

# AP Chemistry— CHAPTER 16 STUDY GUIDE— Acid-Base Equilibrium

## 16.1 Acids and Bases: A Brief Review

- Acids taste sour and cause certain dyes to change color.
- Bases taste bitter and feel soapy.
- Arrhenius concept of acids and bases:
  - An acid is a substance that, when dissolved in water, increases the concentration of  $\text{H}^+$  ions.
  - Example:  $\text{HCl}$  is an acid.
  - An Arrhenius base is a substance that, when dissolved in water, increases the concentration of  $\text{OH}^-$  ions.
  - Example:  $\text{NaOH}$  is a base.
- This definition is quite narrow in scope as it limits us to aqueous solutions.

## 16.2 Brønsted-Lowry Acids and Bases

- We can use a broader, more general definition for acids and bases that is based on the fact that acid-base reactions involve proton transfers.

### The $\text{H}^+$ Ion in Water

- The  $\text{H}^+(\text{aq})$  ion is simply a proton with no surrounding valence electrons.
- In water, clusters of hydrated  $\text{H}^+(\text{aq})$  ions form.
- The simplest cluster is  $\text{H}_3\text{O}^+(\text{aq})$ .
  - We call this a **hydronium ion**.
  - Larger clusters are also possible (such as  $\text{H}_5\text{O}_2^+$  and  $\text{H}_9\text{O}_4^+$ ).
- Generally we use  $\text{H}^+(\text{aq})$  and  $\text{H}_3\text{O}^+(\text{aq})$  interchangeably.

### Proton-Transfer Reactions

- We will focus our attention on  $\text{H}^+(\text{aq})$ .
- According to the Arrhenius definitions, an acid increases  $[\text{H}^+]$  and a base increases  $[\text{OH}^-]$ .
- Another definition of acids and bases was proposed by Brønsted and Lowry.
- In the Brønsted-Lowry system, a **Brønsted-Lowry acid** is a species that donates  $\text{H}^+$  and a **Brønsted-Lowry base** is a species that accepts  $\text{H}^+$ .
  - Therefore a Brønsted-Lowry base does not need to contain  $\text{OH}^-$ .
    - $\text{NH}_3$  is a Brønsted-Lowry base, but not an Arrhenius base.
- Consider  $\text{NH}_3(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{OH}^-(\text{aq})$ :
  - $\text{H}_2\text{O}$  donates a proton to ammonia.
  - Therefore, water is acting as an acid.
  - $\text{NH}_3$  accepts a proton from water.
  - Therefore, ammonia is acting as a base.
  - An **amphiprotic substance** can behave either as an acid or as a base.
  - Thus, water is an example of an amphiprotic species.

### Conjugate Acid-Base Pairs

- Whatever is left of the acid after the proton is donated is called its conjugate base.
- Similarly, a conjugate acid is formed by adding a proton to the base.
- Consider  $\text{HX}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{X}^-(\text{aq})$ :
  - $\text{HX}$  and  $\text{X}^-$  differ only in the presence or absence of a proton.
  - They are said to be a **conjugate acid-base pair**.
  - $\text{X}^-$  is called the **conjugate base**.
- After  $\text{HX}$  (acid) loses its proton it is converted into  $\text{X}^-$  (base).
  - Therefore  $\text{HX}$  and  $\text{X}^-$  are a conjugate acid-base pair.
- After  $\text{H}_2\text{O}$  (base) gains a proton it is converted into  $\text{H}_3\text{O}^+$  (acid).
  - $\text{H}_3\text{O}^+$  is the **conjugate acid**.
  - Therefore,  $\text{H}_2\text{O}$  and  $\text{H}_3\text{O}^+$  are a conjugate acid-base pair.

### Relative Strengths of Acids and Bases

- The stronger an acid is, the weaker its conjugate base will be.
- We can categorize acids and bases according to their behavior in water.
  1. *Strong acids* completely transfer their protons to water.
    - No undissociated molecules remain in solution.
    - Their conjugate bases have negligible tendencies to become protonated.
    - An example is  $\text{HCl}$ .

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