Acids, Bases and Equilibrium

Chapter 9
Educational Goals

1. Given a chemical equation, write the law of mass action.
2. Given the equilibrium constant ($K_{eq}$) for a reaction, predict whether the reactants or products are predominant.
3. Use Le Châtelier’s Principle to explain how a chemical reaction at equilibrium responds when a change is made to the concentration of reactant or product.
4. Know the definitions of Bronsted-Lowry acids and bases.
5. Given the acid form or the base form of a conjugate pair, identify its conjugate.
6. List the properties of acidic and basic solutions.
7. Understand the term "acid strength," and know how acid strength is related to the acidity constant ($K_a$) value.
8. Given the $[H_3O^+]$, be able to calculate the $[OH^-]$ (and vice versa).
9. Given the $[H_3O^+]$, be able to calculate the pH (and vice versa).
10. Given the $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, or pH, be able to characterize a solution as being **acidic**, **basic**, or **neutral**.

11. Given the reactants, predict the products of a **neutralization reaction**.

12. Given the pH of a solution and the p$K_a$ for a particular acid, determine the relative concentrations of the acid and base forms of the conjugate pair.

13. Define a **buffer**, and describe how a buffer solution is made.
In earlier chapters, we assumed all reactions converted reactants into products until one of the reactants was completely used up.

This is often not the case since many reactions are reversible.

Products can be converted back to reactants!
Consider the Following Reaction:

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \]

Colorless Gas \quad Brown Gas

This reaction is reversible!
The rate of forward reaction is equal to the rate of reverse reaction when the reaction has reached equilibrium.

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \]

Colorless Gas \hspace{2cm} Brown Gas

<table>
<thead>
<tr>
<th>Time</th>
<th>Amount Present</th>
</tr>
</thead>
<tbody>
<tr>
<td></td>
<td>( \text{N}_2\text{O}_4 )</td>
</tr>
<tr>
<td></td>
<td>Colorless Gas</td>
</tr>
<tr>
<td></td>
<td>( \text{NO}_2 )</td>
</tr>
</tbody>
</table>
At equilibrium the rate of the forward and reverse reactions are the same and the concentrations (amounts) of reactants and products do not change.
Equilibrium Constants

If $\text{N}_2\text{O}_4(g)$ and $\text{NO}_2(g)$ are allowed to reach equilibrium and the equilibrium concentrations of each are measured, then the following will always be true:

$$K_{eq} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 4.6 \times 10^{-3} \text{ M}$$

**Concentrations in units of molarity**

**Value of the equilibrium constant for this reaction** (determined by experiment)
Law of Mass Action

For any chemical reaction at equilibrium:

$$aA + bB \rightleftharpoons cC + dD$$

The equilibrium concentrations are given by:

$$K_{eq} = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

**Law of Mass Action**
Law of Mass Action

Since the pressure of each gas is proportional to its concentration:

\[ P = \frac{n}{V} \, RT \]

The Law of Mass Action can also be written in terms of the pressures of reactants and products for gas phase reactions:

\[ K_p = \frac{(P_C)^c (P_D)^d}{(P_A)^a (P_B)^b} \]

**Partial pressures of reactants and products**
Law of Mass Action

• Pure liquids or solids do not appear in the equilibrium equation.

• Equilibrium constant values vary with temperature.

• Examples used in this chapter will assume a temperature of 25°C.
Example

Balance the reaction equation and then write the corresponding equilibrium constant expression.

$$2\text{CO}(g) + \text{O}_2(g) \rightleftharpoons 2\text{CO}_2(g)$$

$$K_{eq} = \frac{[\text{CO}_2]^2}{[\text{CO}]^2 \, [\text{O}_2]}$$
You try it

Write the equilibrium constant expression for this reaction.

\[ \text{H}_2(g) + \text{CO}_2(g) \rightleftharpoons \text{CO}(g) + \text{H}_2\text{O}(g) \]
When an equilibrium constant has a value less than 1, the denominator (related to reactant concentration) of the equilibrium constant expression is greater than the numerator, which means that at equilibrium the reactants are favored (more reactants than products).

We say that the equilibrium favors the reactants when $K_{eq} < 1$.
When an equilibrium constant has a value greater than 1, the numerator (related to product concentration) of the equilibrium constant expression is greater than the denominator, which means that at equilibrium the products are favored (more products than reactants).

We say that the equilibrium favors the products when $K_{eq} > 1$.
Think about it:

In which reaction below are *products* favored by equilibrium?

1. **HI** + **H₂O** ⇌ **I⁻** + **H₃O⁺

   \[
   K_{eq} = \frac{[I^-][H_3O^+]}{[HI]} = 2.5 \times 10^{10} \text{ M}
   \]

2. **HF** + **H₂O** ⇌ **F⁻** + **H₃O⁺

   \[
   K_{eq} = \frac{[F^-][H_3O^+]}{[HF]} = 6.6 \times 10^{-4} \text{ M}
   \]
Le Châtelier’s Principle
The response to a loss of equilibrium is predicted by: \textit{Le Châtelier’s Principle}, which states that when a reversible reaction is pushed out of equilibrium, the reaction responds to reestablish equilibrium.
Or......think of it like this

The response to a loss of equilibrium is predicted by:

**Le Châtelier’s Principle**, which states that *the chemical system will try to un-do what you did!*
Le Châtelier’s Principle

- Consider the our reaction from earlier:

\[ \text{N}_2\text{O}_4(g) \rightleftharpoons 2\text{NO}_2(g) \]

At equilibrium the following equation applies:

\[ K_{eq} = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]} = 4.6 \times 10^{-3} \text{ M} \]

If the amount of one species is changed, the amounts of other species must change in order for the equilibrium constant to remain constant!
Another Example of Le Châtelier’s Principle

\[ \text{H}_2\text{O}(l) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{CO}_3(aq) \]

1. Increasing \([\text{CO}_2]\) upsets the equilibrium and a net forward → reaction takes place.

Net Forward Reaction means:
The rate of reactants going to products is fastest until the law of mass action is once again satisfied.

- Forward means “reaction proceeds to the right”
Another Example of Le Châtelier’s Principle

\[ \text{H}_2\text{O}(l) + \text{CO}_2(g) \rightleftharpoons \text{H}_2\text{CO}_3(aq) \]

2. Decreasing [CO\(_2\)] upsets the equilibrium and a net reverse \( \leftarrow \) reaction takes place.

3. Increasing [H\(_2\)CO\(_3\)] upsets the equilibrium and a net reverse reaction takes place.

Net Reverse Reaction means:
The rate of \textit{products} going to \textit{reactants} is fastest until the law of mass action is once again satisfied.

- \textit{Reverse means “reaction proceeds to the left”}
Group Work:

For the equilibrium reaction shown below, predict which reaction (forward or reverse) will be the faster one until equilibrium is reestablished when:

\[2\text{SO}_2(g) + \text{O}_2(g) \rightleftharpoons 2\text{SO}_3(g)\]

a. [SO\(_2\)] is increased  

b. [SO\(_2\)] is decreased  
c. [SO\(_3\)] is decreased  
d. [O\(_2\)] is increased
Catalysts do not Effect Equilibrium

- Catalysts increase reaction rates by lowering the activation energy.

- When a catalyst is used in a reversible reaction, the lowered activation energy speeds up both the forward and reverse reactions.

- The overall result is that a catalyst has no effect on an equilibrium or on the value of $K_{eq}$. 
Acids and Bases
Some Common Acids and Bases

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>Uses</th>
</tr>
</thead>
<tbody>
<tr>
<td><strong>Acids</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Hydroiodic acid</td>
<td>HI</td>
<td>Disinfectant</td>
</tr>
<tr>
<td>Hydrobromic acid</td>
<td>HBr</td>
<td>Veterinary sedative</td>
</tr>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>Household cleaning products, swimming pool maintenance, and metal cleaning</td>
</tr>
<tr>
<td>(muriatic acid)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sulfuric acid</td>
<td>H₂SO₄</td>
<td>Fertilizers, explosives, dyes, and glues</td>
</tr>
<tr>
<td>Nitric acid</td>
<td>HNO₃</td>
<td>Fertilizers, explosives, and dyes</td>
</tr>
<tr>
<td><strong>Bases</strong></td>
<td></td>
<td></td>
</tr>
<tr>
<td>Sodium hydroxide</td>
<td>NaOH</td>
<td>Drain cleaners, soap manufacture</td>
</tr>
<tr>
<td>Potassium hydroxide</td>
<td>KOH</td>
<td>Paint and varnish removers</td>
</tr>
</tbody>
</table>
Properties of Acids

• Acids will **dissolve** some metals

• Acids will turn a plant pigment (blue litmus) from *blue to red*.

• Acids taste *sour*
Properties of Bases

• Bases feel slippery or soapy.

• Bases will turn a plant pigment (red litmus) from red to blue.

• Bases taste bitter
Bronsted-Lowry Acids and Bases
Bronsted-Lowry Acids and Bases

A Bronsted-Lowry acid releases $\text{H}^+$.

A Bronsted-Lowry base accepts $\text{H}^+$.
Bronsted-Lowry Acids and Bases

\[ \text{HCl} + \text{H}_2\text{O} \rightleftharpoons \text{Cl}^- + \text{H}_3\text{O}^+ \]  
\( \text{acid} \quad \text{base} \)

\[ \text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^- \]  
\( \text{base} \quad \text{acid} \)
Example:

For the forward direction of the reaction, identify the Bronsted-Lowry acid and base.

\[ \text{HCN} + \text{CO}_3^{2-} \rightleftharpoons \text{HCO}_3^- + \text{CN}^- \]
For the forward direction of the reaction, identify the Bronsted-Lowry acid and base.

You try it:

\[
\text{HSO}_4^- + \text{HPO}_4^{2-} \rightleftharpoons \text{H}_2\text{SO}_4 + \text{PO}_4^{3-}
\]
Conjugates

Compounds, such as HCN and CN\(^-\) or NH\(_3\) and NH\(_4\)\(^+\), which differ only in the presence or absence of H\(^+\) are called conjugates.

Another example of conjugates is acetic acid and acetate ion:

\[
\text{CH}_3\text{C}–\text{OH} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{C}–\text{O}^- + \text{H}_3\text{O}^+
\]
Possible Misconceptions

You can call this the:
- “Acid”
- “Conjugate Acid”
- “Acid Form”

CH₃C–OH + H₂O ⇌ CH₃C–O⁻ + H₃O⁺

You can call this the:
- “Base”
- “Conjugate Base”
- “Base Form”
Amphoteric Compounds

• Compounds that can act as acids or as bases are called amphoteric.

• Examples:
  – \( \text{H}_2\text{O} \) (water)
    • \( \text{H}_2\text{O} + \text{CN}^- \rightarrow \text{OH}^- + \text{HCN} \) (water as an acid)
    • \( \text{H}_2\text{O} + \text{HCl} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^- \) (water as a base)

  – \( \text{HCO}_3^- \) (hydrogen carbonate)
    • \( \text{HCO}_3^- + \text{CN}^- \rightarrow \text{CO}_3^{2-} + \text{HCN} \) (\( \text{HCO}_3^- \) as an acid)
    • \( \text{HCO}_3^- + \text{HCl} \rightarrow \text{H}_2\text{CO}_3 + \text{Cl}^- \) (\( \text{HCO}_3^- \) as a base)
Another way to write acid-base equations:

• These are equivalent:

1) \( \text{HCN} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{CN}^- \)

2) \( \text{HCN} \rightarrow \text{H}^+ + \text{CN}^- \)

In #2 above, we write \( \text{H}^+ \) even though we know \( \text{H}^+ \) only exist in water as \( \text{H}_3\text{O}^+ \)
Ionization of Water

2 H₂O ⇌ OH⁻ + H₃O⁺
Ionization of Water

\[ 2 \text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{H}_3\text{O}^+ \]

Ionization of water obeys the law of mass action just like **EVERY** other chemical reaction!!

\[ K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2 \]
Ionization of Water

\[ 2 \text{H}_2\text{O} \rightleftharpoons \text{OH}^- + \text{H}_3\text{O}^+ \]

That means whenever we know \([\text{OH}]\), we can calculate the concentration of \([\text{H}_3\text{O}^+]\)

\[ K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2 \]

\[ [\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14} \text{ M}^2}{[\text{OH}^-]} \]
Ionization of Water

\[ 2 \text{H}_2\text{O} \leftrightharpoons \text{OH}^- + \text{H}_3\text{O}^+ \]

Also, whenever we know \([\text{H}_3\text{O}^+]\) we can calculate the concentration of \([\text{OH}^-]\)

\[ K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2 \]

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{M}^2}{[\text{H}_3\text{O}^+]} 
\]
Example:
What is the concentration of [OH\(^-\)] in an aqueous solution when [H\(_3\)O\(^+\)] = 1.0 \times 10^{-3} \text{ M}:

**Strategy:**

\[ K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \]

\[
[\text{OH}^-] = \frac{1.0 \times 10^{-14} \text{M}^2}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14} \text{M}^2}{1.0 \times 10^{-3} \text{ M}} = 1.0 \times 10^{-11} \text{ M}
\]
What is the concentration of \([H_3O^+]\) in an aqueous solution when \([OH^-] = 5.0 \times 10^{-11} \text{ M}\)?

**Strategy:**

\[ K_w = [OH^-][H_3O^+] = 1.0 \times 10^{-14} \text{ M}^2 \]

\[ = 2.0 \times 10^{-4} \text{ M} \]
Group Work

What is the concentration of [OH\(^-\)] in an aqueous solution when:

a. \([H_3O^+] = 1.0 \times 10^{-3} \text{ M}\)
b. \([H_3O^+] = 1.0 \times 10^{-7} \text{ M}\)
c. \([H_3O^+] = 5.0 \times 10^{-11} \text{ M}\)

Strategy:

\[K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14} \text{ M}^2\]
The pH Scale
The pH Scale

• The “p” in pH means “take logarithm then multiply by (-1)”

• pH means “take logarithm of the $\text{H}_3\text{O}^+$ ion concentration, then multiply by (-1)”

• $\text{pH} = -\log[\text{H}_3\text{O}^+]$
Quick Review of Logs

The log(10) is 1 because \(10^1 = 10\)
The log(100) is 2 because \(10^2 = 100\)
The log(1000) is 3 because \(10^3 = 1000\)
The log(.1) is -1 because \(10^{-1} = .1\)
The log(.01) is -2 because \(10^{-2} = .01\)
The log(3.4) is .53 because \(10^{(.53)} = 3.4\)

Most of the time we need our calculators
Review of Logs

Don’t forget about the web site that Dr. Perez make for helping chemistry 108 students with math!

You can link to it from our class webpage.
pH of a Solution:
What is the pH of an aqueous solutions with \([H_3O^+] = 6.3 \times 10^{-1}\)M?

Strategy:
\[
pH = - \log[H_3O^+]
\]

\[
pH = - \log [6.3 \times 10^{-1}] = -(-0.20) = 0.20
\]

NOTE: pH is one of the few *unit-less* numbers used in science!
You try it:
What is the pH of an aqueous solution with \([\text{OH}^-] = 6.3 \times 10^{-1}\text{M}\)?

Strategy:

\[
\text{pH} = -\log[\text{H}_3\text{O}^+] \\
K_w = [\text{OH}^-][\text{H}_3\text{O}^+] = 1.0 \times 10^{-14}
\]
Significant Figures in Logs

In a log value, the digits to the left of the decimal point **are not significant**!

\[ \text{pH} = - \log[H_3O^+] = -\log(1.6 \times 10^{-14}) = 13.80 \]

2 Significant Figures
The Antilog

What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with pH = 5.3?

Start with the definition of pH:

$$\text{pH} = -\log[\text{H}_3\text{O}^+]$$

$$5.3 = -\log[\text{H}_3\text{O}^+]$$

Multiply both sides by (-1)

$$-5.3 = \log[\text{H}_3\text{O}^+]$$

Take the antilog of both sides

$$10^{-5.3} = [\text{H}_3\text{O}^+]$$

Enter -5.3 in to your calculator, then find the $10^x$ key (antilog).

$$[\text{H}_3\text{O}^+] = 5 \times 10^{-6}$$
The Antilog

What is the $[\text{H}_3\text{O}^+]$ of an aqueous solution with $\text{pH} = 12.3$?
The pH Scale

We say a solution is **basic** when there is more $[\text{OH}^-]$ than $[\text{H}_3\text{O}^+]$.

We say a solution is **neutral** when $[\text{OH}^-] = [\text{H}_3\text{O}^+]$.

We say a solution is **acidic** when there is more $[\text{H}_3\text{O}^+]$ than $[\text{OH}^-]$.

<table>
<thead>
<tr>
<th>pH</th>
<th>$[\text{H}_3\text{O}^+]$</th>
<th>$[\text{OH}^-]$</th>
</tr>
</thead>
<tbody>
<tr>
<td>14</td>
<td>$1 \times 10^{-14}$</td>
<td>$1 \times 10^6$</td>
</tr>
<tr>
<td>13</td>
<td>$1 \times 10^{-13}$</td>
<td>$1 \times 10^{-1}$</td>
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<td>12</td>
<td>$1 \times 10^{-12}$</td>
<td>$1 \times 10^{-2}$</td>
</tr>
<tr>
<td>11</td>
<td>$1 \times 10^{-11}$</td>
<td>$1 \times 10^{-3}$</td>
</tr>
<tr>
<td>10</td>
<td>$1 \times 10^{-10}$</td>
<td>$1 \times 10^{-4}$</td>
</tr>
<tr>
<td>9</td>
<td>$1 \times 10^{-9}$</td>
<td>$1 \times 10^{-5}$</td>
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<tr>
<td>8</td>
<td>$1 \times 10^{-8}$</td>
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<td>$1 \times 10^{-7}$</td>
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<td>$1 \times 10^{-8}$</td>
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<td>$1 \times 10^{-5}$</td>
<td>$1 \times 10^{-9}$</td>
</tr>
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<td>$1 \times 10^{-4}$</td>
<td>$1 \times 10^{-10}$</td>
</tr>
<tr>
<td>3</td>
<td>$1 \times 10^{-3}$</td>
<td>$1 \times 10^{-11}$</td>
</tr>
<tr>
<td>2</td>
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<td>$1 \times 10^{-12}$</td>
</tr>
<tr>
<td>1</td>
<td>$1 \times 10^{-1}$</td>
<td>$1 \times 10^{-13}$</td>
</tr>
<tr>
<td>0</td>
<td>$1 \times 10^0$</td>
<td>$1 \times 10^{-14}$</td>
</tr>
<tr>
<td>Solution</td>
<td>pH</td>
<td></td>
</tr>
<tr>
<td>--------------------------------</td>
<td>--------</td>
<td></td>
</tr>
<tr>
<td>Battery acid</td>
<td>0.5</td>
<td></td>
</tr>
<tr>
<td>Gastric fluid in stomach</td>
<td>1.5</td>
<td></td>
</tr>
<tr>
<td>Soft drinks</td>
<td>2.0–4.0</td>
<td></td>
</tr>
<tr>
<td>Vinegar</td>
<td>2.4–3.4</td>
<td></td>
</tr>
<tr>
<td>Grapefruit juice</td>
<td>3.2</td>
<td></td>
</tr>
<tr>
<td>Urine</td>
<td>4.8–7.5</td>
<td></td>
</tr>
<tr>
<td>Rainwater (unpolluted)</td>
<td>6.2</td>
<td></td>
</tr>
<tr>
<td>Milk</td>
<td>6.3–6.6</td>
<td></td>
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<tr>
<td>Pure water</td>
<td>7.0</td>
<td></td>
</tr>
<tr>
<td>Blood</td>
<td>7.35–7.45</td>
<td></td>
</tr>
<tr>
<td>Soap</td>
<td>8.0–10.0</td>
<td></td>
</tr>
<tr>
<td>Antacids</td>
<td>10.5</td>
<td></td>
</tr>
<tr>
<td>Household ammonia</td>
<td>11.5</td>
<td></td>
</tr>
<tr>
<td>Lye (a 1 M NaOH solution)</td>
<td>14.0</td>
<td></td>
</tr>
</tbody>
</table>
Measuring pH:
A pH meter

Measuring pH:
Chemical Color Indicators

Change Colors according to pH
Measuring pH: Chemical Color Indicators

• Acids will turn a plant pigment (blue litmus) from blue to red.

• Bases will turn a plant pigment (red litmus) from red to blue.
Acid and Base Strength
• The stronger the acid, the more $\text{H}_3\text{O}^+$ it produces and the lower the pH of the solution that it forms.

• A strong acid such as HCl **dissociates** almost completely in water.

• $\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{Cl}^-(aq) + \text{H}_3\text{O}^+ (aq)$
When dealing with acids, the equilibrium constant \((K_{eq})\) is also known as the **acidity constant** \((K_a)\).

\[
\text{HCl}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{Cl}^- (aq) + \text{H}_3\text{O}^+ (aq)
\]

\[
K_a = \frac{[\text{Cl}^-][\text{H}_3\text{O}^+]}{[\text{HCl}]} = 1.0 \times 10^7
\]

\[
\text{HF}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{F}^- (aq) + \text{H}_3\text{O}^+ (aq)
\]

\[
K_a = \frac{[\text{F}^-][\text{H}_3\text{O}^+]}{[\text{HF}]} = 6.6 \times 10^{-4}
\]
Acid and Base Strength

- The larger the $K_a$ for an acid, the **stronger** an acid it is.

- Sometimes an acid's strength is tabulated by $pK_a$ ($pK_a = -\log K_a$).

- The **smaller** the $pK_a$, the **stronger** the acid.
  - $pK_a$ for HCl = -7.00
  - $pK_a$ for HF = 3.18
<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>$K_a$</th>
<th>$pK_a$</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>$1.0 \times 10^7$</td>
<td>-7.00</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>$H_3PO_4$</td>
<td>$7.5 \times 10^{-3}$</td>
<td>2.12</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>HF</td>
<td>$6.6 \times 10^{-4}$</td>
<td>3.18</td>
</tr>
<tr>
<td>Lactic acid</td>
<td>$CH_3CH(OH)CO_2H$</td>
<td>$1.4 \times 10^{-4}$</td>
<td>3.85</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>$CH_3CO_2H$</td>
<td>$1.8 \times 10^{-5}$</td>
<td>4.74</td>
</tr>
<tr>
<td>Carbonic acid</td>
<td>$H_2CO_3$</td>
<td>$4.4 \times 10^{-7}$</td>
<td>6.36</td>
</tr>
<tr>
<td>Dihydrogenphosphate ion</td>
<td>$H_2PO_4^-$</td>
<td>$6.2 \times 10^{-8}$</td>
<td>7.21</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>$NH_4^+$</td>
<td>$5.6 \times 10^{-10}$</td>
<td>9.25</td>
</tr>
<tr>
<td>Hydrocyanic acid</td>
<td>HCN</td>
<td>$4.9 \times 10^{-10}$</td>
<td>9.31</td>
</tr>
<tr>
<td>Hydrogencarbonate ion</td>
<td>$HCO_3^-$</td>
<td>$5.6 \times 10^{-11}$</td>
<td>10.25</td>
</tr>
<tr>
<td>Methylammonium ion</td>
<td>$CH_3NH_3^+$</td>
<td>$2.4 \times 10^{-11}$</td>
<td>10.62</td>
</tr>
<tr>
<td>Hydrogenphosphate ion</td>
<td>$HPO_4^{2-}$</td>
<td>$4.2 \times 10^{-13}$</td>
<td>12.38</td>
</tr>
</tbody>
</table>
Acid and Base Strength

• The stronger the base, the more OH\(^-\) it produces and the higher the pH of the solution that it forms.

• A strong base such as NaOH \textit{dissociates} (falls apart) almost completely in water.

\[
\text{NaOH}(s) \Leftrightarrow \text{Na}^+(aq) + \text{OH}^- (aq)
\]
Neutralizing Acids and Bases
Neutralization takes place when an acid and a base react to form water and a salt (an ionic compound).

- $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{O}(l)$

  strong acid   strong base   salt   water

- When a strong acid is neutralized by a strong base, the solution has a pH $= 7.0$
pH Buffers
Notation

Specific notation for an acid base reaction:
Example:

$$HCl(aq) + H_2O(l) \rightleftharpoons Cl^-(aq) + H_3O^+(aq)$$

General notation for an acid base reaction:

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

A$^-$ represents the conjugate base of any acid
An interesting relationship exists between pH and an acid and its conjugate base.

\[ \text{HA}_{(aq)} + \text{H}_2\text{O}_{(l)} \rightleftharpoons \text{A}^-_{(aq)} + \text{H}_3\text{O}^+_{(aq)} \]

For the acid (HA) and its conjugate base (A-),

- There is more HA when the pH is less than the pK_a.
- There is more A- when the pH is greater than the pK_a.
- There are equal amounts of HA and A- when the pH = pK_a.
Henderson-Hasselbalch Equation

\[ K_a = \frac{[H_3O^+][A^-]}{[HA]} \]

Rearrange to solve for \([H_3O^+]\)

\[ [H_3O^+] = \frac{K_a[HA]}{[A^-]} \]

Take the negative log of both sides

\[ -\log[H_3O^+] = -\log\left(\frac{K_a[HA]}{[A^-]}\right) \]

The negative log of translates to "p"

\[ pH = pK_a - \log\left(\frac{[HA]}{[A^-]}\right) \]

To make a negative log positive, take the inverse

\[ pH = pK_a + \log\left(\frac{[A^-]}{[HA]}\right) \]

Henderson-Hasselbalch Equation
Buffers

A buffer is a solution that resists change in pH when small amounts of acid or base are added.
Buffers

Example: Adding acid (H$_3$O$^+$) to an acetic acid/acetate buffer

\[
\text{CH}_3\text{CO}_2\text{H} + \text{H}_2\text{O} \rightleftharpoons \text{CH}_3\text{CO}_2^- + \text{H}_3\text{O}^+
\]

acetic acid \hspace{2cm} acetate ion

(weak acid) \hspace{2cm} (conjugate base)

If we add acid (H$_3$O$^+$), what does Le Chatelier say will happen?

- It is removed in the reverse reaction!
Buffers

Example: Adding base (OH-) to acetic acid/acetate buffer

\[
\text{CH}_3\text{CO}_2\text{H} + \text{OH}^- \rightleftharpoons \text{CH}_3\text{CO}_2^- + \text{H}_2\text{O}
\]

Acetic acid \hspace{1cm} Acetate ion

If we add base (OH-) does Le Chatelier say will happen?

- It is removed in the neutralization reaction!

A buffer is a solution made for fairly high concentration of a weak acid and it’s conjugate base.
Buffers are most resistant to pH changes when the pH = pKₐ of the weak acid and are effective when the pH is within one unit of the pKₐ (pH = pKₐ ± 1).

<table>
<thead>
<tr>
<th>Name</th>
<th>Formula</th>
<th>K_a</th>
<th>pKₐ</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric acid</td>
<td>HCl</td>
<td>1.0 × 10⁷</td>
<td>−7.00</td>
</tr>
<tr>
<td>Phosphoric acid</td>
<td>H₃PO₄</td>
<td>7.5 × 10⁻³</td>
<td>2.12</td>
</tr>
<tr>
<td>Hydrofluoric acid</td>
<td>HF</td>
<td>6.6 × 10⁻⁴</td>
<td>3.18</td>
</tr>
<tr>
<td>Lactic acid</td>
<td>CH₃CH(OH)CO₂H</td>
<td>1.4 × 10⁻⁴</td>
<td>3.85</td>
</tr>
<tr>
<td>Acetic acid</td>
<td>CH₃CO₂H</td>
<td>1.8 × 10⁻⁵</td>
<td>4.74</td>
</tr>
<tr>
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<td>H₂CO₃</td>
<td>4.4 × 10⁻⁷</td>
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<tr>
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<td>H₂PO₄⁻</td>
<td>6.2 × 10⁻⁸</td>
<td>7.21</td>
</tr>
<tr>
<td>Ammonium ion</td>
<td>NH₄⁺</td>
<td>5.6 × 10⁻¹⁰</td>
<td>9.25</td>
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<td>HPO₄²⁻</td>
<td>4.2 × 10⁻¹³</td>
<td>12.38</td>
</tr>
</tbody>
</table>
Buffers in the Blood

• The pH of your blood normally ranges between 7.35 and 7.45.

• A blood pH below the normal range is called **acidosis**, while a blood pH above this range is called **alkalosis**, either one of which is potentially fatal.
Buffers in the Blood

• Blood is kept in this narrow range (pH = 7.35 – 7.45) with the help of buffers.

• The most important buffer system in blood is formed from carbonic acid (H$_2$CO$_3$) and its conjugate base, hydrogen carbonate ion (HCO$_3^-$):

\[
\text{H}_2\text{CO}_3 + \text{H}_2\text{O} \rightleftharpoons \text{HCO}_3^- + \text{H}_3\text{O}^+
\]