Acids:

- Sour taste
- Turn litmus **red**
- Normally have H written first in formula (HCl, HNO_3, etc.)
- Treated like an ionic compound with H^+ / anion$^-$

  H^+, Cl^-

  Binary acids: H and one other element

  HCl, HBr, etc

  hydrochloric acid

  Oxyacids: H and an oxyanion

  HNO_3 nitric acid

  HNO_2 nitrous acid

Bases:

- Bitter taste
- Turn litmus **blue**

Ionic hydroxides: NaOH, $Ca(OH)_2$

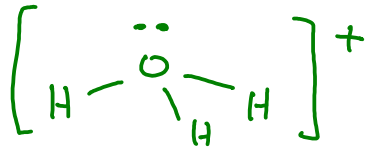
~~CH_3OH alcohol~~ not a base

Carbonates: Na_2CO_3 $NaHCO_3$ sodium bicarbonate (baking soda)

N-containing molecules: NH_3 ammonia

Arrhenius theory of acids and bases

Acids create H_3O^+ (hydronium) ion in water

hydronium ion structure: 

1. Dissociation:

$$HCl(g) \xrightarrow{\text{dissolve}} H^+(aq) + Cl^-(aq)$$

H–Cl molecules

2. Hydronium ion formation:

$$H^+(aq) + H_2O(l) \rightarrow H_3O^+$$

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

Every acid produces hydronium ions in water by a similar process:

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

$$H_2SO_4 + H_2O \rightarrow H_3O^+ + HSO_4^-$$

$$H_2SO_4 + 2H_2O \rightarrow 2H_3O^+ + SO_4^{2-}$$

Bases create OH^- (hydroxide) ion in water:

$$NaOH(s) \xrightarrow{\text{dissolves}} Na^+(aq) + OH^-(aq)$$

Acidity/basicity of solutions

A **strongly acidic** solution has a relatively high concentration of H_3O^+.

A **strongly basic** solution has a relatively high concentration of OH^-.

A **neutral** solution has equal conc's of H_3O^+ and OH^-

A basic solution still has some H_3O^+ ions,
but $[OH^-] > [H_3O^+]$ and vice-versa for acidic solutions.

↑ molar concentration = [ ]

The **pH scale** is a 0-14 scale which measures the acidity or basicity of solution.

1 pH unit means 10x in acidity

0 most acidic
1
2
3
4 ↖ 10x more acidic
5
6
7.00 neutral pH
8
9
10
11
12
13
14 most basic (alkaline)

pH values and calculations

pH = 2 means $[H_3O^+] = 10^{-2}$ *M*

pH = 3 means $[H_3O^+] =$ 10^{-3} M

Which is **more acidic**, pH = 2 or pH = 3?

$[H_3O^+] = 0.01$ M ↙ 0.001 M ↙

If $[H_3O^+] = 10^{-9}$ *M*, then pH = 9

Logarithms:

log 10^9 = 9

log 10^2 = 2

log 10^{-9} = −9

-log 10^{-9} = 9

pH = -log $[H_3O^+]$

$10^0 = 1$

$10^{-1} = \frac{1}{10} = .1$

$10^{-2} = \frac{1}{100} = .01$

$10^{-3} = \frac{1}{1000} = .001$

	$[H_3O^+]$
0	
1	
2	$= 10^{-2}$ M
3	10^{-3} M
4	
5	
6	
7	
8	
9	$= 10^{-9}$ M
10	
11	
12	
13	
14	

If $[H_3O^+] = 10^{-6}$ *M*, then pH = -log 10^{-6} = 6

If $[H_3O^+] = 1.0 \times 10^{-12}$ *M*, pH = 12.00

sig figs in $[H_3O^+]$ become decimal places in pH

If $[H_3O^+] = 4.2 \times 10^{-12}$ *M*, pH = $-\log(4.2 \times 10^{-12}) = 11.38$

Text calculator: (-) LOG 4.2 E (-) 12 =

Numerical calculator: 4.2 E 12 +/- LOG +/-

pH calculations

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

If $[H_3O^+] = 2.87 \times 10^{-3}$ *M*, what is pH?

$pH = -\log(2.87 \times 10^{-3}) = 2.542$

(in pH values, the sig figs are after the decimal point)

If pH = 8, $[H_3O^+] =$ 10^{-8} M

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

needs exact power of 10

If pH = 2.87, $[H_3O^+] = 10^{-2.87} =$ 0.0013 M $= 1.3 \times 10^{-3}$ M

text calculator: [10^x] [(-)] 2.87 [=]

numerical calculator: 2.87 [+/-] [10^x]

If pH = 6.43, $[H_3O^+] = 10^{-6.43} = 3.7 \times 10^{-7}$ M

(Use MODE or SCI if your calculator gives you 0.00000...)

$[OH^-]$ and pOH calculations

If $[OH^-] = 10^{-4}$ *M*, pOH = 4

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

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

$pH + pOH = 14$

pH = 8.73

pOH = 14 - 8.73

=

pH		pOH
0	acidic	14
1		13
2		12
3		11
4		10
5		9
6		8
7	neut.	7
8		6
9		5
10		4
11		3
12		2
13		1
14	basic	0

$[H_3O^+]$ $\xrightarrow{-\log [H_3O^+]}$ pH $\xrightarrow{pH+pOH=14}$ pOH $\xrightarrow{10^{-pOH}}$ $[OH^-]$

$\xleftarrow{10^{-pH}}$ $\xleftarrow{-\log [OH^-]}$

If $[OH^-] = 0.0230$ *M*, pOH = $-\log(.0230)$ = 1.638

$14 - 1.638$

pH = 12.362

$10^{-12.362}$

$[H_3O^+] = 4.35 \times 10^{-13}$ M

10^x (-) 12.362 =

12.362 = 10^x

pH shows this solution is: acidic / basic / neutral

Calculations for specific solutions

acids → H_3O^+
bases → OH^-

What is the pH, pOH, $[OH^-]$, and $[H_3O^+]$ of 1.50×10^{-3} *M* HCl(*aq*)?

Which is HCl? acid or base

$[HCl] = 1.50 \times 10^{-3}$ M

$[H_3O^+] = 1.50 \times 10^{-3}$ M

pH = 2.824 (acidic)

pOH = 11.176

$[OH^-] = 6.67 \times 10^{-12}$ M

These equations will be given on the final exam:

- pH = -log $[H_3O^+]$
- $[H_3O^+] = 10^{-pH}$
- pH + pOH = 14
- pOH = -log $[OH^-]$
- $[OH^-] = 10^{-pOH}$

What is the pH, pOH, $[OH^-]$, and $[H_3O^+]$ of 2.43×10^{-5} *M* KOH(*aq*)?

Which is KOH? acid or base

$[OH^-] = 2.43 \times 10^{-5}$ M

pOH = 4.614

pH = 9.386 (basic)

$[H_3O^+] = 4.12 \times 10^{-10}$ M

When an acid and a base react, treat them like ionic compounds and do a double displacement reaction, remembering that H^+ and OH^- combine to form:

$$\underline{2}\ HNO_3 + \underline{\quad}\ Ca(OH)_2 \rightarrow 2\,H_2O + Ca(NO_3)_2$$

H^+/NO_3^- Ca^{2+}/OH^- H^+/OH^- Ca^{2+}/NO_3^-

$$\underline{\quad}\ H_2SO_4 + \underline{2}\ NaOH \rightarrow 2\,H_2O + Na_2SO_4$$

H^+/SO_4^{2-} Na^+/OH^- H^+/OH^- Na^+/SO_4^{2-}