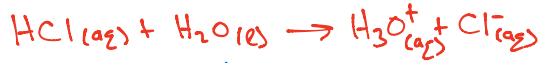


Acids and Bases Fall into 3 categories. Starting w/ most specific?

### Arrhenius

**Arrhenius Acid** - When added to  $H_2O$ , the  $[H_3O^+]$  increases  
 - Release  $H^+$  (or more accurately,  $H_3O^+$ )



**Arrhenius Base** - Release  $OH^-$  when added to  $H_2O$

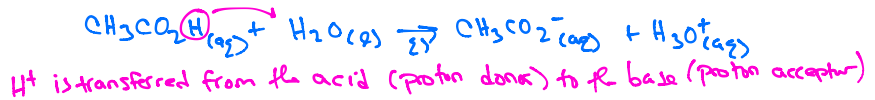


\*Note that including  $H_2O$  is not necessary for Arrhenius base. " $OH^-$ " is built into the formula &

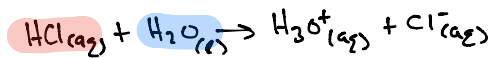
### Bronsted-Lowry

Bronsted-Lowry Acid = Proton Donor  
 $\swarrow$   
 $H^+$

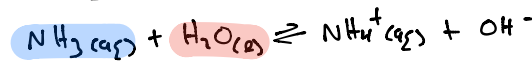
Bronsted-Lowry Base: Proton Acceptor



Identify the Bronsted-Lowry Acid + Base in each reaction:

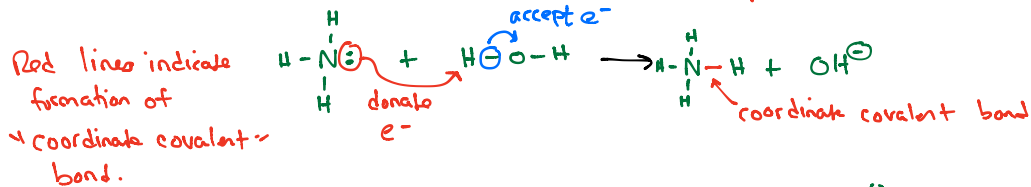


acid  
base



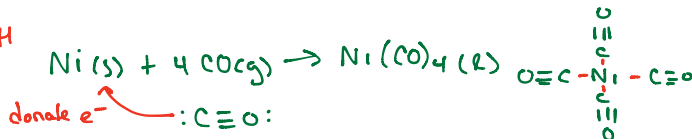
### Lewis Acid/Base

Lewis Acid = Electron acceptor  
 Lewis Base = Electron Donor  
 } this definition goes well beyond traditional acids + bases; Encompasses LOTS of molecules



Coordinate bond formed with

②  $e^-$  from one atom the base



CO = Lewis Base  
 Ni = Lewis Acid

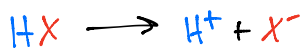


NOT an acid/base reaction.

## Nomenclature Review

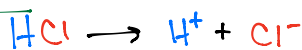
Acids are named based on the anions that the proton is paired with

General reaction =



↑ the name of this anion directs the name of the acid

-ide Rule



↑ chloride

Hydrochloric acid

① Hydro

② drop -ide

③ ic acid

\* This type of acid exclusively when halogen forms acid



-ate Rule



↑ nitrate

Nitric Acid

① Drop -ate

② add -ic acid

-ite rule



↑ nitrite

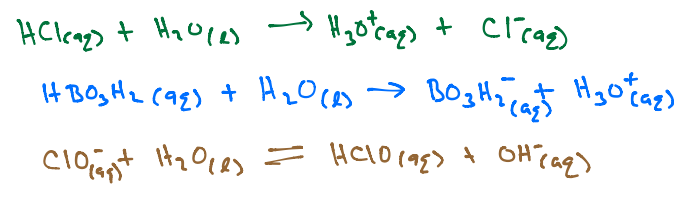
Nitrous acid

① Drop -ite

② add -ous acid

Acids and Bases react with water. You should be able to write an acid/base reaction if given a molecule in aqueous solution.

HCl, Boric acid ( $\text{HBO}_3\text{H}_2$ ), hypochlorite



Here is the cool thing, these reactions predict how acidic or basic a solution is. But how do we measure acidity? **pH scale**

$\text{pH} \leftarrow [\text{H}_3\text{O}^+]$       $\text{pH} = -\log[\text{H}_3\text{O}^+]$   
 $\swarrow \text{p} \equiv -\log$

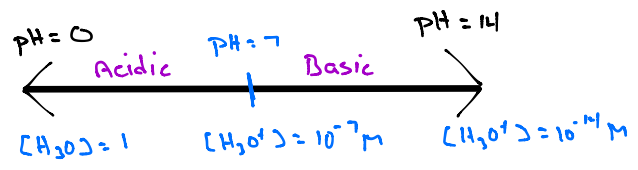
calculate  $[\text{H}_3\text{O}^+]$  for a solution at pH 8.3  
 $\text{pH} = -\log[\text{H}_3\text{O}^+]$   
 $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$   
 $[\text{H}_3\text{O}^+] = 10^{-8.3} = 5.01 \times 10^{-9} \text{ M}$

calculate pH for each of the following hydronium ion concentrations:

- $[\text{H}_3\text{O}^+] = 1.00 \text{ M}$       $\rightarrow -\log 1 = 0$       $\text{pH} = 0$
- $[\text{H}_3\text{O}^+] = 1.00 \times 10^{-7} \text{ M}$       $-\log (1 \times 10^{-7}) = 7$       $\text{pH} = 7$
- $[\text{H}_3\text{O}^+] = 1.00 \times 10^{-14} \text{ M}$       $\text{pH} = 14$
- $[\text{H}_3\text{O}^+] = 10 \text{ M}$       $\text{pH} = 0$
- $[\text{H}_3\text{O}^+] = 1.00 \times 10^{-15} \text{ M}$       $\text{pH} = 15$

The traditional pH scale goes from 1  $\rightarrow$  14, but it is NOT restricted to that range!

The higher  $[\text{H}_3\text{O}^+]$ , the lower the pH



- "Acidic"  $\equiv \text{pH} < 7$
- "Basic"  $\equiv \text{pH} > 7$
- "neutral"  $\equiv \text{pH} = 7$

What's so special about this number?  
 $[\text{H}_3\text{O}^+] = [\text{OH}^-]$

So,  $[H_3O^+] = 10^{-7} = [OH^-]$  → this leads to a VERY useful relationship:

$$[OH^-][H_3O^+] = 1 \times 10^{-14} \quad (@ 25^\circ C)$$

This comes from the autoionization of water equilibrium:



$K_w$  has awesome impact: if you know  $[H_3O^+]$ , you also know  $[OH^-]$ !

calculate  $[OH^-]$  for each of the following hydronium ion concentrations:

$[H_3O^+] = 1.00 \text{ M}$	→ $10^{-14} = (1)[OH^-]$	$[OH^-] = 10^{-14} \text{ M}$	$pOH = 14$
$[H_3O^+] = 1.00 \times 10^{-7} \text{ M}$	$10^{-14} = 10^{-7}[OH^-]$	$[OH^-] = 10^{-7} \text{ M}$	$pOH = 7$
$[H_3O^+] = 1.00 \times 10^{-14} \text{ M}$	$[OH^-] = 1.00 \text{ M}$		$pOH = 0$
$[H_3O^+] = 10 \text{ M}$	$[OH^-] = 10^{-15} \text{ M}$		$pOH = 15$
$[H_3O^+] = 1.00 \times 10^{-15} \text{ M}$	$[OH^-] = 10 \text{ M}$		$pOH = -1$

calculate  $pOH$  for each of the conditions above → this is calculated exactly like  $pH$

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

Another useful relationship that comes from  $K_w$ :

$$10^{-14} = [OH^-][H_3O^+] \quad -\log \text{ of both sides}$$

$$14 = pOH + pH \quad \text{The sum of } pH \text{ and } pOH \text{ always equals } 14$$

(Only at  $25^\circ C$  → remember  $K$  is Temp dependent)

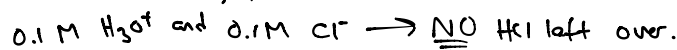
## Strong and Weak Acids / Bases



-these reactions are under the control of an equilibrium constant. However, some of these reactions favor products so much that an equilibrium constant cannot be measured. These are called Strong Acids and Strong Bases.



If you start with 0.1 M HCl, you will create



this is a strong acid, so the reaction proceed to completion.

### Strong Acids (Table 20.1)

HClO <sub>4</sub>	Perchloric Acid
HNO <sub>3</sub>	Nitric Acid
H <sub>2</sub> SO <sub>4</sub>	Sulfuric Acid
HCl	Hydrochloric Acid
HBr	Hydrobromic Acid
HI	Hydroiodic acid

(NOT HF!)

Strong Bases → anything that has OH<sup>-</sup> and is soluble in H<sub>2</sub>O

NaOH
LiOH
KOH
RbOH
Ca(OH) <sub>2</sub>
Ba(OH) <sub>2</sub>

Strong Acid/Base reactions boil down to stoichiometry (no equilibrium)

Calculate the [H<sub>3</sub>O<sup>+</sup>], [OH<sup>-</sup>], pH, and pOH of 1.3mM HBr at 25°C  
Start by writing out the reaction.



↑ strong acid, so not an equilibrium

$$\frac{1.3 \text{ mmol HBr}}{L} \left| \frac{1 \text{ mmol H}_3\text{O}^+}{1 \text{ mmol HBr}} \right| \frac{10^{-3} \text{ mol}}{1 \text{ mmol}} = 1.3 \times 10^{-3} \text{ M H}_3\text{O}^+ = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log[\text{H}_3\text{O}^+] = -\log 1.3 \times 10^{-3} = 2.87 = \text{pH}$$

$$\text{pH} + \text{pOH} = 14$$

$$2.87 + \text{pOH} = 14$$

$$\text{pOH} = 11.11$$

$$[\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14}$$

$$[\text{OH}^-] = 10^{-14} / 1.3 \times 10^{-3} = 7.69 \times 10^{-12} = [\text{OH}^-]$$

You have 2.8 L of 10 mM HCl. How much water do you need to add to have a pH=5?

10 mM = 0.01 M      pH = 2

current concentration      pH 5       $[H_3O^+] = 10^{-5} = [HCl]$

$$\frac{0.01 \text{ mol}}{L} (2.8 \text{ L}) = 0.028 \text{ mol HCl}$$



$$\frac{0.028 \text{ mol HCl}}{10^{-5} \text{ mol HCl}} \left| \frac{1 \text{ L}}{10^{-5} \text{ mol HCl}} \right| = 2800 \text{ L!}$$

desired concentration

Need to add  $2800 - 2.8 = 2797.2 \text{ L}$  of  $H_2O$

You have 2.8 L of 10 mM  $HNO_3$ . Your goal is to neutralize this solution by adding NaOH. If the NaOH you are adding is 1.00 M, calculate the volume that is needed. What is  $[H_3O^+]$  at this point?

① How many moles of  $H_3O^+$  are present?



$$\frac{2.8 \text{ L } HNO_3}{L} \left| \frac{0.01 \text{ mol}}{L} \right| \left| \frac{1 \text{ mol } H_3O^+}{1 \text{ mol } HNO_3} \right| = 0.028 \text{ mol } H_3O^+$$

So to neutralize this solution, 0.028 mol  $OH^-$  are needed (remember neutral is when  $[H_3O^+] = [OH^-]$ )

② Calculate volume of NaOH needed.

$$\frac{0.028 \text{ mol } OH^-}{1 \text{ mol } OH^-} \left| \frac{1 \text{ mol NaOH}}{1 \text{ mol } OH^-} \right| \left| \frac{1 \text{ L}}{1 \text{ mol NaOH}} \right| = 0.028 \text{ L NaOH needed}$$

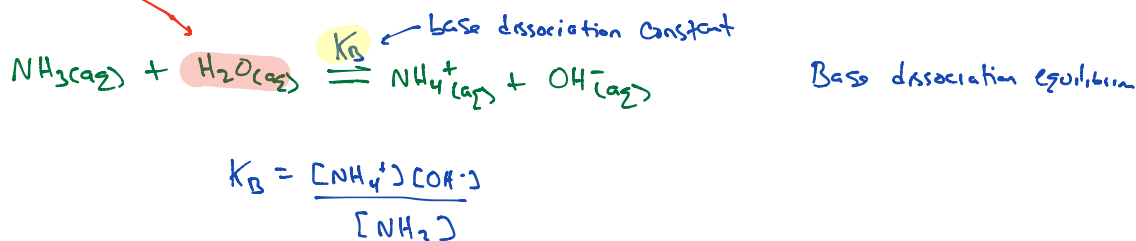
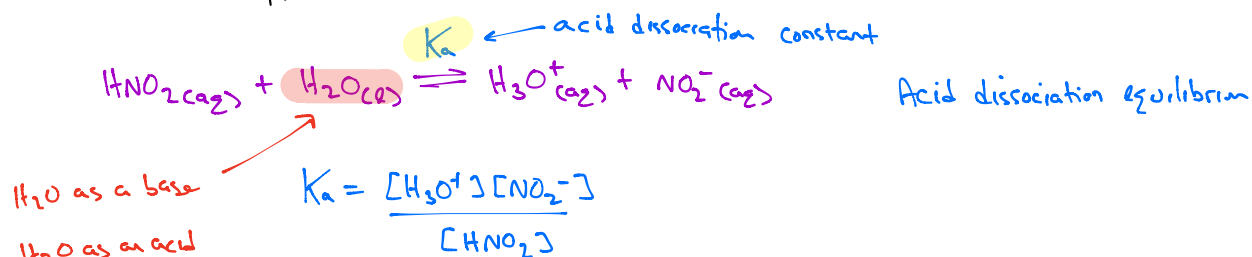
Once the NaOH has been added, a neutralization reaction occurs:



Since  $[OH^-]_{added} = [H_3O^+]_{from HNO_3}$ , only the autoionization of water matters

$$2H_2O(l) \rightleftharpoons [H_3O^+][OH^-] = 10^{-14} \quad [H_3O^+] = 10^{-7} M$$

Acids and bases react with water - the 1<sup>st</sup> thing you should do when you think about these molecules is to react them with H<sub>2</sub>O



H<sub>2</sub>O can be an acid or a base!



$$\text{pH} + \text{pOH} = 14$$

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 10^{-14} \text{ (@ } 25^\circ\text{C)}$$

this is awesome because we can now calculate [OH<sup>-</sup>] from [H<sub>3</sub>O<sup>+</sup>]

Strong acids and bases =  $K_a / K_b$  are not measurable! The equilibrium lies completely on the side of the products!

HCl, HBr, HI, HNO<sub>3</sub>, H<sub>2</sub>SO<sub>4</sub>

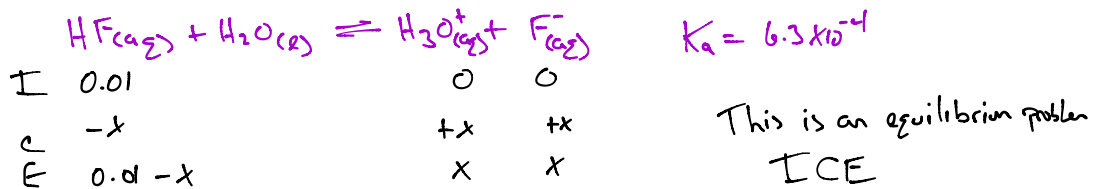
NOT HF → H-F bond is stronger than H-Cl bond ∴ harder to break. HF is a weak acid

## WEAK ACIDS/BASES

We saw above that 10 mM HCl has a pH of 2 → strong acid, so  $[H_3O^+] = 10 \text{ mM}$

But what is the pH of 10 mM HF

← NOT a strong acid!



$$6.3 \times 10^{-4} = \frac{x^2}{0.01 - x}$$

$$6.3 \times 10^{-6} - 6.3 \times 10^{-4} x = x^2$$

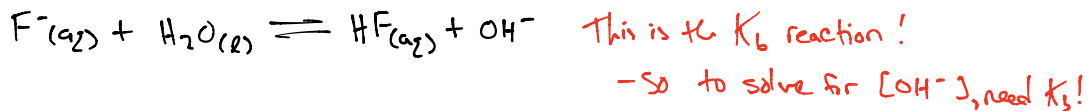
$$0 = x^2 + 6.3 \times 10^{-4} x - 6.3 \times 10^{-6}$$

$$x = \frac{-6.3 \times 10^{-4} + \sqrt{(6.3 \times 10^{-4})^2 - 4(-6.3 \times 10^{-6})}}{2} = 2.21 \times 10^{-3} = [H_3O^+]$$

$$pH = -\log = 2.65$$

= 10 mM F<sup>-</sup>

What about 10 mM NaF? → NOTE that NaF is soluble... in solution it is Na<sup>+</sup> + F<sup>-</sup>  
- we can't just reverse the K<sub>a</sub> reaction - Bases react with H<sub>2</sub>O, NOT H<sub>3</sub>O<sup>+</sup>!



$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

$$K_b = \frac{[OH^-][HF]}{[F^-]}$$

$$K_a \cdot K_b = K_w!$$

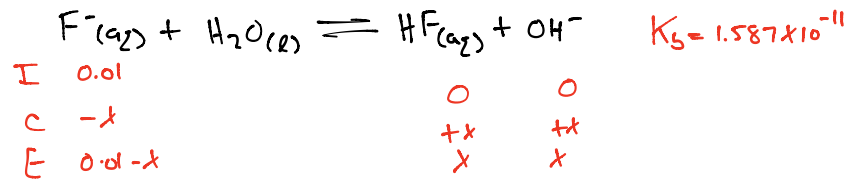
$$pK_a + pK_b = 14$$

$$\frac{[H_3O^+][F^-]}{[HF]} \cdot \frac{[OH^-][HF]}{[F^-]} = [H_3O^+][OH^-] = K_w$$

$$K_a \text{ for HF} = 6.3 \times 10^{-4}$$

$$K_b = \frac{10^{-14}}{6.3 \times 10^{-4}} = 1.587 \times 10^{-11}$$





$$1.587 \times 10^{-11} = \frac{x^2}{0.01-x}$$

$$x^2 = 1.587 \times 10^{-13} - 1.587 \times 10^{-11}x$$

$$0 = x^2 + 1.587 \times 10^{-11}x - 1.587 \times 10^{-13}$$

$$x = \frac{-1.587 \times 10^{-11} + \sqrt{(1.587 \times 10^{-11})^2 - 4(-1.587 \times 10^{-13})}}{2}$$

$$x = 3.984 \times 10^{-7} = [OH^{-}]$$

$$pOH = -\log [OH^{-}] = 6.4$$

$$pH = 14 - pOH = 7.6$$