Acids And Bases

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Introduction:

- The concept of acids and bases is fundamental to chemistry.
- Over time, several definitions have been proposed to explain acid-base behavior under different conditions (aqueous, gaseous, non-aqueous, etc.).

Three Major Theories:

Theory	Acid Definition	Base Definition	Scope
Arrhenius	Produces H^+ in water	Produces OH^- in water	Limited to aqueous solutions
Brønsted-Lowry	Proton donor	Proton acceptor	Broader (not limited to water)
Lewis	Electron pair acceptor	Electron pair donor	Most general; includes non-protonic reactions

- We'll explain Brønsted-Lowry in detail in the next topic, so let's focus here on an overview and understanding of acid/base strength, behavior, and properties.

General Characteristics

1. Acids

- Taste sour (in edible substances).
- Turn blue litmus paper red.
- React with metals to release hydrogen gas.
- React with bases in neutralization reactions to form salts and water.
- Have pH < 7 in aqueous solutions.

2. Bases

- Taste bitter and feel slippery (soapy).
- Turn red litmus paper blue.
- React with acids to neutralize them.
- Often contain hydroxide ions (OH^-) or accept protons.
- Have pH > 7 in aqueous solutions.

Acid and Base Strength

A. Strong Acids:

- Completely dissociate in water.
- **Examples:** HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4 (first proton)

B. Weak Acids:

- Partially dissociate.
- **Example:** CH_3COOH (acetic acid), HF, H_2CO_3

C. Strong Bases:

- Fully dissociate into OH^-.
- **Examples:** NaOH, KOH, $Ba(OH)_2$

D. Weak Bases:

- Partially accept protons (often produce OH^- indirectly).
- **Examples:** NH_3, CH_3NH_2

Quantifying Strength

A. For Acids:

- **Ka** = Acid dissociation constant
  - $HA \rightleftharpoons H^+ + A^-$
  - $Ka = [H^+][A^-]/[HA]$
- **pKa = –log(Ka)**
  - Lower pKa = Stronger acid
  - Higher pKa = Weaker acid

B. For Bases:

- **Kb** = Base dissociation constant
  - $B + H_2O \rightleftharpoons BH^+ + OH^-$
  - $Kb = [BH^+][OH^-]/[B]$
- **pKb = –log(Kb)**
  - Lower pKb = Stronger base
- **Relation**: pKa + pKb = 14 (at 25°C)

Acid-Base Reactions:

- **Neutralization**:
  - Acid + Base → Salt + Water
  - Example: HCl + NaOH → NaCl + H_2O
- **Acid-Base Equilibrium**:

- Favor the formation of the weaker acid-base pair.
- Proton transfers are governed by relative strengths (ΔpKa).

Conjugate Acid-Base Pairs:

- Every acid has a conjugate base (what remains after it donates a proton).
- Every base has a conjugate acid (what forms after it accepts a proton).
- Example:
  - Acid: HCl → Conjugate base: Cl^-
  - Base: NH_3 → Conjugate acid: NH_4^+
- Stronger acid → weaker conjugate base, and vice versa.

Amphoteric Substances:

- Compounds that can act as both acid and base, depending on the environment.
- Example: Water (H_2O)
  - Acts as base with HCl (accepts H^+)
  - Acts as acid with NH_3 (donates H^+)

Acid-Base Properties in Non-Aqueous Solvents:

- Acid-base behavior can differ outside water.
- In liquid ammonia, for example:
  - NH_4^+ is an acid
  - NH_2^- is a base
- This shows the importance of solvent system in defining acid-base strength and reactions.

Leveling Effect:

- In water, strong acids like HCl, HNO_3, and H_2SO_4 all appear to have the same strength because they completely dissociate, forming H_3O^+.
- Water levels the strength of all strong acids to that of H_3O^+.
- The same applies to strong bases and OH^- in water.

Importance in Chemistry:

- Determines solubility, reactivity, and equilibrium in chemical systems.
- Central to titrations, buffer solutions, reaction mechanisms, and biological functions (e.g., enzyme activity).

Brønsted–Lowry Theory (Proton Transfer Theory)

Definition:

- **Brønsted–Lowry Acid**: A substance that donates a proton (H^+).
- **Brønsted–Lowry Base**: A substance that accepts a proton (H^+).
- This theory extends beyond aqueous solutions and is more general than Arrhenius theory.

Key Concepts:

A. Acid-Base Reaction as Proton Transfer:

- **General form:**

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

  - HA: acid (proton donor)
  - B: base (proton acceptor)
  - A^-: conjugate base
  - HB^+: conjugate acid

- **Example:**

$$HCl + H_2O \rightarrow Cl^- + H_3O^+$$

  - HCl is the Brønsted acid
  - H_2O is the Brønsted base
  - Cl^- is the conjugate base
  - H_3O^+ is the conjugate acid

B. Conjugate Acid-Base Pairs:

- Always exist in **pairs**.
- **Acid ↔ Conjugate base** (loses a proton)
- **Base ↔ Conjugate acid** (gains a proton)
- Example:

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

  - NH_3 is base, NH_4^+ is its conjugate acid
  - H_2O is acid, OH^- is its conjugate base

Strength Relationship:

- Strong acid → Weak conjugate base
- Weak acid → Strong conjugate base
- **Example:**
  - HCl (strong acid) ⇌ Cl^- (very weak base)

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