

# The Nature of Acid-Base Equilibria

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# Properties of Acids and Bases

## Acids

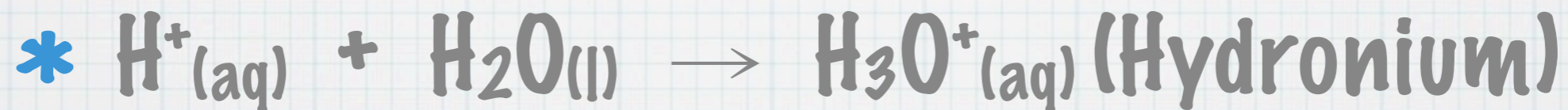
- Sour tasting
- React with some metal to form  $H_2$  gas
- Turns blue litmus red
- Colourless in phenolphthalein
- Conduct electricity

## Bases

- Bitter tasting
- Feels slippery
- Turns red litmus blue
- Pink in phenolphthalein
- Conduct electricity

# Arrhenius Acids and Bases

\* **Acid:** a substance that produces  $\text{H}_3\text{O}^+$  when dissolved in water.

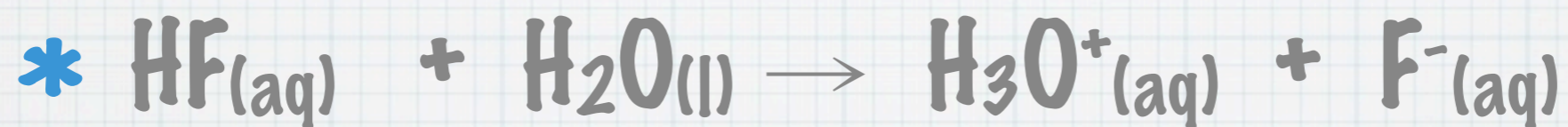


\* **Base:** a substance that produces  $\text{OH}^-$  when dissolved in water



# Bronsted-Lowry Acids and Bases

\* Acids: proton ( $\text{H}^+$ ) donors



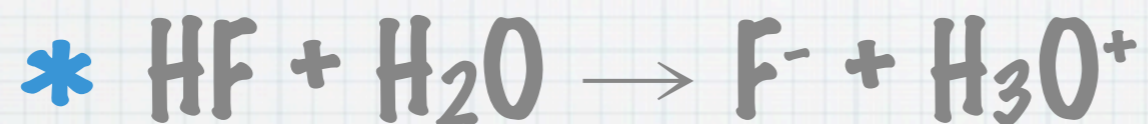
\* Bases: proton acceptors



**$\text{H}_2\text{O}$ : acts as an acid and a base  
AMPHOTERIC**

# Bronsted-Lowry Acids and Bases

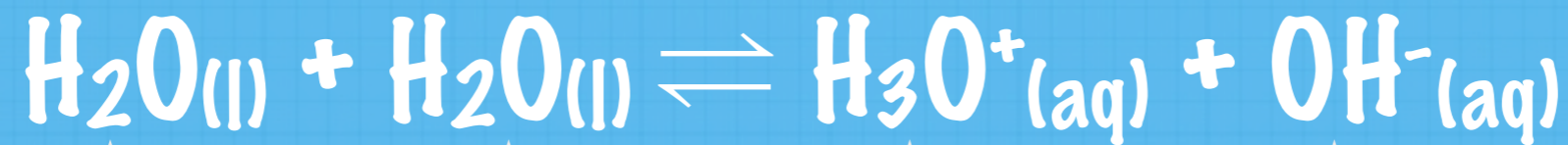
\* **Acid-Base Conjugate Pairs:** two species that differ by a proton



\* HF and F<sup>-</sup> make up an acid-base conjugate pair. F<sup>-</sup> is the conjugate base of HF.

# Ion Product Constant of Water ( $K_w$ )

## Autoionization of Water



Acid

Base

Acid

Base

# Ion Product Constant of Water ( $K_w$ )

- \* The equilibrium constant for this reaction is called  $K_w$
- \*  $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$
- \*  $K_w$  changes with temperature and is equal to  $1.0 \times 10^{-14}$  at  $25^\circ\text{C}$ .

# Example

- \* Calculate the hydroxide ion concentration of a solution in which the hydrogen ion concentration is  $3.6 \times 10^{-3}$  M



# Solution

\*  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$

\*  $(3.6 \times 10^{-3})[\text{OH}^-] = 1.0 \times 10^{-14}$

\*  $[\text{OH}^-] = 2.8 \times 10^{-12} \text{ M}$

# pH Scale

\*  $\text{pH} = -\log_{10}[\text{H}_3\text{O}^+]$

# Example

- \* The pH reading of a solution is 10.33.  
What is its hydrogen ion concentration?

# Solution

- \*  $\text{pH} = -\log[\text{H}^+]$

- \*  $[\text{H}^+] = 10^{-\text{pH}}$

- \*  $[\text{H}^+] = 10^{-10.33}$

- \*  $[\text{H}^+] = 4.7 \times 10^{-11} \text{ M}$

# P-notation

\* The p-notation can be used for any scale with low numbers.

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

# The Relationship between pH and pOH



\*  $pH + pOH = 14.00$

# Example 1

- \* Water taken from a lake was found to have  $[H^+] = 3.2 \times 10^{-5}$  M. Calculate the pH and pOH.

# Solution

\*  $\text{pH} = -\log[3.2 \times 10^{-5}]$

\*  $\text{pH} = 4.49$

\*  $\text{pOH} = 14 - 4.49$

\*  $\text{pOH} = 9.51$



# Strong Acids and Cases

- \* ionize (splits up into ions) almost 100% in water
- \* NO EQUILIBRIUM
- \* mostly ions in solution

# Example

- \* Calculate the pH of a 0.017 M  $\text{Ba}(\text{OH})_2$  solution

# Solution



\*  $[\text{OH}^{-}] = 2 \times [\text{Ba}(\text{OH})_2] = 0.034 \text{ M}$

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

\*  $= -\log(0.034)$

\*  $\text{pOH} = 1.47$

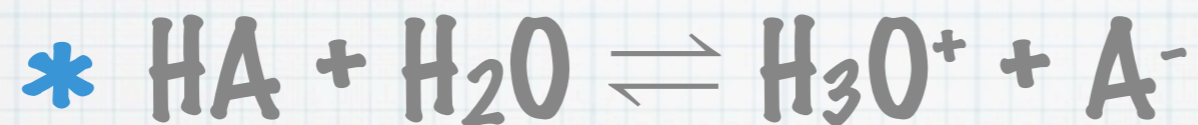
\*  $\text{pH} = 14 - \text{pOH}$

\*  $= 12.53$

# Weak Acids and Bases

- \* Weak acids and weak bases are those that do not ionize completely in water
- \* A weak acid exists in equilibrium with its conjugate base.
- \*  $\text{HCN}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{H}_3\text{O}^+_{(aq)} + \text{CN}^-_{(aq)}$

# Competition for protons



- \* Both  $\text{H}_2\text{O}$  and  $\text{A}^-$  are bases and compete for the proton
- \* If  $\text{HA}$  is a strong acid,  $\text{A}^-$  is an extremely poor proton acceptor
- \* If  $\text{HA}$  is a weak acid,  $\text{A}^-$  is good proton acceptor

# Homework

\* p 502 # 21, 22, 23