Advance Warning

Quiz in Week 12:
- Gases (mainly last two lectures)
- Kinetics
- Equilibrium
- Acids/Bases (from first three lectures)

So Gases and Kinetics problem sets should now be finished.
* This week’s tutes - Equilibrium and Gases

Equilibria In Acids & Bases

In water: an acid (e.g., HCl) ionises to produce H⁺ (aq)
- actually H₃O⁺ (aq), but we usually just write H⁺ (aq)

Conjugate Acid-base Pairs

- NH₄⁺ is the conjugate acid of NH₃
- NH₃ is the conjugate base of NH₄⁺

A conjugate base has one less proton than its conjugate acid
- HSO₄⁻: conjugate base is SO₄²⁻
  conjugate acid is H₂SO₄
- H₂SO₄ is dibasic or diprotic acid:
  H₂SO₄ + H₂O ↔ H₃O⁺ + SO₄²⁻
  (lies ~100% to right)

Definitions

- Arrhenius:
  H⁺ + OH⁻ ⇔ H₂O
  - ACID:
  - BASE:

- Brønsted-Lowry:
  H⁺ + A⁻ ⇔ HA
  - ACID:
  - BASE:

- Lewis:
  A + B ⇔ A:B
  - ACID:
  - BASE:

Examples

Write the formula of the conjugate bases Write the formula of the conjugate acids

- H₃O⁺
  OH⁻
- H₂SO₄
  H₂O
- HClO₄
  CN⁻
- CH₃COOH
  NH₃
- HPO₄²⁻
  HPO₄³⁻

Acid-Base Reactions (see Silberberg)

Table 18.4 The Conjugate Pairs in Some Acid-Base Reactions

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Acid</th>
<th>Base</th>
<th>Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>Reaction 1</td>
<td>HF</td>
<td>H₂O</td>
<td>H₃O⁺</td>
</tr>
<tr>
<td>Reaction 2</td>
<td>HCOOH</td>
<td>CN⁻</td>
<td>HCOO⁻</td>
</tr>
<tr>
<td>Reaction 3</td>
<td>NH₄⁺</td>
<td>CO₃²⁻</td>
<td>NH₃</td>
</tr>
<tr>
<td>Reaction 4</td>
<td>H₃PO₄</td>
<td>OH⁻</td>
<td>H₃PO₄²⁻</td>
</tr>
<tr>
<td>Reaction 5</td>
<td>H₂SO₄</td>
<td>NH₃</td>
<td>H₂SO₄⁻²</td>
</tr>
<tr>
<td>Reaction 6</td>
<td>H₃PO₄²⁻</td>
<td>SO₄²⁻</td>
<td>PO₄³⁻</td>
</tr>
</tbody>
</table>

Autoionisation of Water

H₂O (l) ↔ H⁺ (aq) + OH⁻ (aq)

- Equilibrium constant given special symbol:
  Kₖw = [H⁺][OH⁻]
  NB: [H₂O (l)]: actually activity, which is = 1

At 25 °C: Kₖw = 1 × 10⁻¹⁴

- Neutral solution: [H⁺] = [OH⁻] = 1 × 10⁻⁷ mol L⁻¹
- Acidic solution: [H⁺] > 1 × 10⁻⁷ M
- Basic: [H⁺] < 1 × 10⁻⁷ M
Autoionisation of Water

\[
pH = - \log_{10}[H^+] \\
pOH = - \log_{10}[OH^-] \\
pK_w = - \log_{10}[K_w] = 14 \text{ at } 25 \degree C
\]

Acid: \( pH < 7 \)
Neutral: \( pH = 7 \)
Basic: \( pH > 7 \)

The ‘p’ Convention

\[
\text{pH} + \text{pOH} = 14 \\
\text{pOH} = 14 - \text{pH}
\]

Strong Acids & Bases

- Completely ionise in water:
  - e.g. HCl \( \rightarrow H^+ + Cl^- \)
  - equilibrium lies completely to right, \( K_c = \infty \)

Strong acids
- \( H_2SO_4, HCl, HBr, HI, HNO_3, HClO_4 \)

Strong bases
- All hydroxides of Groups 1 & 2 (except Be): NaOH, Ca(OH)$_2$, ...

HF : NOT strong acid !
because H–F bond stronger than O–H

Weak Acids and Bases

- Any acid or base not on the list of strong ones is weak
  - it does not completely ionise in water.
    - e.g., acetic (ethanoic) acid, \( CH_3CO_2H \) (HA for short):
      \[
      \text{HA} \leftrightarrow H^+ + A^- \\
      \]

\[
K_a = [H^+] [A^-] \\
[HA]
\]

for acetic acid, \( K_a = 10^{-5.7} \) M
\( pK_a = 4.7 \)

pH of 0.1 M solution of acetic acid > 1
(pH would be \(-\log(0.1) = 1\) only if it were completely ionised)

Examples

- What is the pH of a 0.1 M HCl solution?
- What is the pH of a 0.002 M NaOH solution?

More Examples

- Calculate the pH of:
  0.001 M HNO$_3$
  0.001 M NaOH
  0.001 M Ca(OH)$_2$

- What is the pH of a solution formed by mixing 400 mL of 0.05 M HCl with 600 mL of 0.05 M NaOH?

- What is the [H$^+$] of a solution with a pH of 4.5?

Comparing Strong and Weak Acids

Strong acid: HA\( (aq) + H_2O \rightarrow H_3O^+ (aq) + A^- (aq) \)
Weak acid: HA\( (aq) \rightarrow H_2O + A^- (aq) \)