Important Notice

Group 1: You are in the Tutorial: Stoichiometry 1 and 2

Group 2: You are doing the lab: Synthesis

Section 003

Student Number	Initials	
250377285	N. G.	
250417065	Z. Mc.	
250419357	Z. Mo.	
250418455	T. N.	
250184967	H. R.	**Find Jan or Sandy**
250382992	M. S.	**asap in the Chem. labs**
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Section 006

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250343897	Y. X.

Review Quiz Section 006

- **Class Average:** 52.6 % (186 wrote)
- **Breakdown:**

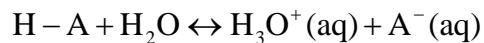A^+: 1% A: 6% B: 7% C: 17% D: 27%
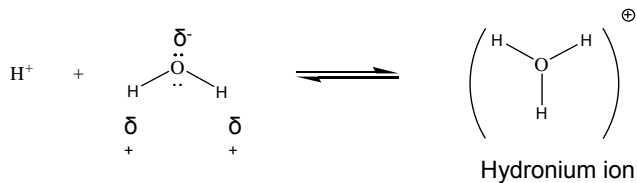F: 42%

Strong Acids and Bases C020

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.
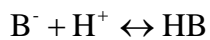
Hydronium ion

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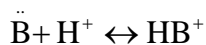
BrΦnsted Base

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:

$$\delta^+$$
$$H$$
$$\ddot{O}\ \delta^-$$
$$H$$
$$\delta^+$$

$$\delta^+$$
$$H\ \ \delta^-\ \ H\ \ \delta^+$$
$$N$$
$$H$$
$$\delta^+$$

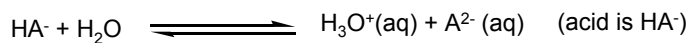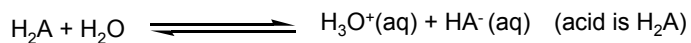etc.
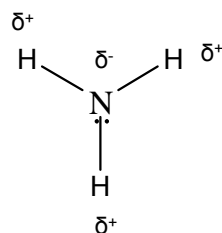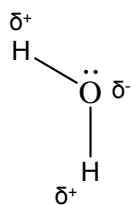
An acid capable of loosing more than 1 H is termed a **polyprotic acid**.

For example:
H_2SO_4 sulfuric acid

In general:

$H_2A + H_2O \rightleftharpoons H_3O^+(aq) + HA^- (aq)$ (acid is H_2A)

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

Such acids are **di**protic

Phosphoric acid is an example of a **tri**protic acid

$H_3PO_4 + H_2O \longrightarrow 3\ H^+(aq) + PO_4^{3-}(aq)$
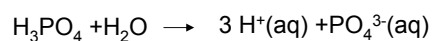
A base capable of accepting more than 1 H is termed a **polyprotic base**.

For example:
CO_3^{2-} carbonate anion

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

$$HCO_3^- + H_3O^+ \; \rightleftharpoons \; H_2CO_3 + H_2O \qquad \text{(base is } HCO_3^-\text{)}$$
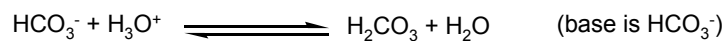
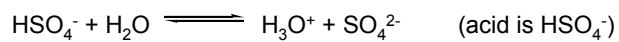
It is not unusual for a single substance to be able to function as either an acid or a base.

These are called **amphiprotic** substances

For example:
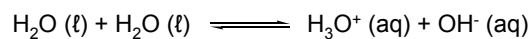
$$HSO_4^- + H_2O \; \rightleftharpoons \; H_3O^+ + SO_4^{2-} \qquad \text{(acid is } HSO_4^-\text{)}$$
$$HSO_4^- + H_3O^+ \; \rightleftharpoons \; H_2SO_4 + H_2O \qquad \text{(base is } HSO_4^-\text{)}$$

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 $[H_2O]^2 \sim$ 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

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

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

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

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You can work backwards too.

For example:

If pH = 2.49

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

If pOH = 9.08

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

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For neutral water at equilibrium

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

$\therefore$ pH = pOH = 7.0

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

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

$\therefore$ $\boxed{pK_W = pH + pOH \quad = 14}$

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

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

$\therefore$ 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

Concentration	pH	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 Vineger
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

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^-

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

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 = $[B]_o$