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 $H^+_{(aq)}$ ions and bases produce $OH^-_{(aq)}$ ions.
- Brønsted-Lowry: a proton is transferred from one reactant to the other.

Brønsted-Lowry Theory

(H^+)

- A Brønsted-Lowry acid is a proton donor and a Brønsted-Lowry base is a proton acceptor.
- E.g. $H_2O_{(l)} + HCl_{(g)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$ *(water forms the hydronium ion)*
  $H_2O_{(l)}$ is the Brønsted-Lowry base and $HCl_{(g)}$ Brønsted-Lowry acid.
- E.g. $NH_{3(g)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$ *(water forms the hydroxide ion)*
  $H_2O_{(l)}$ is the Brønsted-Lowry acid and $NH_{3(g)}$ 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)

Conjugate pair

Conjugate pair

Acetic Acid is a weak acid and mostly stays in its non ionic form

- E.g. $\underset{acid}{HC_2H_3O_{2(aq)}} + \underset{base}{H_2O_{(l)}} \leftrightarrow \underset{base}{C_2H_3O_2^-{}_{(aq)}} + \underset{acid}{H_3O^+{}_{(aq)}}$

Acid Base Conjugate base Conjugate acid

H^+ H^+

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.
- $H_2O_{(l)} \leftrightarrow H^+{}_{(aq)} + OH^-{}_{(aq)}$ 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.

$$\frac{[H^+{}_{(aq)}][OH^-{}_{(aq)}]}{[H_2O_{(l)}]} = K_e$$

$$H_2O_{(l)} + \cancel{H_2}O \leftrightarrow \cancel{H_3O}^+ + OH^-$$

H^+

  - Since the $[H_2O]$ is constant we include it in K_e to get…

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

in pure water the concentration of $[H^+_{(aq)}] = [OH^-_{(aq)}] = 1.0 \times 10^{-7}$ 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.
- E.g. $HCl_{(g)} \xrightarrow[100\%]{H_2O_{(l)}} H^+_{(aq)} + Cl^-_{(aq)}$
- 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.
- E.g. $NaOH_{(s)} \xrightarrow[100\%]{H_2O_{(l)}} Na^+_{(aq)} + OH^-_{(aq)}$
- Strong bases: group 1 and 2 hydroxides.

Hydrogen Ion Concentration and pH

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

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

$$[H^+] \xrightarrow{-\log} pH$$

$$pH \xrightarrow{10^{-pH}} [H^+]$$

pOH and pK_w

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

$pH = 4$

$[H^+] = 1 \times 10^{-4}$

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.

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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