

8.1 The Nature of Acid-Base Equilibria

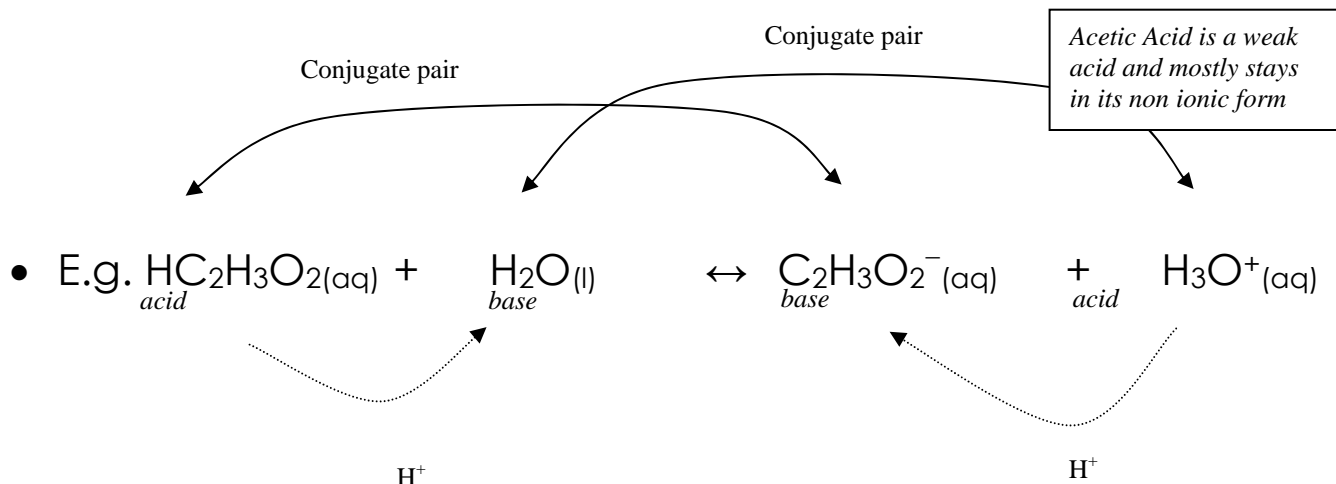
- Acidic solutions are sour tasting, conduct electricity, and turn blue litmus red.
- Basic solutions, like aqueous ammonia, also conduct electricity, are bitter tasting, feel slippery, and turn red litmus blue.
- Arrhenius: acids produce $\text{H}^+_{(\text{aq})}$ ions and bases produce $\text{OH}^-_{(\text{aq})}$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

Brønsted-Lowry Theory

- 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.
- $\text{H}_2\text{O}(\text{l}) \leftrightarrow \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq})$
- 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 $[H^+_{(aq)}] = [OH^-_{(aq)}] = 1.0 \times 10^{-7} \text{ mol/L}$

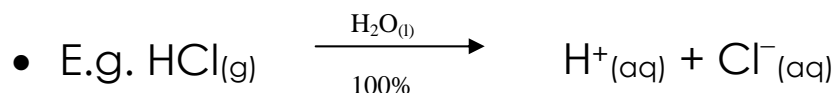
Therefore $K_w = 1.0 \times 10^{-14}$ (at SATP)

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

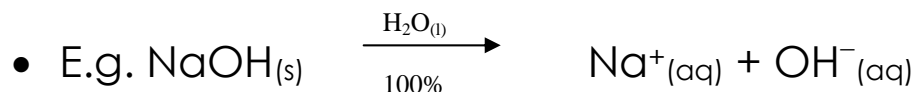
- A strong acid is an acid that ionizes completely in water to form hydrogen ions.



- 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 4.7×10^{-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 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)

Measuring pH

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

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