Strong Acids and Bases
C020

Before discussing acids and bases examine the concept of chemical equilibrium

At reaction is at equilibrium when it is proceeding forward and backwards at the same rate, and the concentrations of all species involved remain constant

For the following reaction:

$$aA + bB \rightleftharpoons cC + dD \qquad K_{eq} = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

Here, [X] rersents the concentration of reactant A (usually in molarity; moles of solute in a litre of solvent; mol L^{-1}) and K_{eq} is the equilibrium constant at a given temperature

The value of K_{eq} indicates the relative position of the equilibrium; a large K_{eq} means the reaction favours products while a small K_{eq} means the reaction favours reactants.

Note that solids do not appear in K_{eq} expressions because their concentrations (densities) are constant

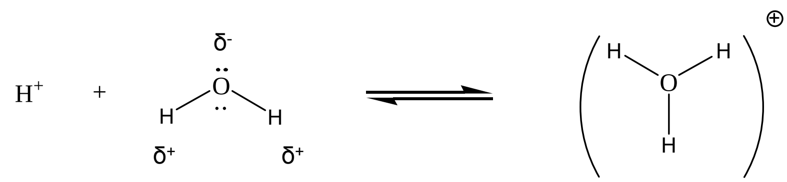
Start with a General Definition:

BrΦnsted Acid

A BrΦnsted definition of an acid is “any substance that can donate a H^+ to another substance”.

$$H - A + H_2O \leftrightarrow H_3O^+(aq) + A^-(aq)$$

In water, the association of H^+ with water is so extensive that the solvated species is written H_3O^+ (aq) where H_3O^+ is the hydronium ion.



Hydronium ion

All molecules with a $H^{\delta+}$ are capable of acting as acids (e.g. $H^{\delta+}$-$O^{\delta-}$R; $H^{\delta+}$-$X^{\delta-}$)

BrΦnsted Base

A BrΦnsted definition of an base is “any substance that is capable of accepting a H^{+}” from another substance; (by forming some sort of bond with it).

$$B^{-} + H^{+} \leftrightarrow HB$$

or even

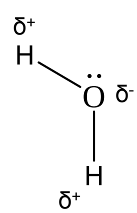
$$\ddot{B} + H^{+} \leftrightarrow HB^{+}$$

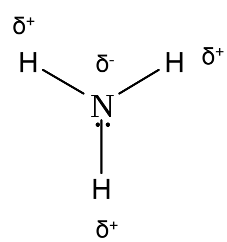
All anions are generally good bases:

For example: HO^{-} H_2N^{-} H_3C^{-}

Similarly so are most molecules containing atoms with a lone pair and a partial negative charge.

For example:





etc.

An acid capable of loosing more than 1 H is termed a polyprotic acid.

For example:
H_2SO_4 sulfuric acid

In general:

$$H_2A + H_2O \rightleftharpoons H_3O^+(aq) + HA^-(aq) \quad \text{(acid is } H_2A\text{)}$$

$$HA^- + H_2O \rightleftharpoons H_3O^+(aq) + A^{2-}(aq) \quad \text{(acid is } HA^-\text{)}$$

Such acids are diprotic

Phosphoric acid is an example of a triprotic acid

$$H_3PO_4 + H_2O \longrightarrow 3\ H^+(aq) + PO_4^{3-}(aq)$$

A base capable of accepting more than 1 H is termed a polyprotic base.

For example:
CO_3^{2-} carbonate anion

$$CO_3^{2-} + H_3O^+ \rightleftharpoons HCO_3^- + H_2O \quad \text{(base is } CO_3^{2-}\text{)}$$

$$HCO_3^- + H_3O^+ \rightleftharpoons H_2CO_3 + H_2O \quad \text{(base is } HCO_3^-\text{)}$$

It is not unusual for a single substance to be able to function as either an acid or a base.

These are called amphiprotic substances

For example:

$HSO_4^- + H_2O \rightleftharpoons H_3O^+ + SO_4^{2-}$ (acid is HSO_4^-)

$HSO_4^- + H_3O^+ \rightleftharpoons H_2SO_4 + H_2O$ (base is HSO_4^-)

Therefore, HSO_4^- is amphiprotic

The most common amphiprotic substance is H_2O which can autoionize:

$H_2O\ (\ell) + H_2O\ (\ell) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$

acid base

For this reaction, the equilibrium constant K is given by:

$$K = \frac{[H_3O^+][OH^-]}{[H_2O]^2}$$

But in dilute solution $[H_2O]^2$ ~ constant

Therefore:
$K[H_2O]^2 = K_W \equiv$ water ionization constant

At 25° C, $K_W = 1.008 \times 10^{-14} \sim 10^{-14}$

Therefore:
$K_W = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$

This leads to the concept of pH and pOH

Since in pure "neutral" water $[H_3O^+][OH^-] = 1.0 \times 10^{-14}$ and $[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ mol L^{-1}

$\Rightarrow$ water is arbitrarily accepted as the measure of neutrality for defining acids and base.

If $[H_3O^+] > 1.0 \times 10^{-7}$ M, the solution is said to be acidic

If $[H_3O^+] < 1.0 \times 10^{-7}$ M, the solution is said to be basic

$[H_3O^+]$ and $[OH^-]$ are usually given as simple numbers instead of exponents by using the following conversions:

$$pH = -\log_{10}[H_3O^+]$$

and

$$pOH = -\log_{10}[OH^-]$$

Example:

Given $[H_3O^+] = 0.00357$ M

$\therefore$ $pH = -\log_{10}(3.57 \times 10^{-3}) = 2.45$

Given $[OH^-] = 3.0 \times 10^{-7}$ M

$\therefore$ $pOH = -\log_{10}(3.0 \times 10^{-7}) = 6.52$

You can work backwards too.

For example:

If pH = 2.49

$\Longrightarrow$ $[H_3O^+] = 10^{-2.49} = 3.24 \times 10^{-3}$ M

If pOH = 9.08

$\Longrightarrow$ $[OH^-] = 10^{-9.08} = 8.32 \times 10^{-10}$ M

For neutral water at equilibrium

$[H_3O^+] = [OH^-] = 1.0 \times 10^{-7}$ M

$\therefore$ pH = pOH = 7.0

$\because$ $K_W = [H_3O^+][OH^-]$ $\Longrightarrow$ $-\log_{10}(K_W) = -\log_{10}([H_3O^+][OH^-])$

$= -\log_{10}([H_3O^+]) - \log_{10}([OH^-])$

$\therefore$ pK_W = pH+pOH = 14

Note: pH or pOH can be negative for very high concentrations

Example:

What will be the pH and pOH of a solution where 0.73 g HCl are dissolved In 100 mL of H_2O?

HCl is a strong acid and therefore fully dissociated.

$\Longrightarrow$ $[H_3O^+]=[HCl]_o$ = initial concentration of hydrochloric acid

$$\therefore \quad [HCl]_o = \frac{0.73g}{(0.1L)(36.5g\,mol^{-1})} = 0.20M$$

$\therefore$ pH = $-\log_{10}(0.2)$ = 0.7

and

pOH = 14-0.7 = 13.3

Concentration of Hydrogen ions compared to distilled water		Examples of solutions at this pH
10,000,000	pH= 0	Battery acid, Strong Hydrofluoric Acid
1,000,000	pH = 1	Hydrochloric acid secreted by stomach lining
100,000	pH = 2	Lemon Juice, Gastric Acid Vinegar
10,000	pH = 3	Grapefruit, Orange Juice, Soda
1,000	pH = 4	Acid rain Tomato Juice
100	pH = 5	Soft drinking water Black Coffee
10	pH = 6	Urine Saliva
1	pH = 7	"Pure" water
1/10	pH = 8	Sea water
1/100	pH = 9	Baking soda
1/1,000	pH = 10	Great Salt Lake Milk of Magnesia
1/10,000	pH = 11	Ammonia solution
1/100,000	pH = 12	Soapy water
1/1,000,000	pH =13	Bleaches Oven cleaner
1/10,000,000	pH = 14	Liquid drain cleaner

Conjugate Acid/Base Pairs

Acids and base reactions are complementary reactions;
that is, they must always occur together

For example:

$$\overset{\delta^+}{H}\text{-}\overset{\delta^-}{Cl} + \overset{\delta^-}{N}\text{-}\overset{\delta^+}{H_3} \rightleftharpoons NH_4^+ + Cl^-$$

acid, base ⇌ conjugate acid, conjugate base

In general:

$$HA + B \rightleftharpoons BH^+ + A^-$$

acid, base ⇌ conjugate acid, conjugate base

After an acid loses a proton it becomes a substance capable of accepting the proton back again;
that is, an acid is transformed into its conjugate base by loss of a proton.

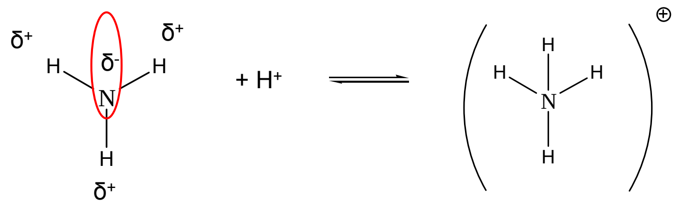
Similarly, after a base accepts a proton it becomes a substance capable of donating a proton;
that is, a base is transformed into is conjugate acid by gain of a proton.

Example:

For each of the following:

i) State whether the substance is an acid, base or amphiprotic
ii) Underline the acidic H or encircle the basic atom
iii) Draw the structure of the conjugate acid/base

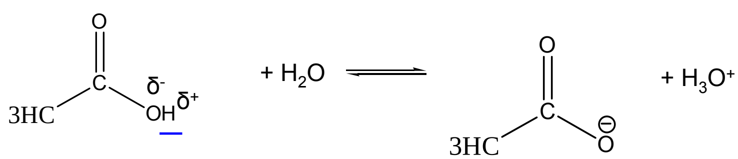
a)



base

conjugate acid

b)

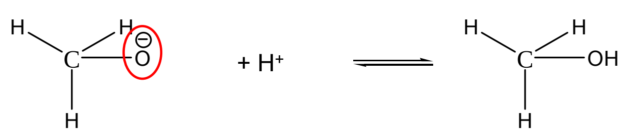


acid

conjugate base

Strong Acids and Bases 19

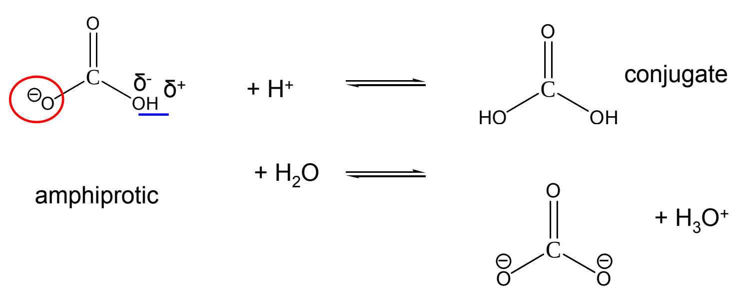
c)



base

conjugate acid

d)



conjugate acid

amphiprotic

conjugate base

Strong Acids and Bases 20

Strong Acids and Bases

(Weak acids and Bases later)

For the ionization of all acids; that is:

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

the concentration of all species at equilibrium is given by:

$$K = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$

and since in aqueous solution $[H_2O]$~ constant, one writes:

$K[H_2O]=K_a$ = acid ionization constant

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

K_a is a measure of acid strength
and $pK_a = -\log_{10}K_a$

If at equilibrium $[H_3O^+][A^-] >> [HA]$ then $K_a >> 1$ and $pK_a << 1$

The acid is then said to be a **strong acid**

Strong acids are fully dissociated in water.

There are six common strong acids:

HCl, HBr, HI, HNO_3, $HClO_4$ and H_2SO_4 (first ionization only)

All strong acids HA will dissociate **completely** to H_3O^+ and A^-

$\therefore$ $[H_3O^+]$ = initial concentration of HA = $[HA]_o$

For the ionization of all bases:

$$B^- + H_2O \rightleftharpoons BH + OH^-$$

The concentration of all species at equilibrium is given by:

$$K = \frac{[BH][OH^-]}{[B^-][H_2O]}$$

Again since $[H_2O]$ ~ constant:

$$K[H_2O] = K_b = \frac{[BH][OH^-]}{[B^-]}$$

where K_b = base ionization constant and $pK_b = -\log_{10}K_b$

Strong bases correspond to the case where $[BH][OH^-] >> [B^-]$ and $K_b >> 1$

All species containing OH^-, and molecules with atoms having a negative charge will be strong bases.

For example: LiOH, NaOH, KOH, $Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$, H3C-, H_2N^-, etc.

All halogen ions, $X^- = F^-, Cl^-, Br^-, I^-$, and neutral atoms with a lone pair of electrons will be weaker bases

A strong base B will react **completely** with water to form BH and OH^-

∴ at equilibrium $[OH^-]$ = initial concentration of base = $[B]_o$

Acids and bases react to form a salt and water

For example: $HA + NaOH \rightarrow Na^{+}A^{-} + H_2O$

pH Indicators

Many natural substances change color when they change from acid to conjugate base

These substances can be used as **indicators** of acidic and basic solutions

For example: let an indicator molecule = HIn

$$HIn + H_2O \rightleftharpoons H_3O^{+} + In^{-}$$

Color A Color B

In acid solution the indicator will show color A and in basic solution, color B

The change in color occurs over a range of pH values (usually ≤ 2) but can occur quite suddenly when the pH changes rapidly during a titration.