#### **Chapter 17: Acids and Bases I**

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#### Acids and Bases

#### <u>Acids</u>

- Sour taste (vinegar)
- Dissolve many metals
- Ability to neutralize bases
- Strong or Weak

#### <u>Bases</u>

Bitter taste

(caffeine, poisons from plants - alkaloids)

- Slippery feel
- Ability to neutralize acids
- Strong or Weak

• Definitions: Arrhenius, Bronsted-Lowry, Lewis





#### **Common acids**

TABLE 16.1 Some Common Acids									
Name	Occurrence/Uses								
Hydrochloric acid (HCI)	Metal cleaning; food preparation; ore refining; primary component of stomach acid								
Sulfuric acid (H <sub>2</sub> SO <sub>4</sub> )	Fertilizer and explosives manufacturing; dye and glue production; in automobile batteries; electroplating of copper								
Nitric acid (HNO <sub>3</sub> )	Fertilizer and explosives manufacturing; dye and glue production								
Acetic acid $(HC_2H_3O_2)$	Plastic and rubber manufacturing; food preservation; active component of vinegar								
Citric acid (H <sub>3</sub> C <sub>6</sub> H <sub>5</sub> O <sub>7</sub> )	In citrus fruits such as lemons and limes; used to adjust pH in foods and beverages								
Carbonic acid (H <sub>2</sub> CO <sub>3</sub> )	In carbonated beverages due to the reaction of carbon dioxide with water								
Hydrofluoric acid (HF)	Metal cleaning; glass frosting and etching								
Phosphoric acid (H <sub>3</sub> PO <sub>4</sub> )	Fertilizer manufacturing; biological buffering; beverage preservation								

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#### **Common bases**

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Name	Occurrence/Uses				
Sodium hydroxide (NaOH)	Petroleum processing; soap and plastic manufacturing				
Potassium hydroxide (KOH)	Cotton processing; electroplating; soap production; batteries				
Sodium bicarbonate (NaHCO <sub>3</sub> )	Sold as baking soda; antacid; source of CO <sub>2</sub>				
Sodium carbonate (Na <sub>2</sub> CO <sub>3</sub> )	Glass and soap manufacturing; general cleanser; water softener				
Ammonia (NH <sub>3</sub> )	Detergent; fertilizer and explosives manufacturing; synthetic fiber production				

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#### **Arrhenius Definition**





- Acid: A substance that produces H<sup>+</sup> ions in aqueous solution
- Base: A substance that produces OH<sup>-</sup> (hydroxide) ions in aqueous solution



HCl(aq) →H<sup>+</sup>(aq) + Cl<sup>-</sup>(aq) NaOH(aq) →Na<sup>+</sup>(aq) + OH<sup>-</sup>(aq) Svante Arrhenius 1859-1927 Sweden



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#### **Hydronium ion**

- The H<sup>+</sup> ion is a proton
- In reality, the H<sup>+</sup> ion always associates with water molecules as H(H<sub>2</sub>O)<sub>n</sub><sup>+</sup>



- Often said to be  $H_3O^+$ , the hydronium ion
- Chemists use  $H^+$  and  $H_3O^+$  interchangeably

#### **Acid-base neutralization**

• Arrhenius:

Acids and bases react to form water, a neutral compound

### $\mathsf{H}^+(aq) + \mathsf{OH}^-(aq) \rightarrow \mathsf{H}_2\mathsf{O}(I)$

Warning: Neutralization reactions can release lot of heat!



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#### **Brønsted-Lowry Definition**

- Acid: Proton (H<sup>+</sup>) donor
- Base: Proton (H<sup>+</sup>) acceptor



Martin Lowry Johannes Brønsted

### HCl(aq) + H<sub>2</sub>O(/) → H<sub>3</sub>O<sup>+</sup>(aq) + Cl<sup>-</sup>(aq) Bronsted-Lowry Base Bronsted-Lowry Acid

 $\begin{array}{c} \mathsf{NH}_3(aq) + \mathsf{H}_2\mathsf{O}(I) \rightleftharpoons \mathsf{NH}_4^+(aq) + \mathsf{OH}^-(aq) \\ & \uparrow \\ & \mathsf{Bronsted-Lowry\ Acid} \\ & \mathsf{Bronsted-Lowry\ Base} \end{array}$ 



#### **Polyprotic acids and bases**



- Polyprotic acids: Can donate two or more protons
- Polyprotic bases: Can accept two or more protons

**Example:** H<sub>2</sub>SO<sub>4</sub> (sulfuric acid)

 $H_2SO_4(aq) + H_2O(\ell) \longrightarrow H_3O^+(aq) + HSO_4^-(aq)$  (1)

 $HSO_4^{-}(aq) + H_2O(\ell) \longrightarrow H_3O^{+}(aq) + SO_4^{2-}(aq)$  (2)

(produces two  $H_3O^+$  in a sequence)



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#### **Carboxylic acids**

Organic molecules with a -COOH group:





#### **Amphoteric compounds**

 They can act as either Bronsted-Lowry acids or bases (for example, water):

 $H_2PO_4^{-}(aq) + H_2O(\ell) \Longrightarrow HPO_4^{2-}(aq) + H_3O^{+}(aq)$ 

 $H_2PO_4^{-}(aq) + H_2O(\ell) \Longrightarrow H_3PO_4(aq) + OH^{-}(aq)$ 





#### Bronsted-Lowry Conjugate Acids and Bases



### $NH_3(aq) + H_2O(I) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

#### Forward: $NH_3(aq) + H_2O(I) \rightarrow NH_4^+(aq) + OH^-(aq)$ Bronsted-Lowry Acid Bronsted-Lowry Base

### **Reverse:** $NH_4^+(aq) + OH^-(aq) \rightarrow NH_3(aq) + H_2O(/)$ Bronsted-Lowry Base Bronsted-Lowry Acid



#### **Conjugate Acid-Base Pairs**



### $NH_3(aq) + H_2O(I) \Rightarrow NH_4^+(aq) + OH^-(aq)$



- Base accepts a proton and becomes an acid in the reverse reaction (conjugate acid).
- Acid donates a proton and becomes a base in the reverse reaction (conjugate base).



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#### **Conjugate Acid-Base Pairs**

### $NH_3(aq) + H_2O(I) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$



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#### Example 1

Identify the Bronsted-Lowry acid and base and their conjugates in the following reaction:



- Forward:  $H_2SO_4$  (acid) donates  $H^+$  to  $H_2O$  (base).
- Reverse:  $H_3O^+$  (acid) donates  $H^+$  to  $HSO_4^-$  (base).



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#### Example 2

Identify the Bronsted-Lowry acid and base and their conjugates in the following reaction:



- Forward:  $H_2O$  (acid) donates  $H^+$  to  $HCO_{3^-}$  (base).
- Reverse: H<sub>2</sub>CO<sub>3</sub> (acid) donates H<sup>+</sup> to OH<sup>-</sup> (base).



# Example 3

Identify the Bronsted-Lowry acid and base and their conjugates in the following reaction:



- Forward:  $H_2O$  (acid) donates  $H^+$  to  $CH_3NH_2$  (base).
- Reverse:  $CH_3NH_3^+$  (acid) donates  $H^+$  to  $OH^-$  (base).

#### Lewis Acids and Bases

- Lewis model focuses on the electrons instead of the protons
- Lewis acid: Accepts electron pairs
- Lewis base: Donates electron pairs
- Lewis acid: Has an empty orbital (or can rearrange electrons to create an empty orbital) that can accept an electron pair



CO<sub>2</sub> rearrange & accept







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#### Acid strength

- Strong acid: Completely ionizes in solution
- Weak acid: Partially ionizes in solution
- The 6 strong acids (memorize)

TABLE 16.3 Strong Acids	
Hydrochloric acid (HCI)	Nitric acid (HNO <sub>3</sub> )
Hydrobromic acid (HBr)	Perchloric acid (HClO <sub>4</sub> )
Hydriodic acid (HI)	Sulfuric acid (H <sub>2</sub> SO <sub>4</sub> ) ( <i>diprotic</i> )

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#### Strong acid

Completely dissociates as ions in aqueous solution:

$$HCl(aq) + H_2O(\ell) \longrightarrow H_3O^+(aq) + Cl^-(aq)$$

$$Forward arrow only$$

- A solution that is "1.0 M HCl" is actually 1.0 M  $H_3O^+$  and 1.0 M  $Cl^-$
- Practically no undissociated HCI exists in the solution.



#### Weak acid

Partially dissociates into ions in aqueous solution establishing equilibrium:

### $HF(aq) + H_2O(\ell) \Longrightarrow H_3O^+(aq) + F^-(aq)$ Equilibrium arrows

For example, a solution that is "1.0 M HF":

- Some of the HF dissociates, but some (usually most) stays as HF in the solution
- Some H<sub>3</sub>O<sup>+</sup> is formed
- Need to solve equilibrium problem (ICE table)







#### Acid ionization constant, K<sub>a</sub> (dissociation constant)

How do we determine the concentrations of a weak acid at equilibrium?

Generic weak acid reaction:

$$HA(aq) + H_2O(\ell) \Longrightarrow H_3O^+(aq) + A^-(aq)$$
  
or 
$$HA(aq) \Longrightarrow H^+(aq) + A^-(aq)$$

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{[H^+][A^-]}{[HA]}$$

**Remember:** No pure liquids in equilibrium expression!



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#### Example

Write the acid dissociation expression ( $K_a$ ) for the following reaction:

 $HNO_2(aq) + H_2O(\ell) \Longrightarrow H_3O^+(aq) + NO_2^-(aq)$ 

# $K_a = \frac{[H_3O^+][NO_2^-]}{[HNO_2]}$

Note:  $H_2O$  is pure liquid and not included!



#### Range of K<sub>a</sub> values



• For a strong acid,  $K_a$  is very large ( $\infty$ ):

Use one arrow to indicate that there is essentially zero reactant left after the reaction

• For weak acids, *K<sub>a</sub>* can vary depending on the "strength" of the acid:

Set up as an equilibrium

Weak Acid	Formula	Ka			
Hydrofluoric acid	HF	1.8 x 10 <sup>-4</sup>			
Acetic acid	$HC_2H_3O_2$	1.8 x 10 <sup>-5</sup>			
Hydrocyanic acid	HCN	4.9 x 10 <sup>-10</sup>			



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#### Autoionization of water

Recall that water can act as an acid or a base (amphoteric):

$$\begin{split} \mathrm{HCl}(\mathrm{aq}) + \mathrm{H_2O}(\ell) &\longrightarrow \mathrm{H_3O^+}(\mathrm{aq}) + \mathrm{Cl^-}(\mathrm{aq}) \\ \\ & \mathsf{base} \end{split}$$

#### $\operatorname{NH}_{3}(\operatorname{aq}) + \operatorname{H}_{2}O(\ell) \rightleftharpoons \operatorname{NH}_{4}^{+}(\operatorname{aq}) + \operatorname{OH}^{-}(\operatorname{aq})$ acid

#### Autoionization of water

• Even by itself, in pure water,  $H_2O$  can react with itself; autoionization:

### $H_2O(\ell) + H_2O(\ell) \Longrightarrow H_3O^+(aq) + OH^-(aq)$ $H_2O(\ell) \Longrightarrow H^+(aq) + OH^-(aq)$

• This reaction has a special equilibrium constant,  $K_w$ "ion product constant" or "autoionization constant"

$$K_w = [H_3O^+][OH^-] = [H^+][OH^-]$$







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#### <u>Pure</u> water

 For pure water, H<sub>2</sub>O is the only source of H<sup>+</sup> or OH<sup>-</sup>, so they must have equal concentrations:

For every H<sup>+</sup> produced, there must be an OH<sup>-</sup> (neutral)

• At 25 °C (\*),  $K_w = 1.0 \times 10^{-14}$ , therefore:

$$K_w = 1.0 \times 10^{-14} = [H^+][OH^-] = x^2$$
  
 $[H^+] = [OH^-] = 1.0 \times 10^{-7} M$ 

\*Use in this class, assume this temperature for acids & bases unless otherwise noted.



#### **Acidic solution**



- [H<sup>+</sup>] > [OH<sup>-</sup>]  $K_w = 1.0 \times 10^{-14} = [H^+][OH^-]$
- Product of [H<sup>+</sup>] times [OH<sup>-</sup>] must equal 1.0 x 10<sup>-14</sup>
- Example: if  $[H^+] = 1.0 \times 10^{-6} \text{ M}$ ,  $[OH^-]$  must be  $1.0 \times 10^{-8} \text{ M}$

#### Try filling out the table below:

[H+] (M)	[OH <sup>-</sup> ] (M)	Product
1.0 x 10 <sup>-5</sup>		1.0 x 10 <sup>-14</sup>
1.0 x 10 <sup>-4</sup>		1.0 x 10 <sup>-14</sup>
2.5 x 10 <sup>-5</sup>		1.0 x 10 <sup>-14</sup>
	3.7 x 10 <sup>-8</sup>	1.0 x 10 <sup>-14</sup>



#### **Basic solution**



- [OH<sup>-</sup>] > [H<sup>+</sup>]  $K_w = 1.0 \times 10^{-14} = [H^+][OH^-]$
- Product of [H<sup>+</sup>] times [OH<sup>-</sup>] must equal 1.0 x 10<sup>-14</sup>
- Example: if  $[H^+] = 1.0 \times 10^{-8} \text{ M}$ ,  $[OH^-]$  must be 1.0 x 10<sup>-6</sup> M

#### Try filling out the table below:

[H+] (M)	[OH <sup>-</sup> ] (M)	Product
	1.0 x 10 <sup>-3</sup>	1.0 x 10 <sup>-14</sup>
	1.0 x 10 <sup>-5</sup>	1.0 x 10 <sup>-14</sup>
	8.5 x 10 <sup>-6</sup>	1.0 x 10 <sup>-14</sup>
6.2 x 10 <sup>-8</sup>		1.0 x 10 <sup>-14</sup>



#### pH scale

- It is annoying to have to keep writing these low concentrations using scientific notation.
- Using the pH scale helps:

"*p*" = -log

$$pH = -log[H_3O^+] = -log[H^+]$$
$$pOH = -log[OH^-]$$

Conversion:  $10^{-pH} = [H^+]$  or  $10^{-pOH} = [OH^-]$ 

In this class:  $\log = 10$ -based,  $\ln = e$ -based.





#### A word about sig. figs. and logarithms

For logarithms, exponents  $n = 10^a \Leftrightarrow log(n) = a$ log(12.0) = 1.0791812... characteristic mantissa  $10^{1.0792...} = 12.0$ 

<u>Log</u>: Number of digits in mantissa of log x is number of significant figures in x: log(12.0) = 1.079 (3 sf. for mantissa).

Exponent: Number of digits in  $10^{\times}$  is number of significant figures in mantissa of x:  $10^{2.3456} = 2.216 \times 10^{2}$  (4 sf. for mantissa).





### A word about sig. figs. and logarithms

#### To exemplify:









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#### pH of pure water

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$
  
 $pK_w = pH + pOH = 14$ 

For pure water:

$$[H^+] = [OH^-] = 1.0 \times 10^{-7} M$$

$$pH = -log[H^+] = -log(1.0 \times 10^{-7} M) = 7$$

$$pOH = -log[OH^{-}] = -log(1.0 \times 10^{-7} M) = 7$$



#### pH scale

- Pure water (neutral): pH = 7 (at 25 °C)
- Acidic solution: pH < 7 (or pOH > 7)
- Basic solution: pH > 7 (or pOH < 7)
- Change in <u>1 pH unit</u> is a change in [H<sup>+</sup>] of a <u>factor of 10</u> <u>M units</u>

"logarithmic" scale:



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#### pH of acid solutions



Acidic solutions: pH < 7 ([H<sup>+</sup>] > 1 x 10<sup>-7</sup> M)

Example: 
$$[H^+] = 1.0 \times 10^{-6} M \Rightarrow pH = 6.00$$

Try filling out the table below:

[H+] (M)	[OH <sup>-</sup> ] (M)	рН
1.0 x 10 <sup>-5</sup>		
1.0 x 10 <sup>-4</sup>		
2.5 x 10 <sup>-5</sup>		
	3.7 x 10 <sup>-8</sup>	



#### pH of basic solutions



Basic solutions: pH > 7 ([H<sup>+</sup>] < 1 x 10<sup>-7</sup> M)

Example: 
$$[H^+] = 1.0 \times 10^{-8} M \Rightarrow pH = 8.00$$

Try to fill out the table below:

[H+] (M)	[OH <sup>-</sup> ] (M)	рН
	1.0 x 10 <sup>-3</sup>	
	1.0 x 10 <sup>-5</sup>	
	8.5 x 10 <sup>-6</sup>	
6.2 x 10 <sup>-8</sup>		



#### pOH scale

- Sometimes it is easier to calculate pOH first, then pH
- Sometimes it is easier to convert concentrations first
- Remember: they are related to each other

$$K_w = [H^+][OH^-] = 1.0 \times 10^{-14}$$
  
 $pK_w = pH + pOH = 14$ 

$\langle$	0.0	1.0	2.0	3.0	4.0	5.0	6.0	7.0	8.0	9.0	10.0	11.0	12.0	13.0	14.0
Ac	Acidic											Basic			
$\langle$	14.0	13.0	12.0	11.0	10.0	9.0	8.0	7.0	6.0	5.0	4.0	3.0	2.0	1.0	0.0
								рОН							

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#### Note: pH can be negative (e.g., $[H^+] = 3.0 \text{ M}$ )!



#### **Tips for this section**

- Learn the definitions:
  - Arrhenius Acid/Base, Bronsted-Lowry Acid/Base, Lewis Acid/Base
  - Memorize the six strong acids
- Be able to predict what might be an acid or base:
  - Think about what happens in aqueous solution
  - e.g. HCI, NaOH, or carboxylic acid
- Be able to identify the conjugate acid/base pairs in a reaction
- Practice pH calculations!