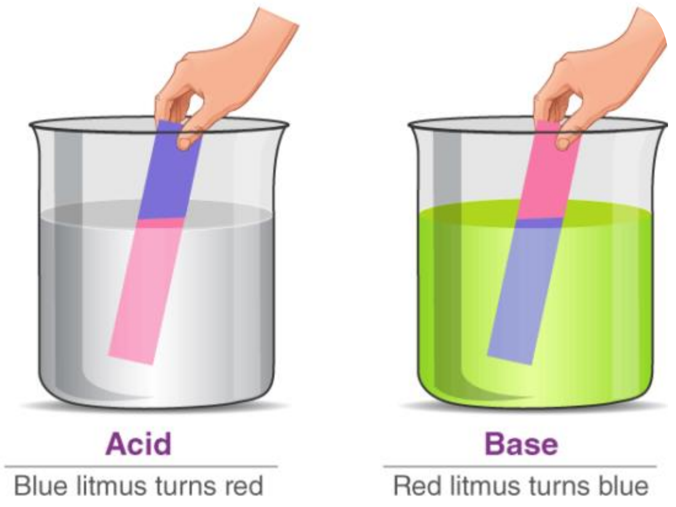
Acids and Bases

Dr. Md. Mahbub Alam
Department of Chemistry
BUET, Dhaka-1000



The contents of this presentation is prepared to provide a brief idea about the topics, details will be discussed in the classes.
Contents have been collected from multiple textbooks and internet.

Arrhenius Concept of Acids and Bases

- *The earliest acid-base definition, which classifies these substances in terms of their behavior in water.*
- *An acid is a substance with H in its formula that dissociates to yield H_3O^+; Example: HCl, H_2SO_4 etc.*
- *A base is a substance with OH in its formula that dissociates to yield OH^-. Example: NaOH, $Ca(OH)_2$ etc.*
- *When an acid reacts with a base, they undergo neutralization:*

$$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$$

Lewis Concept of Acids and Bases

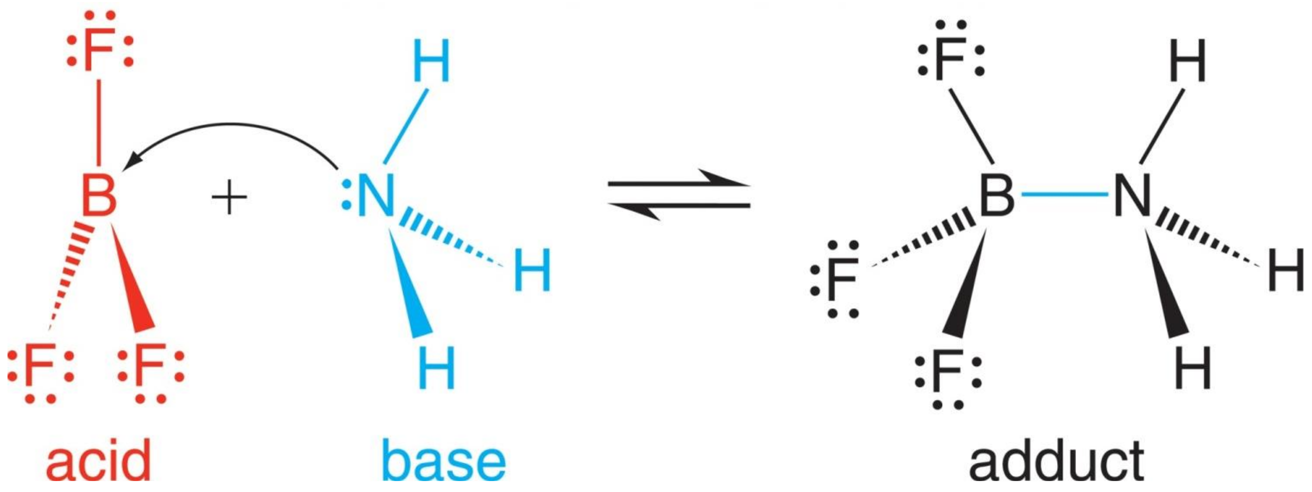
- *A Lewis base is any species that **donates an electron pair** to form a bond.*
- *A Lewis acid is any species that **accepts an electron pair** to form a bond.*

$$\underset{\text{base}}{B{:}} + \underset{\text{acid}}{H^+} \rightleftharpoons \underset{\text{adduct}}{B\text{—}H^+}$$

- *The Lewis definition views an acid-base reaction as the **donation and acceptance** of an electron pair to form a **covalent bond**.*
- *A Lewis base must have **a lone pair of electrons** to donate.*

Electron-Deficient Molecules as Lewis Acids

- *B and Al often form electron-deficient molecules, and these atoms have an unoccupied p orbital that can accept a pair of electrons.*
- *BF_3 accepts an electron pair from ammonia to form a covalent bond.*



Brønsted-Lowry Concept of Acids and Bases

- *An acid is a proton donor, any species that donates an H+ ion.*
- *An acid must contain H in its formula.*
- *A base is a proton acceptor, any species that accepts an H+ ion.*
- *An acid-base reaction is a proton-transfer process.*
- *A Brønsted acid is a proton donor and a Brønsted base is a proton acceptor.*

$$\underset{\text{base}}{NH_3(aq)} + \underset{\text{acid}}{H_2O(l)} \rightleftharpoons \underset{\text{conjugate acid}}{NH_4^+(aq)} + \underset{\text{conjugate base}}{OH^-(aq)}$$

Conjugate Acid-Base Pairs

NH_3 *accepts* an H^+ to form NH_4^+.

$$H_2S + NH_3 \rightleftharpoons HS^- + NH_4^+$$

H_2S *donates* an H^+ to form HS^-.

- *H_2S and HS^- are a **conjugate acid-base** pair. HS^- is the conjugate base of the acid H_2S.*
- *NH_3 and NH_4^+ are a **conjugate acid-base** pair. NH_4^+ is the conjugate acid of the base NH_3.*
- *A Brønsted-Lowry acid-base reaction occurs when an acid and a base react to form their conjugate base and conjugate acid, respectively.*

Strong and Weak Acids

- A strong acid dissociates completely into ions in water:

$$HA(g \text{ or } l) + H_2O(l) \rightarrow H_3O+(aq) + A-(aq)$$

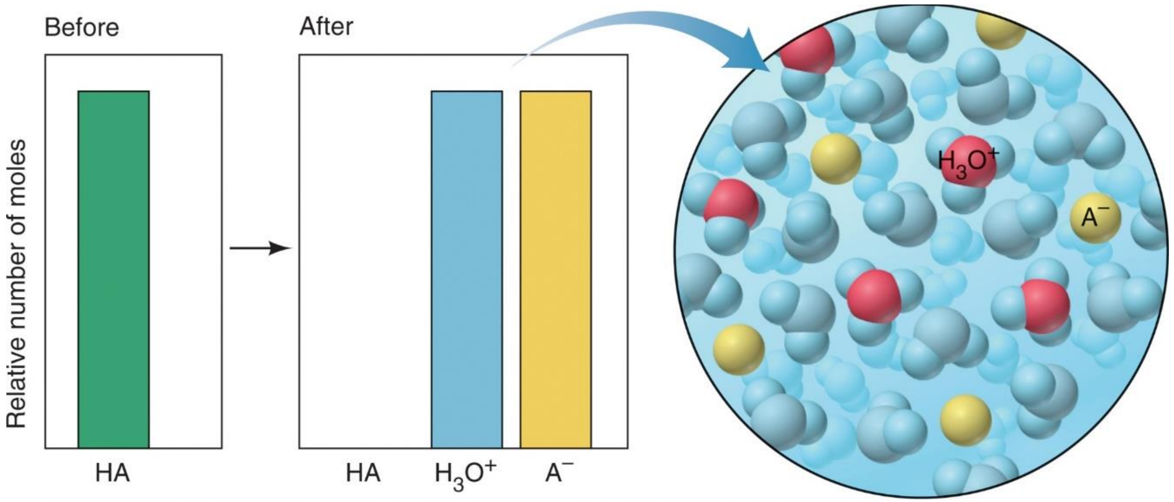
- A dilute solution of a strong acid contains no HA molecules.
- A weak acid dissociates slightly to form ions in water:

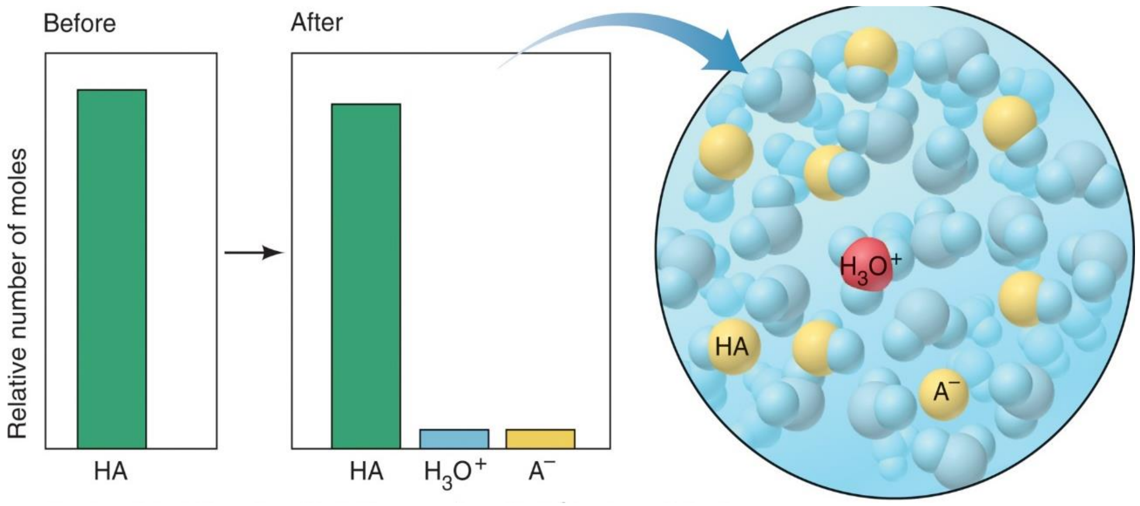
$$HA(aq) + H2O(l) \rightarrow \quad H_3O+(aq) + A-(aq)$$

- In a dilute solution of a weak acid, most HA molecules are undissociated.

$$K_c = \frac{[H_3O^+][A^-]}{[HA][H_2O]}$$ has a very **small** value.

Strong and Weak Acids





Strong acid: $HA(g \text{ or } l) + H_2O(l) \rightarrow H_3O^+(aq) + A^-(aq)$

There are no HA molecules in solution.

Weak acid: $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$

Most HA molecules are undissociated.

Strong and Weak Acids

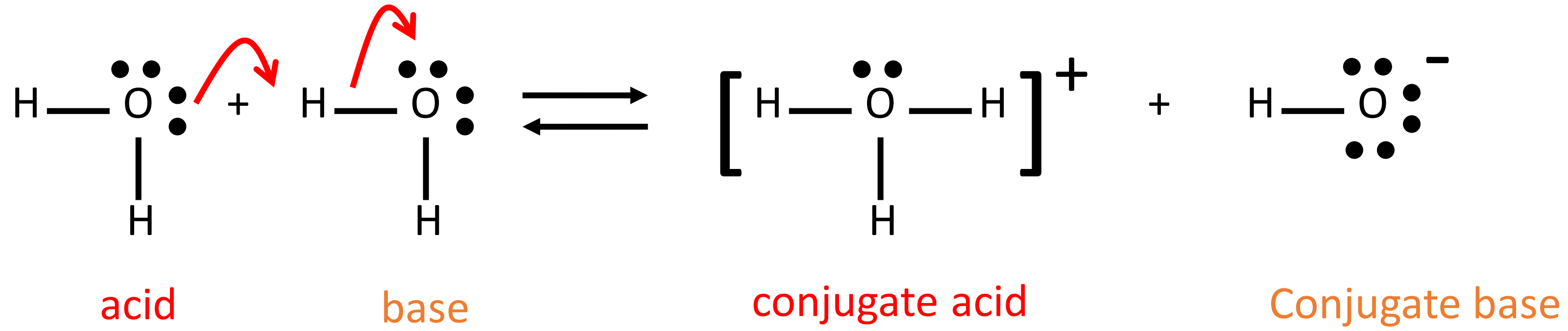
Acid strength increases ↑

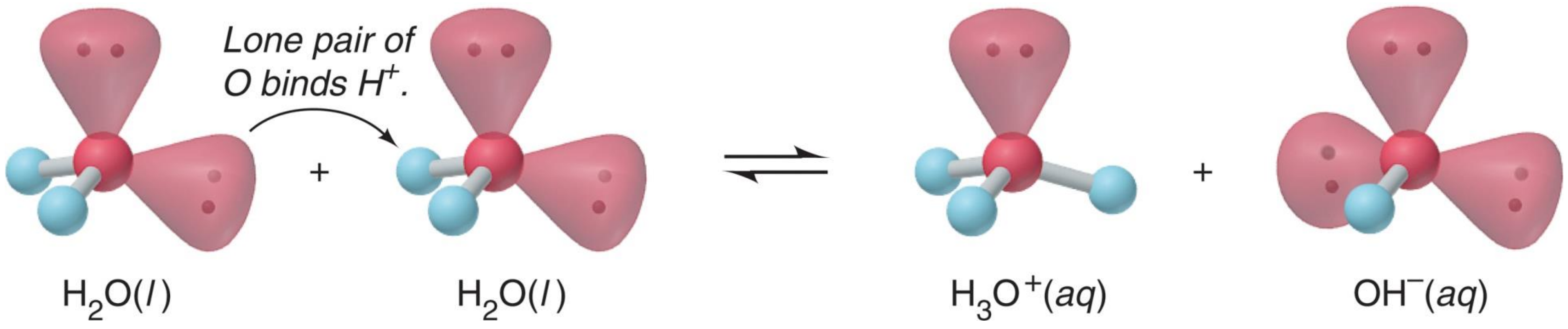
Acid	Conjugate Base
Strong acids	
$HClO_4$ (perchloric acid)	ClO_4^- (perchlorate ion)
HI (hydroiodic acid)	I^- (iodide ion)
HBr (hydrobromic acid)	Br^- (bromide ion)
HCl (hydrochloric acid)	Cl^- (chloride ion)
H_2SO_4 (sulfuric acid)	HSO_4^- (hydrogen sulfate ion)
HNO_3 (nitric acid)	NO_3^- (nitrate ion)
H_3O^+ (hydronium ion)	H_2O (water)
Weak acids	
HSO_4^- (hydrogen sulfate ion)	SO_4^{2-} (sulfate ion)
HF (hydrofluoric acid)	F^- (fluoride ion)
HNO_2 (nitrous acid)	NO_2^- (nitrite ion)
HCOOH (formic acid)	$HCOO^-$ (formate ion)
CH_3COOH (acetic acid)	CH_3COO^- (acetate ion)
NH_4^+ (ammonium ion)	NH_3 (ammonia)
HCN (hydrocyanic acid)	CN^- (cyanide ion)
H_2O (water)	OH^- (hydroxide ion)
NH_3 (ammonia)	NH_2^- (amide ion)

Base strength increases ↓

Acid-Base Properties of Water

- ***Water can act either as an acid or as a base donating or accepting proton.***
- ***This is also sometimes called the autoionization or self ionization of water.***





The Ion Product of Water

$$2H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + OH^-\ (aq)$$

$$H_2O\ (l) \rightleftharpoons H^+\ (aq) + OH^-\ (aq)$$

$$K_c = \frac{[H_3O^+][OH^-]}{[H_2O]^2} \qquad \text{Or, } K_c = \frac{[H^+][OH^-]}{[H_2O]^2} \qquad \text{Or, } K_c[H_2O]^2 = [H^+][OH^-]$$

- *Because only a very small fraction of water molecules are ionized, the concentration of water, $[H_2O]$, remains virtually unchanged. Then we can replace $K_c[H_2O]^2$ by K_w (the ion-product constant).*

$$K_w = [H^+][OH^-]$$

The Ion Product of Water

- *In pure water at 25°C, the concentrations of H^+ and OH^- ions are equal and found to be $[H^+] = 1.0 \times 10^{-7}$ M and $[OH^-] = 1.0 \times 10^{-7}$ M.*

$$K_w = [H^+][OH^-] = (1.0 \times 10^{-7})\,(1.0 \times 10^{-7}) = 1.0 \times 10^{-14}$$

$$[H^+][OH^-] = 1.0 \times 10^{-14}$$

Taking negative logarithm in both sides,

$$-\log [H^+] - \log [OH^-] = -\log (1.0 \times 10^{-14})$$

$$\text{or, } pH + pOH = 14$$

pH – A Measure of Acidity

$$pH = -\log [H^+]$$

Solution type		At 25^0C	
neutral	$[H^+] = [OH^-]$	$[H^+] = 1 \times 10^{-7}$	$pH = 7$
acidic	$[H^+] > [OH^-]$	$[H^+] > 1 \times 10^{-7}$	$pH < 7$
basic	$[H^+] < [OH^-]$	$[H^+] < 1 \times 10^{-7}$	$pH > 7$

pH – A Measure of Acidity

$$pH = -\log [H^+]$$

pH Meter

pH paper

The pHs of Some Common Fluids

Sample	pH Value
Gastric juice in the stomach	1.0–2.0
Lemon juice	2.4
Vinegar	3.0
Grapefruit juice	3.2
Orange juice	3.5
Urine	4.8–7.5
Water exposed to air*	5.5
Saliva	6.4–6.9
Milk	6.5
Pure water	7.0
Blood	7.35–7.45
Tears	7.4
Milk of magnesia	10.6
Household ammonia	11.5

Weak Acids and Acid Ionization Constants

$$HA\ (aq) + H_2O\ (l) \rightleftharpoons H_3O^+\ (aq) + A^-\ (aq)$$

$$HA\ (aq) \rightleftharpoons H_3O^+\ (aq) + A^-\ (aq)$$

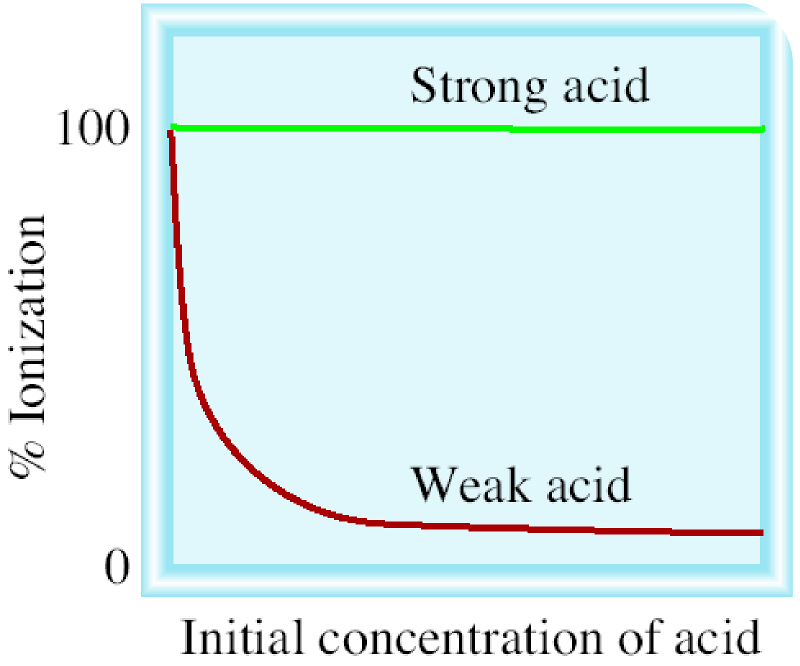
$$K_a = \frac{[H^+][A^-]}{[HA]}$$

K_a is the *acid ionization constant*.

Another measure of the strength of an acid is its percent of ionization.

$$\text{percent ionization} = \frac{\text{ionized acid concentration at equilibrium}}{\text{initial concentration of acid}} \times 100\%$$

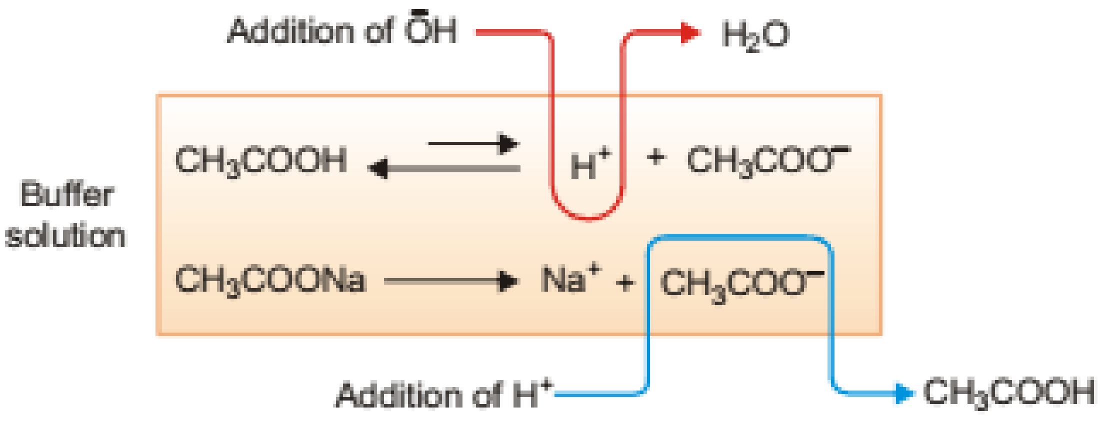
$$\text{percent ionization} = \frac{[H^+]}{[HA]_0} \times 100\%$$

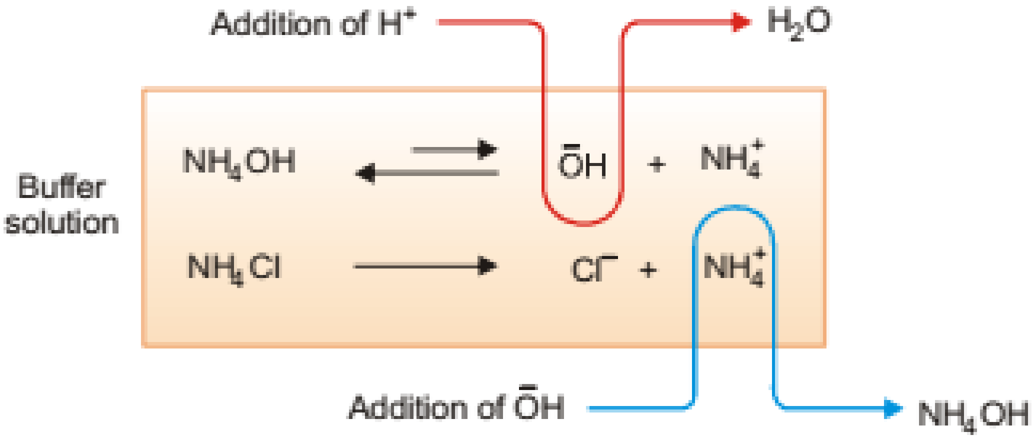


Buffer Solution

- *A buffer solution has the ability to resist changes in pH upon the addition of small amounts of either acid or base.*
- *A buffer solution must contain –*
  - ✓ *A weak acid or a weak base and*
  - ✓ *The salt of the weak acid or weak base*
- *Two types –*
  - ✓ *Acid buffers – $CH_3COOH + CH_3COONa$*
  - ✓ *Basic buffers – $NH_4OH + NH_4Cl$*

How Does a Buffer Solution Work?





Thank You