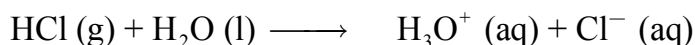


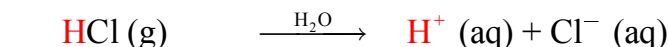
ACIDS & BASES

- Many common substances in our daily lives are **acids and bases**. **Oranges, lemons and vinegar** are examples of **acids**. In addition, our stomachs contain acids that help digest foods. **Antacid** tablets taken for heartburn and **ammonia** cleaning solutions are examples of **bases**.
- General properties associated with acids include the following:
 - sour taste
 - change color of litmus from blue to red
 - react with metals to produce H₂ gas
 - react with bases to produce salt and water
- General properties associated with bases include the following:
 - bitter taste
 - slippery soapy feeling
 - change color of litmus from red to blue
 - react with acids to produce salt and water
- The most common definition of acids and bases was formulated by the Swedish chemist Svante **Arrhenius** in 1884.
- According to the **Arrhenius** definition,

Acids are substances that produce hydronium ion (H₃O⁺) in aqueous solution

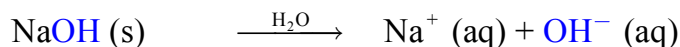


Commonly written as

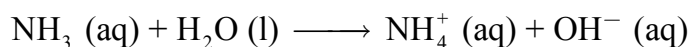


polar covalent

Bases are substances that produce hydroxide ion (OH⁻) in aqueous solution



ionic compound

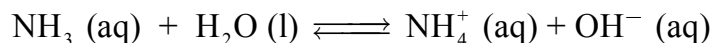


ACID & BASE STRENGTH

- According to the **Arrhenius** definition, the **strength** of acids and bases is based on the amount of their **ionization** in water.
- **Strong acids and bases are those that ionize completely in water.**
- Strong acids and bases are **strong electrolytes**.



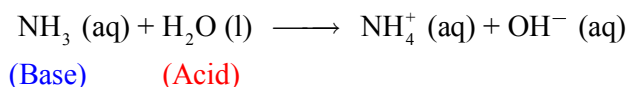
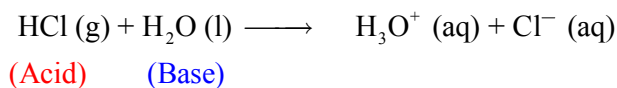
- **Weak acids and bases are those that ionize partially in water.**
- Weak acids and bases are **weak electrolytes**.



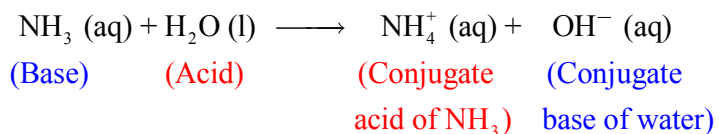
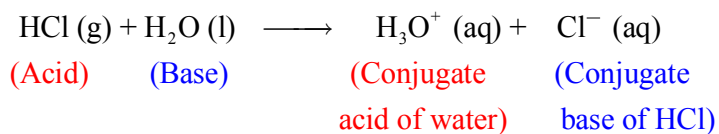
COMMON STRONG ACIDS & BASES			
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HBr	Hydrobromic acid	NaOH	Sodium hydroxide
HI	Hydroiodic acid	KOH	Potassium hydroxide
HNO ₃	Nitric acid	Ba(OH) ₂	Barium hydroxide
H ₂ SO ₄	Sulfuric acid		
COMMON WEAK ACIDS & BASES			
HC ₂ H ₃ O ₂	Acetic acid	NH ₃	Ammonia
H ₂ CO ₃	Carbonic acid	CO(NH ₂) ₂	Urea
HF	Hydrofluoric acid		
HCN	Hydrocyanic acid		
H ₂ S	Hydrosulfuric acid		

BRØNSTED-LOWRY ACIDS & BASES

- The **Arrhenius** definition of acids and bases is **limited to aqueous solutions**.
- A broader definition of acids and bases was developed by **Brønsted and Lowry** in the early 20th century.
- According to **Brønsted-Lowry** definition, an **acid** is a **proton donor**, and a **base** is a **proton acceptor**.

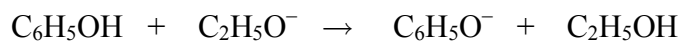
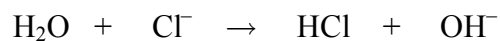


- A substance that can act as a **Brønsted-Lowry acid and base** (such as **water**) is called **amphiprotic**.
- In Brønsted-Lowry definition, any pair of molecules or ions that can be **interconverted by transfer of a proton** is called **conjugate acid-base pair**.



BRØNSTED-LOWRY ACIDS & BASES**Examples:**

1. Identify the conjugate acid-base pairs for each reaction shown below:



2. Write the formula for the conjugate acid for each base shown:

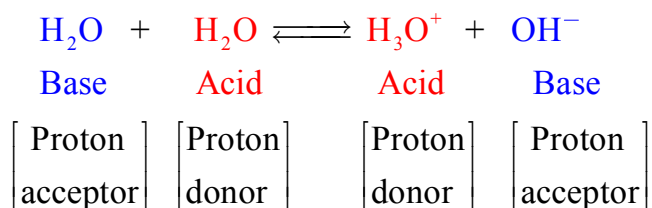


3. Write the formula for the conjugate base for each acid shown:



IONIZATION OF WATER

- As noted previously, **water** can act both as an **acid and a base**.
- In pure water, one **water** molecule **donates a proton** to another water molecule to produce **ions**.



- In pure water, the **transfer of protons** between water molecules produces **equal numbers of H₃O⁺ and OH⁻ ions**. However, the number of ions produced in pure water is very small, as indicated below:

$$\text{Pure water} \quad [\text{H}_3\text{O}^+] = [\text{OH}^-] = 1.0 \times 10^{-7} \text{ M}$$

- When the **concentrations of H₃O⁺ and OH⁻ are multiplied** together, the **ion-product constant (K_w)** is formed.

$$\begin{aligned}
 K_w &= [\text{H}_3\text{O}^+] \times [\text{OH}^-] \\
 &= (1.0 \times 10^{-7} \text{ M}) \times (1.0 \times 10^{-7} \text{ M}) = 1.0 \times 10^{-14}
 \end{aligned}$$

- All aqueous solutions have **H₃O⁺ and OH⁻ ions**. An **increase in the concentration** of one of the ions will cause an equilibrium shift that causes a **decrease in the other one**.

ACIDIC & BASIC SOLUTIONS

- When $[\text{H}_3\text{O}^+]$ and $[\text{OH}^-]$ are equal in a solution, it is **neutral**.
- When $[\text{H}_3\text{O}^+]$ is **greater** than $[\text{OH}^-]$ in a solution, it is **acidic**.

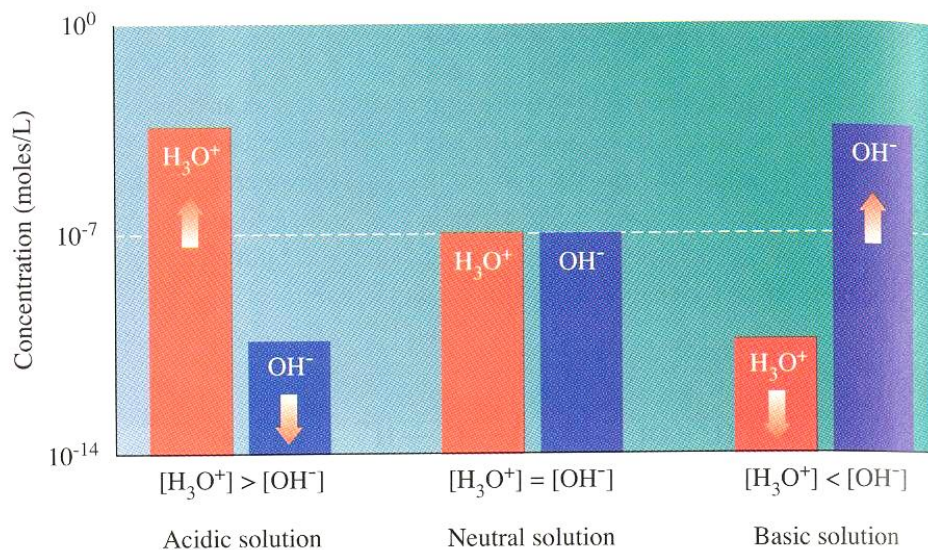
For example, if $[\text{H}_3\text{O}^+]$ is 1.0×10^{-4} M, then $[\text{OH}^-]$ would be 1.0×10^{-10} M.

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-4}} = 1.0 \times 10^{-10} \text{ M}$$

- When $[\text{OH}^-]$ is **greater** than $[\text{H}_3\text{O}^+]$ in a solution, it is **basic**.

For example, if $[\text{OH}^-]$ is 1.0×10^{-6} M, then $[\text{H}_3\text{O}^+]$ would be 1.0×10^{-8} M.

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} = \frac{1.0 \times 10^{-14}}{1.0 \times 10^{-6}} = 1.0 \times 10^{-8} \text{ M}$$



ACIDIC & BASIC SOLUTIONS**Examples:**

1. Calculate the $[\text{OH}^-]$ in a solution with $[\text{H}_3\text{O}^+] = 2.3 \times 10^{-4} \text{ M}$. Classify the solution as acid or basic.

$$[\text{OH}^-] = \frac{K_w}{[\text{H}_3\text{O}^+]} =$$

2. Calculate the $[\text{H}_3\text{O}^+]$ in a solution with $[\text{OH}^-] = 2.3 \times 10^{-4} \text{ M}$. Classify the solution as acid or basic.

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} =$$

3. Calculate the $[\text{OH}^-]$ in a solution with $[\text{H}_3\text{O}^+] = 5.8 \times 10^{-8} \text{ M}$. Classify the solution as acid or basic.

4. Calculate the $[\text{H}_3\text{O}^+]$ in a solution with $[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$. Classify the solution as acid or basic.

THE pH SCALE

- The **acidity** of a solution is commonly measured on a **pH scale**.

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

- The **pH scale** ranges from **0-14**, where **acidic** solutions are **less than 7** and **basic** solutions are **greater than 7**.

Acidic solutions	pH < 7	$[\text{H}_3\text{O}^+] > 1.0 \times 10^{-7}$
Neutral solutions	pH = 7	$[\text{H}_3\text{O}^+] = 1.0 \times 10^{-7}$
Basic solutions	pH > 7	$[\text{H}_3\text{O}^+] < 1.0 \times 10^{-7}$

Examples:

- The $[\text{H}_3\text{O}^+]$ of a liquid detergent is 1.4×10^{-9} M. Calculate its pH.

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = -\log [1.4 \times 10^{-9}] = -(-8.85) = 8.85$$

- The pH of black coffee is 5.3. Calculate its $[\text{H}_3\text{O}^+]$.

$$[\text{H}_3\text{O}^+] = \text{antilog} (-\text{pH}) = 10^{-\text{pH}} = 10^{-5.3} = 5 \times 10^{-6}$$

- The $[\text{H}_3\text{O}^+]$ of a solution is 3.5×10^{-3} M. Calculate its pH.

- The pH of tomato juice is 4.1. Calculate its $[\text{H}_3\text{O}^+]$.

- The $[\text{OH}^-]$ of a cleaning solution is 1.0×10^{-5} M. What is the pH of this solution?

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

- The pH of a solution is 11.50. Calculate the $[\text{H}_3\text{O}^+]$ for this solution.

THE pOH SCALE

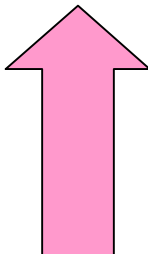

- Similar to pH, a convenient way to designate the **basicity** of a solution is the **pOH scale**.

$$\text{pOH} = -\log [\text{OH}^-]$$

- In aqueous solutions, the **sum of pH and pOH** of a solution equal to **14**.

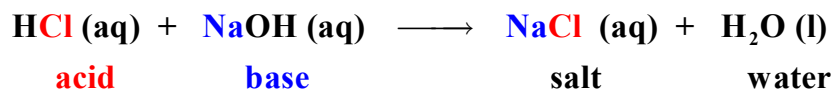
$$\text{pH} + \text{pOH} = 14$$

Comparison of $[\text{H}_3\text{O}^+]$, $[\text{OH}^-]$, pH and pOH Values

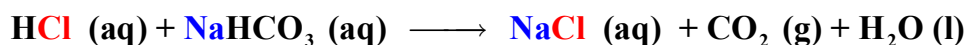
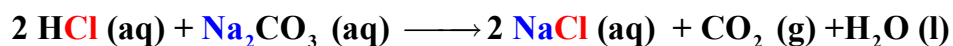
<p style="color: red; font-weight: bold;">ACIDIC</p> 	<table style="width: 100%; border-collapse: collapse;"> <thead> <tr> <th style="border: none;">$[\text{H}_3\text{O}^+]$</th> <th style="border: none;">pH</th> <th style="border: none;">pOH</th> <th style="border: none;">$[\text{OH}^-]$</th> </tr> </thead> <tbody> <tr><td style="border: none;">10^0</td><td style="border: none;">0</td><td style="border: none;">14</td><td style="border: none;">10^{-14}</td></tr> <tr><td style="border: none;">10^{-1}</td><td style="border: none;">1</td><td style="border: none;">13</td><td style="border: none;">10^{-13}</td></tr> <tr><td style="border: none;">10^{-2}</td><td style="border: none;">2</td><td style="border: none;">12</td><td style="border: none;">10^{-12}</td></tr> <tr><td style="border: none;">10^{-3}</td><td style="border: none;">3</td><td style="border: none;">11</td><td style="border: none;">10^{-11}</td></tr> <tr><td style="border: none;">10^{-4}</td><td style="border: none;">4</td><td style="border: none;">10</td><td style="border: none;">10^{-10}</td></tr> <tr><td style="border: none;">10^{-5}</td><td style="border: none;">5</td><td style="border: none;">9</td><td style="border: none;">10^{-9}</td></tr> <tr><td style="border: none;">10^{-6}</td><td style="border: none;">6</td><td style="border: none;">8</td><td style="border: none;">10^{-8}</td></tr> <tr><td style="border: none;">10^{-7}</td><td style="border: none;">7</td><td style="border: none;">7</td><td style="border: none;">10^{-7}</td></tr> <tr><td style="border: none;">10^{-8}</td><td style="border: none;">8</td><td style="border: none;">6</td><td style="border: none;">10^{-6}</td></tr> <tr><td style="border: none;">10^{-9}</td><td style="border: none;">9</td><td style="border: none;">5</td><td style="border: none;">10^{-5}</td></tr> <tr><td style="border: none;">10^{-10}</td><td style="border: none;">10</td><td style="border: none;">4</td><td style="border: none;">10^{-4}</td></tr> <tr><td style="border: none;">10^{-11}</td><td style="border: none;">11</td><td style="border: none;">3</td><td style="border: none;">10^{-3}</td></tr> <tr><td style="border: none;">10^{-12}</td><td style="border: none;">12</td><td style="border: none;">2</td><td style="border: none;">10^{-2}</td></tr> <tr><td style="border: none;">10^{-13}</td><td style="border: none;">13</td><td style="border: none;">1</td><td style="border: none;">10^{-1}</td></tr> <tr><td style="border: none;">10^{-14}</td><td style="border: none;">14</td><td style="border: none;">0</td><td style="border: none;">10^0</td></tr> </tbody> </table>	$[\text{H}_3\text{O}^+]$	pH	pOH	$[\text{OH}^-]$	10^0	0	14	10^{-14}	10^{-1}	1	13	10^{-13}	10^{-2}	2	12	10^{-12}	10^{-3}	3	11	10^{-11}	10^{-4}	4	10	10^{-10}	10^{-5}	5	9	10^{-9}	10^{-6}	6	8	10^{-8}	10^{-7}	7	7	10^{-7}	10^{-8}	8	6	10^{-6}	10^{-9}	9	5	10^{-5}	10^{-10}	10	4	10^{-4}	10^{-11}	11	3	10^{-3}	10^{-12}	12	2	10^{-2}	10^{-13}	13	1	10^{-1}	10^{-14}	14	0	10^0	<p style="color: blue; font-weight: bold;">BASIC</p> 
$[\text{H}_3\text{O}^+]$	pH	pOH	$[\text{OH}^-]$																																																															
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REACTIONS OF ACIDS & BASES

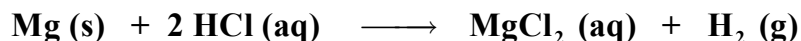
- The most important reaction of acids and bases is called **neutralization**. In these reactions an acid combines with a base to form a **salt and water**. For example:



- Acids also react with **carbonates and bicarbonates** to produce **salt, carbon dioxide gas and water**. For example:

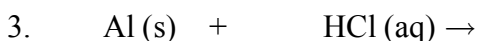
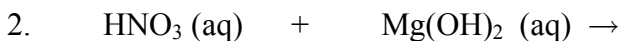


- Acids also react with **active metals** to produce a **salt and hydrogen gas**. For example:



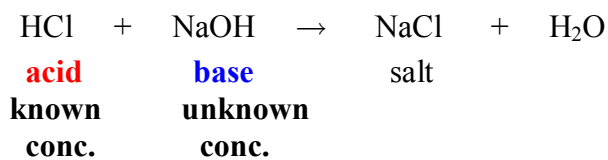
Examples:

Complete and balance each of the following equations:



TITRATION

- Determination of concentration of an acid or base from the known concentration of the other is called **titration**.



- When solving titration problems,
 - Write a balanced equation.
 - Determine moles of known from concentration and volume.
 - Determine moles of unknown based on reaction stoichiometry.
 - Determine the concentration of unknown from moles and volume titrated.

Examples:

- If 32.6 mL of 1.85 M NaOH is required to titrate 25.0 mL of an HCl solution to the end point, what is the molarity of HCl?

Bal. Eq.	HCl + NaOH → NaCl + H ₂ O
Vol (mL)	
Conc. (M)	
mol NaOH	
mol HCl	
Conc. HCl	

Examples:

2. How many mL of 2.15 M KOH are required to titrate 25.0 mL of 0.300 M HC₂H₃O₂?

Bal. Eq.	
Vol (mL)	
Conc. (M)	
mol HAc	
mol NaOH	
Vol NaOH (mL)	

3. If 50.0 mL of 0.300 M NaOH solution is required to titrate 20.0 mL of sulfuric acid, what is the molarity of the acid?

Bal. Eq.	
Vol (mL)	
Conc. (M)	
mol NaOH	
mol H ₂ SO ₄	
Conc. H ₂ SO ₄	