250 M KOH solution are needed to neutralize 20.0 mL of the HCl solution of unknown concentration?

$M_A = ?$
$V_A = 20$
$M_B = 0.250$
$V_B = 50$

$$M_A V_A = M_B V_B$$
$$M_A = \frac{M_B V_B}{V_A} = \frac{(0.250 \times 50)}{20} = \boxed{0.625\,M}$$

Sample Problem 3: A particular acid has an H+ concentration of 0.1 M and a volume of 100 mL. What volume of a base with a 0.5 M [OH-] will be required to neutralize the reaction?

$M_A = 0.1$
$V_A = 100$
$M_B = 0.5$
$V_B = ?$

$$M_A V_A = M_B V_B$$
$$0.1 \times 100 = 0.5\, V_B$$
$$V_B = \boxed{20\text{ mL}}$$

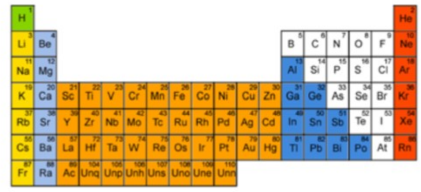
****Sample Problem 4:** You have 50 mL of 1.0 M $H_2SO_{4(aq)}$. What volume of 0.5 M NaOH would be required to neutralize the acid? Remember → Diprotic Acids yield 2 H^+ ions in solution!

$V_A = 50\text{ ml}$
$M_A = 2(1.0)$
$M_B = 0.5$
$V_B = ?$

$$M_A V_A = M_B V_B$$
$$2 \times 1.0 \times 50 = 0.5\, V_B$$
$$V_B = \boxed{200\text{ mL}}$$

Reactions of Metals with Acids:

According to **Table J** in your Reference Tables, any **METAL LOCATED ABOVE H_2 WILL REACT WITH AN ACID** to produce $H_2(g)$ + **SALT**



Standard Hydrogen Electrode:

Example: Which metal, Mg or Cu will react with HCl? Mg

Table J
Activity Series**

Metals (Most → Least)	Nonmetals (Most → Least)
Li	F_2
Rb	Cl_2
K	Br_2
Cs	I_2
Ba	
Sr	
Ca	
Na	
Mg	
Al	
Ti	
Mn	
Zn	
Cr	
Fe	
Co	
Ni	
Sn	
Pb	
**H_2	
Cu	
Ag	
Au	

**Activity Series based on hydrogen standard
Note: H_2 is *not* a metal

General Rxn:

Metal + Acid → H_2 + salt

Ex: Zn + 2HCl → H_2 + $ZnCl_2$ (salt)

Of the four types of reactions that we have learned, this is a **SINGLE REPLACEMENT** reaction; notice how **H_2 GAS** is produced!

Ex: Cu, Ag, & Au do not react with acids because they are located below H_2 on Table J (notice these are metals used for JEWELRY!)

Reactions of Metals with Water:

GROUP I METALS: $Li + H_2O \rightarrow LiOH + H_2$

- All Group I metals react vigorously with "cold" water

GROUP II METALS:

- Berryllium has no reaction with "cold" water or steam, even at red heat
- Magnesium has only a slight reaction with "cold" water, and magnesium burns in steam to produce white magnesium oxide and hydrogen gas:

$$\overset{0}{Mg}(s) + \overset{+1\,-2}{H_2O}(g) \rightarrow \underline{\overset{+2\,-2}{MgO}}(s) + \underline{\overset{0}{H_2}}\ (g)$$

- Calcium, Strontium, and Barium all react with "cold" water with increasing vigor to produce metal hydroxide and hydrogen gas:

$$X(s) + 2H_2O(g) \rightarrow \underline{\overset{+2}{X}(OH)_2}\ (aq) + \underline{H_2}\ (g)$$

- Summary of the trend for Group II Metals: the Group II Metals become more reactive with water as you go **DOWN** the group

Regents Chemistry **Reactivity of Metals with Acid**

1. Which metal is more active than H_2?

   A) Au B) Cu C) Ag D) Pb

2. According to Reference Table *J*, which of these metals will react most readily with 1.0 M HCl to produce $H_2(g)$?

   A) Zn B) Ca C) Mg D) K

3. Under standard conditions, which metal will react with 0.1 M HCl to liberate hydrogen gas?

   A) Ag B) Cu C) Mg D) Au

4. Referring to Reference Table J, which reaction will not occur under standard conditions?

   A) $Ba(s) + 2\ HCl(aq) \rightarrow BaCl_2(aq) + H_2(g)$
   B) $Cu(s) + 2\ HCl(aq) \rightarrow CuCl_2(aq) + H_2(g)$
   C) $Sn(s) + 2\ HCl(aq) \rightarrow SnCl_2(ag) + H_2(g)$
   D) $Mg(s) + 2\ HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$

5. According to the Activity Series, which metal will react spontaneously with hydrochloric acid?

   A) Hg B) Ni C) Cu D) Ag

6. According to Reference Table J, which redox reaction occurs spontaneously?

   A) $Cu(s) + 2\ H^+ \rightarrow Cu^{2+} + H_2(g)$
   B) $Mg(s) + 2\ H^+ \rightarrow Mg^{2+} + H_2(g)$
   C) $2\ Ag(s) + 2\ H^+ \rightarrow 2\ Ag + H_2(g)$
   D) $2\ Ag(s) + 2\ H^+ \rightarrow 2\ Ag^{2+} + H_2(g)$

7. Based on Reference Table J, which metal will *not* react with 1 M HCl?

   A) Ni(s) B) Au(s)
   C) Sn(s) D) Zn(s)

8. According to Reference Table J, which metal will react spontaneously with H^+?

   A) Cu B) Ag C) Cr D) Au

9. According to Reference Table *J*, which metal will react spontaneously with hydrochloric acid?

   A) silver B) zinc
   C) copper D) gold

10. Based on Reference Table J, which reaction will take place spontaneously?

    A) $2\ Ag + 2\ H^+ \rightarrow 2\ Ag^+ + H_2$
    B) $Cu + 2\ H^+ \rightarrow Cu^{2+} + H_2$
    C) $2\ Au + 6\ H^+ \rightarrow 2\ Au^{3+} + 3\ H_2$
    D) $Pb + 2\ H^+ \rightarrow Pb^{2+} + H_2$