

## 8.1 The Nature of Acid-Base Equilibria

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- Acidic solutions are sour tasting, conduct electricity, and turn blue litmus red. (Ionize) - reaction
- Basic solutions, like aqueous ammonia, also conduct electricity, are bitter tasting, feel slippery, and turn red litmus blue. (Dissociate) - dissolving
- Arrhenius: acids produce  $\text{H}^+_{(\text{aq})}$  ions and bases produce  $\text{OH}^-_{(\text{aq})}$  ions.
- Brønsted-Lowry: a proton is transferred from one reactant to the other.

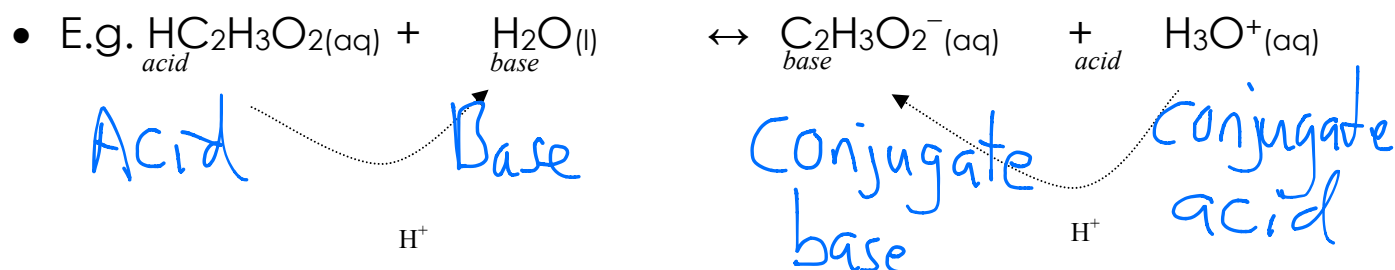
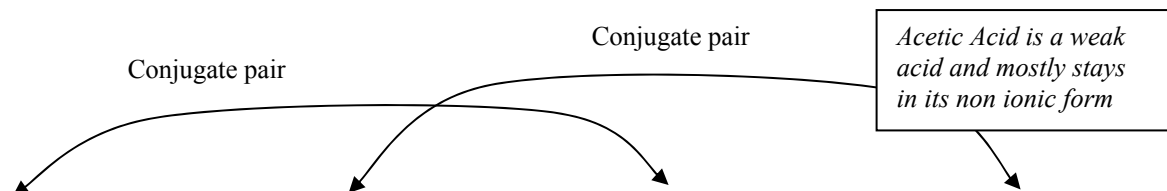
### Brønsted-Lowry Theory

( $\text{H}^+$ )

- A Brønsted-Lowry acid is a proton donor and a Brønsted-Lowry base is a proton acceptor.
- E.g.  $\text{H}_2\text{O}_{(\text{l})} + \text{HCl}_{(\text{g})} \rightarrow \text{H}_3\text{O}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$  (water forms the hydronium ion)  
 $\text{H}_2\text{O}_{(\text{l})}$  is the Brønsted-Lowry base and  $\text{HCl}_{(\text{g})}$  Brønsted-Lowry acid.
- E.g.  $\text{NH}_3_{(\text{g})} + \text{H}_2\text{O}_{(\text{l})} \rightarrow \text{NH}_4^+_{(\text{aq})} + \text{OH}^-_{(\text{aq})}$  (water forms the hydroxide ion)  
 $\text{H}_2\text{O}_{(\text{l})}$  is the Brønsted-Lowry acid and  $\text{NH}_3_{(\text{g})}$  Brønsted-Lowry base.
- Water is amphoteric (amphiprotic) which is a substance capable of acting as an acid or a base in different chemical reactions.

## Reversible Acid-Base Reactions

- In a reversible reaction there is an acid and a base in each of the 2 reactions. They are known as conjugate acid-base pairs (difference of only a single proton)



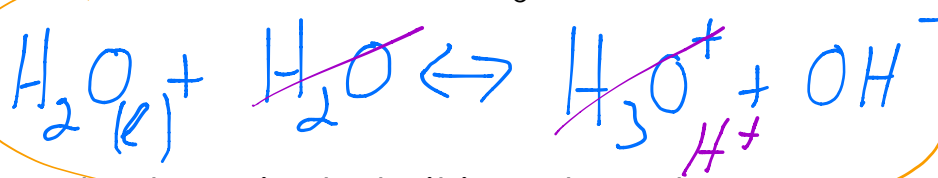
## Competition for Protons

- In the reaction above the acetic acid is a weak acid and the reactants are favoured. HCl is a strong acid and products are favoured.
- The stronger the acid, the weaker its conjugate base, and conversely, the weaker an acid, the stronger its conjugate base.

## The Autoionization of Water

- Autoionization of water is the reaction between two water molecules producing a hydronium ion and a hydroxide ion.
- $\text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$  ~~X~~ *Simplicity*
- It will occur in pure water. Put into perspective it will only happen to about 2 water molecules out of every billion (at SATP).
- We can apply the equilibrium law to find the  $K_e$  value.

$$\frac{[\text{H}^+(\text{aq})][\text{OH}^-(\text{aq})]}{[\text{H}_2\text{O}(\text{l})]} = K_e$$



- Since the  $[\text{H}_2\text{O}]$  is constant we include it in  $K_e$  to get...

$$K_w = [H^+_{(aq)}][OH^-_{(aq)}]$$

in pure water the concentration of  $[H^+_{(aq)}] = [OH^-_{(aq)}] = 1.0 \times 10^{-7} \text{ mol/L}$

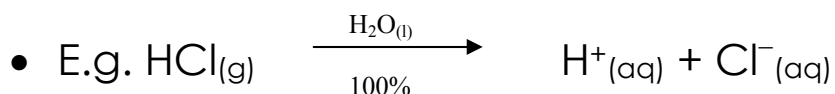
Therefore  $K_w = 1.0 \times 10^{-14}$  (at SATP)

The equation can be rewritten to find concentrations of  $[H^+_{(aq)}]$  and  $[OH^-_{(aq)}]$ .

- $[H^+_{(aq)}] = [OH^-_{(aq)}]$  therefore the solution is neutral
- $[H^+_{(aq)}] > [OH^-_{(aq)}]$  the solution is acidic
- $[H^+_{(aq)}] < [OH^-_{(aq)}]$  the solution is basic

## Strong Acids

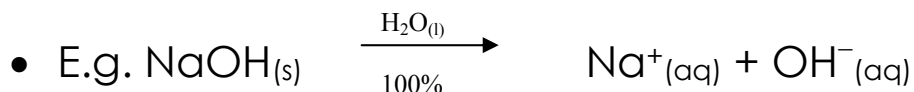
- A strong acid is an acid that ionizes completely in water to form hydrogen ions.



- Strong acids: hydrochloric, hydrobromic, sulfuric, nitric, and phosphoric.
- Monoprotic acids are acids that possess only one ionizable (acidic) proton. There are also diprotic and triprotic acids.

## Strong Bases

- A strong base is an ionic substance that (according to Arrhenius) disassociates completely in water to release hydroxide ions.



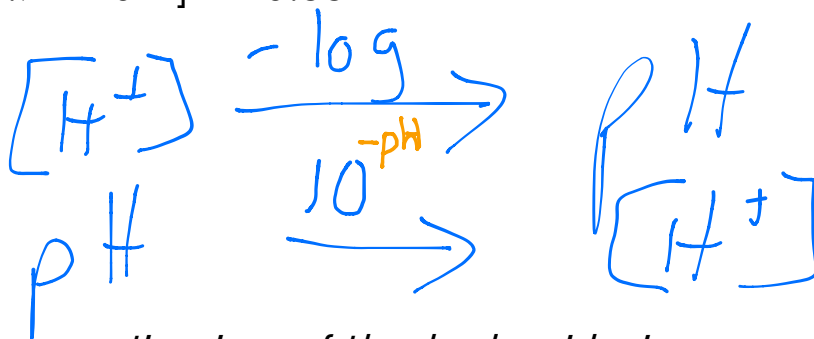
- Strong bases: group 1 and 2 hydroxides.

## Hydrogen Ion Concentration and pH

- $\text{pH} = -\log[\text{H}^+_{(\text{aq})}]$  (*the negative log of the hydrogen ion concentration*)
- a logarithmic scale a pH of 2 is 10 times more acidic than a pH of 3.
- E.g. pH of a solution with a hydrogen ion conc. of  $4.7 \times 10^{-11}$  mol/L

$$\text{pH} = -\log[\text{H}^+_{(\text{aq})}] = -\log[4.7 \times 10^{-11}] = 10.33$$

- pH = 7 is neutral
- pH > 7 is basic
- pH < 7 is acidic



## pOH and $\text{pK}_w$

- $\text{pOH} = -\log[\text{OH}^+_{(\text{aq})}]$  (*the negative log of the hydroxide ion concentration*)
- $\text{pK}_w = -\log K_w$  (*the negative log of the equilibrium constant for water*)
- $\text{pH} + \text{pOH} = \text{pK}_w$        $\text{pH} + \text{pOH} = 14$  (at SATP)

Handwritten example:  $\text{pH} = 4$   
 $[\text{H}^+] = 1 \times 10^{-4}$

## Measuring pH

- We will use acid base indicators such as litmus and pH meters.
- Litmus:  $\text{pH} < 4.7$  = red,  $\text{pH} > 8.3$  = blue,  $4.7 > \text{pH} < 8.3$  = brown.

## pH of Strong Acids and pH of Strong Bases

- Excellent example for acids on page 545.
- Excellent example for bases on page 547.

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## Homework

- Practice 1,2,3,4,5,6,8,9,10,11,12,13,14,15,17,18,19
- Questions 1,2,3,4,5,6,7