Chemistry 1002

Assistance - administrative
• First Year Enquiry Office (10 am - 3.15 pm)
• E-mail: firstyear@chem.usyd.edu.au

Assistance - Course Work
• Duty Tutor Room (Tuesday – Friday, various times)
• E-mail: dutytutor@chem.usyd.edu.au

Information and Resources
• First Year Chemistry web site (http://firstyear.chem.usyd.edu.au/)
• WebCT
• ChemCAL
• On-line laboratory pre-work
Chemistry 1002

Tutorials
• Start in week 2
• Tutorial quizzes held in weeks 3 and 6

Laboratory Work
• Start in week 2

Assessment
• 10 % laboratory assessment
• 15 % tutorial quizzes (4 per semester; the best three marks used)
• 75 % 3 hour exam at the end of semester

Staff-Student Liaison Committee
Equilibria In Acids & Bases

In water: an acid (e.g., HCl) ionises to produce $\text{H}^+ (\text{aq})$

- actually $\text{H}_3\text{O}^+ (\text{aq})$, but we usually just write $\text{H}^+ (\text{aq})$
Definitions

- **Arrhenius**: \( H^+ + OH^- \rightleftharpoons H_2O \)
  - ACID:
  - BASE:

- **Brønsted - Lowry**: \( H^+ + A^- \rightleftharpoons HA \)
  - ACID:
  - BASE:

- **Lewis**: \( A + :B \rightleftharpoons A:B \)
  - ACID:
  - BASE:
Brønsted-Lowry Model

- acid = proton donor
- base = proton acceptor
Conjugate Acid-base Pairs

- NH$_4^+$ is the conjugate acid of NH$_3$
- NH$_3$ is the conjugate base of NH$_4^+$

**A conjugate base has one less proton than its conjugate acid**

- HSO$_4^-$: conjugate **base** is SO$_4^{2-}$
  - conjugate **acid** is H$_2$SO$_4$

- H$_2$SO$_4$ is a **dibasic** or **diprotic** acid:
  
  $$
  \text{H}_2\text{SO}_4 + \text{H}_2\text{O} \iff \text{H}_3\text{O}^+ + \text{HSO}_4^- \\
  \text{HSO}_4^- + \text{H}_2\text{O} \iff \text{H}_3\text{O}^+ + \text{SO}_4^{2-}
  $$

  (lies $\sim$100% to right)
### Examples

<table>
<thead>
<tr>
<th>Write the formula of the conjugate bases</th>
<th>Write the formula of the conjugate acids</th>
</tr>
</thead>
<tbody>
<tr>
<td>$\text{H}_3\text{O}^+$</td>
<td>$\text{OH}^-$</td>
</tr>
<tr>
<td>$\text{H}_2\text{SO}_4$</td>
<td>$\text{H}_2\text{O}$</td>
</tr>
<tr>
<td>$\text{HClO}_4$</td>
<td>$\text{CN}^-$</td>
</tr>
<tr>
<td>$\text{CH}_3\text{COOH}$</td>
<td>$\text{NH}_3$</td>
</tr>
<tr>
<td>$\text{HPO}_4^{2-}$</td>
<td>$\text{HPO}_4^{2-}$</td>
</tr>
</tbody>
</table>
# Acid-Base Reactions

## The Conjugate Pairs in Some Acid-Base Reactions

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Acid</th>
<th>+</th>
<th>Base</th>
<th>↔</th>
<th>Base</th>
<th>+</th>
<th>Acid</th>
</tr>
</thead>
<tbody>
<tr>
<td>1</td>
<td>HF</td>
<td>+</td>
<td>H₂O</td>
<td></td>
<td>F⁻</td>
<td>+</td>
<td>H₃O⁺</td>
</tr>
<tr>
<td>2</td>
<td>HCOOH</td>
<td>+</td>
<td>CN⁻</td>
<td></td>
<td>HCOO⁻</td>
<td>+</td>
<td>HCN</td>
</tr>
<tr>
<td>3</td>
<td>NH₄⁺</td>
<td>+</td>
<td>CO₃²⁻</td>
<td></td>
<td>NH₃</td>
<td>+</td>
<td>HCO₃⁻</td>
</tr>
<tr>
<td>4</td>
<td>H₂PO₄⁻</td>
<td>+</td>
<td>OH⁻</td>
<td></td>
<td>HPO₄²⁻</td>
<td>+</td>
<td>H₂O</td>
</tr>
<tr>
<td>5</td>
<td>H₂SO₄</td>
<td>+</td>
<td>N₂H₅⁺</td>
<td></td>
<td>HSO₄⁻</td>
<td>+</td>
<td>N₂H₆²⁺</td>
</tr>
<tr>
<td>6</td>
<td>HPO₄²⁻</td>
<td>+</td>
<td>SO₃²⁻</td>
<td></td>
<td>PO₄³⁻</td>
<td>+</td>
<td>HSO₃⁻</td>
</tr>
</tbody>
</table>
Autoionisation of Water

\[ \text{H}_2\text{O} (\ell) \rightleftharpoons \text{H}^+ (\text{aq}) + \text{OH}^- (\text{aq}) \]

- Equilibrium constant given special symbol:
  \[ K_w = [\text{H}^+][\text{OH}^-] \]

  NB: [H$_2$O (\ell)]: actually activity, which is = 1

At 25 °C: \( K_w = 1.0 \times 10^{-14} \)

- Neutral solution: \([\text{H}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ mol L}^{-1}\)
- Acidic solution  \([\text{H}^+] > 1.0 \times 10^{-7} \text{ M}\)
- Basic : \([\text{H}^+] < 1.0 \times 10^{-7} \text{ M}\)
Autoionisation of Water
The ‘p’ Convention

\[ \text{pH} = -\log_{10}[H^+] \]

\[ \text{pOH} = -\log_{10}[OH^-] \]

\[ pK_w = -\log_{10}[K_w] = 14 \text{ at } 25 \, ^\circ\text{C} \]

Acid: \( \text{pH} < 7 \)

Neutral: \( \text{pH} = 7 \)

Basic: \( \text{pH} > 7 \)

Since \( K_w = [H^+][OH^-] \):

\[ \log_{10} K_w = \log_{10} [H^+] + \log_{10} [OH^-] \]

\[ -\log_{10} [H^+] - \log_{10} [OH^-] = -\log_{10} K_w \]

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

\[ \text{pOH} = 14 - \text{pH} \]
A pH scale showing the range from acidic to basic. The pH scale is labeled from 0 to 14, with 7 being neutral. At pH 0, the concentration of hydrogen ions $[H^+]$ is $10^{-0}$. As the pH increases, the concentration of $[H^+]$ decreases, and vice versa. The diagram also includes examples of substances at different pH levels:

- **Acidic** (low pH): Stomach Acid, Lemon Juice, Vinegar, 1 M HCl
- **Basic** (high pH): 1 M NaOH, Household Ammonia, Sea Water, Blood, Pure Water, Milk
- **Neutral** (pH 7): Whole Blood
Revision: Strong Acids & Bases

Completely ionise in water:

e.g. \( \text{HCl} \rightleftharpoons \text{H}^+ + \text{Cl}^- \)

equilibrium lies completely to right, \( K_c \approx \infty \)

**Strong acids**
\( \text{H}_2\text{SO}_4, \text{HCl}, \text{HBr}, \text{HI}, \text{HNO}_3, \text{HClO}_4 \)

**Strong bases**
All hydroxides of Groups 1 & 2 (except Be):
\( \text{NaOH}, \text{Ca(OH)}_2, \ldots \)

HF : NOT strong acid !!
because H–F bond stronger than O–H
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 HNO₃
- 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?
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, CH₃CO₂H (HA for short):

  \[
  \text{HA} \rightleftharpoons \text{H}^+ + \text{A}^{-}
  \]

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

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

  \( pK_a = 4.7 \)

  pH of 0.1 M solution of acetic acid \( > 1 \)

  (pH would be \( -\log(0.1) = 1 \) only if it were completely ionised)
Comparing Strong and Weak Acids

**Strong acid**: HA(g or l) + H₂O(l) → H₃O⁺(aq) + A⁻(aq)

**Weak acid**: HA(aq) + H₂O(l) ⇌ H₃O⁺(aq) + A⁻(aq)