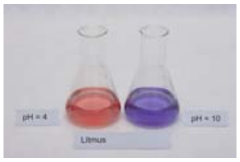
The Nature of Acids and Bases

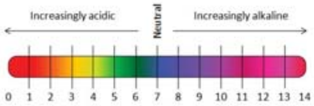
Section 8.1

Homework
Pg. 492 #1 d-e, 2
Pg. 493 #1
Pg. 494 #1-9

Observable Properties

	Acid	Base
pH value	< 7	> 7
taste	sour	bitter
feel	no characteristic feel	slippery
electrolytic	yes	yes
colour of litmus	red	blue





Earliest Definitions: Arrhenius Acids and Bases

An **acid** ionizes in water to produce **hydronium** ions (H_3O^+)	A **base** dissociates in water to produce **hydroxide** ions (OH^-)
$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$	$Ba(OH)_{2\,(s)} \rightarrow Ba^{2+}_{(aq)} + 2\,OH^-_{(aq)}$

Simpler definitions state that acids make H^+, but H^+ is so reactive it combines with H_2O:

$H_2O_{(l)} + H^+_{(aq)} \rightarrow H_3O^+_{(aq)}$

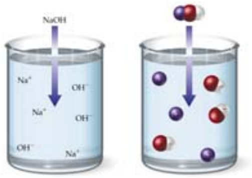
Updated Definitions: Brønsted-Lowry

An **acid** is a proton donor	A **base** is a proton acceptor
$HCl_{(aq)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$	$NH_{3\,(aq)} + H_2O_{(l)} \rightarrow NH_4^+{}_{(aq)} + OH^-_{(aq)}$
very similar to Arrhenius' definition	more explanatory power than Arrhenius' definition
Any substance that donates a proton to water increases the $[H_3O^+]$	Expands Arrhenius' definition: Substances that accept protons from water increase the $[OH^-]$

Formation of an aqueous acidic or basic solution involves **reaction** with H_2O.

Example 1: Sodium Hydroxide, NaOH

- NaOH is a base by **Arrhenius'** definition
  - ➢Contains an OH^- ion



- NaOH is also a base by **B/L** definition, since:

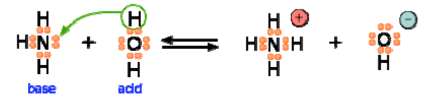
(Dissociation equation) $NaOH_{(s)} \rightleftharpoons Na^+_{(aq)} + OH^-_{(aq)}$,

and OH^- can then act as a proton acceptor:

(Example reaction) $OH^-_{(aq)} + H_3O^+_{(aq)} \rightleftharpoons 2\,H_2O_{(l)}$

Example 2: Ammonia, NH_3

- NH_3 is **NOT** an Arrhenius base
  - does NOT possess an OH^- ion



$NH_{3\,(g)} + H_2O_{(l)} \rightleftharpoons NH_4^+{}_{(aq)} + OH^-_{(aq)}$

- but it **IS** a Brønsted-Lowry base:
  - the lone pair on N can readily accept a H^+
  - in an aqueous solution, this reaction produces $OH^-_{(aq)}$
    - responsible for the basic properties of the solution

Strengths of the Brønsted-Lowry theory

- More general, therefore more explanatory power
  - Explains the basic properties of solutions where solutes do NOT possess an OH^- ion (e.g., $NH_{3\,(aq)}$)
  - Can be applied to reactions that do not occur in aqueous solutions.

According to Brønsted-Lowry theory, all acid-base reactions are **reversible H^+ exchange** (proton transfer) reactions

Practice. Label the acid and base in the following reactions

a) $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$

b) $PO_4^{3-} + H_2O \rightleftharpoons HPO_4^{2-} + OH^-$

c) $HCO_3^- + H_3O^+ \rightleftharpoons H_2CO_3 + OH^-$

Classification as a Bronsted-Lowry acid or base is not a permanent one – it depends on the particular reaction.

- An **amphiprotic substance** is one that acts as an acid in some reactions, and as a base in another.

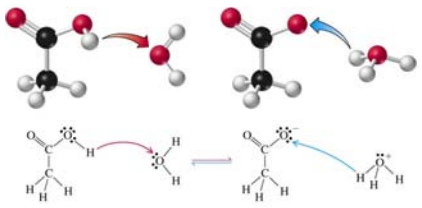
$$\underset{\text{base}}{CO_3^{2-}}{}_{(aq)} + \underset{\text{acid}}{H_2O}{}_{(l)} \rightleftharpoons OH^-{}_{(aq)} + HCO_3^-{}_{(aq)}$$

$$\underset{\text{acid}}{NH_4^+}{}_{(aq)} + \underset{\text{base}}{H_2O}{}_{(l)} \rightleftharpoons H_3O^+{}_{(aq)} + NH_{3\,(aq)}$$

Learning Checkpoint

- According to Bronsted-Lowry definitions,
  - An acid is a proton donor
  - A base is a proton acceptor
  - All acid-base reactions involve a proton exchange
  - Formation of an aqueous acidic or basic solution involves a REACTION with water (vs. simple dissociation or ionization)

Consider the ionization of acetic acid, $HC_2H_3O_2$

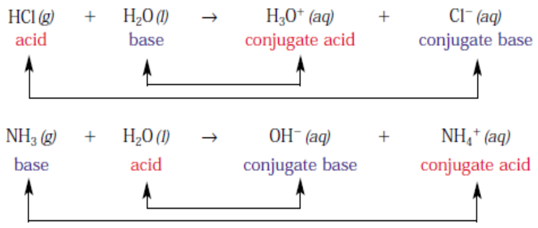


Forward reaction $\underset{\text{acid}}{HC_2H_3O_2} + \underset{\text{base}}{H_2O} \rightleftharpoons \underset{\text{conjugate base}}{C_2H_3O_2^-} + \underset{\text{conjugate acid}}{H_3O^+}$

Reverse reaction $\underset{\text{base}}{C_2H_3O_2^-} + \underset{\text{acid}}{H_3O^+} \rightleftharpoons HC_2H_3O_2 + H_2O$

Conjugate acid-base pairs

- Every acidic reactant has a corresponding basic product, and vice versa.
- These corresponding acid-base pairs are called **conjugate acids** and **conjugate bases**.

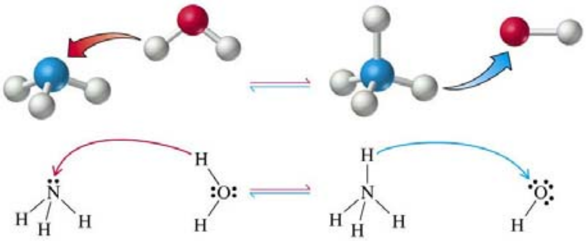


$$\underset{\text{acid}}{HCl\,(g)} + \underset{\text{base}}{H_2O\,(l)} \rightarrow \underset{\text{conjugate acid}}{H_3O^+\,(aq)} + \underset{\text{conjugate base}}{Cl^-\,(aq)}$$

$$\underset{\text{base}}{NH_3\,(g)} + \underset{\text{acid}}{H_2O\,(l)} \rightarrow \underset{\text{conjugate base}}{OH^-\,(aq)} + \underset{\text{conjugate acid}}{NH_4^+\,(aq)}$$

$$NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$$

NH_3	H_2O	NH_4^+	OH^-
base	acid	*conjugate acid*	*conjugate base*



Conjugate acid-base pairs differ in formula by one proton (H^+)

- one hydrogen
- charge of one

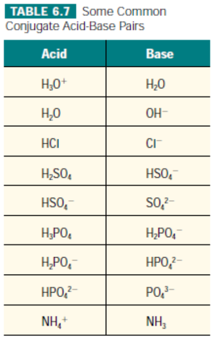


TABLE 6.7 Some Common Conjugate Acid-Base Pairs

Acid	Base
H_3O^+	H_2O
H_2O	OH^-
HCl	Cl^-
H_2SO_4	HSO_4^-
HSO_4^-	SO_4^{2-}
H_3PO_4	$H_2PO_4^-$
$H_2PO_4^-$	HPO_4^{2-}
HPO_4^{2-}	PO_4^{3-}
NH_4^+	NH_3

Example 1. The carbonate ion, **CO_3^{2-}**, forms a **basic** solution in water.

a) Write out the balanced equation for the reaction of the carbonate ion with water.

b) Identify the conjugate acid-base pairs.

$$CO_3^{2-}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons OH^-{}_{(aq)} + \_\_\_\_\_\_\_\_\_\_\_\_$$

Example 2. Do the same for ammonium, **NH_4^+**, which forms an **acidic** solution in water.

$$NH_4^+{}_{(aq)} + H_2O_{(l)} \rightleftharpoons \_\_\_\_\_\_\_\_\_\_ + \_\_\_\_\_\_\_\_\_\_\_\_$$

Learning Checkpoint

- When an acid donates a proton, it forms a conjugate base.
- When a base accepts a proton, it forms a conjugate acid.

- Conjugate acid-base pairs differ in their formulas by one H^+.

All acid-base reactions can involve a COMPETITION for protons.

The ionization of acetic acid:

$$\boxed{HC_2H_3O_2{}_{(aq)} + H_2O_{(l)} \overset{1.3\%}{\rightleftharpoons} C_2H_3O_2^-{}_{(aq)} + H_3O^+{}_{(aq)}}$$

$$C_2H_3O_2^- \dashleftarrow H^+ \dashrightarrow H_2O$$

acetate ion — water

$C_2H_3O_2^-$ has a stronger hold on the proton.
Because of this, it ionizes very little (1.3%).

The ionization of hydrochloric acid:

$$\boxed{HCl_{(aq)} + H_2O_{(l)} \xrightarrow{99\%} Cl^-_{(aq)} + H_3O^+_{(aq)}}$$

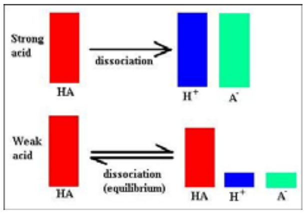
$\boxed{}$ ⟵ H^+ ⟶ H_2O

water

Who has the stronger pull on the proton?

Based on how well it ionizes in water, an acid can be classified as strong or weak.

- **strong acids:** ionize completely (100%)
  - HCl, HI, HBr, $HClO_4$, H_2SO_4, HNO_3
- **weak acids**: partial ionization; some still exists in molecular form
  - e.g., $HC_2H_3O_2$ (acetic acid)



You must MEMORIZE the strong acids!

Weak acids form a dynamic equilibrium with their conjugate bases.

$$\underbrace{HA_{(aq)}}_{\text{acid}} + \underbrace{H_2O_{(l)}}_{\text{base}} \rightleftharpoons \underbrace{H_3O^+_{(aq)}}_{\text{conjugate acid}} + \underbrace{A^-_{(aq)}}_{\text{conjugate base}} \qquad K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

acid ionization constant

Simplify by representing as an ionization (omit H_2O)

$$\underbrace{HA_{(aq)}}_{\text{acid}} \rightleftharpoons H^+_{(aq)} + \underbrace{A^-_{(aq)}}_{\text{conjugate base}} \qquad K_a = \frac{[H^+][A^-]}{[HA]}$$

Example. Writing a K_a expression

Write a K_a expression for acetic acid, $HC_2H_3O_{2\,(aq)}$

Strategy

1. Write the equilibrium reaction equation. Simplify the scenario by treating it as a simple ionization (instead of a reaction with water)
2. Write the K_a expression – don't forget the subscript!

❶ $HC_2H_3O_{2\,(aq)} \rightleftharpoons H^+_{(aq)} + C_2H_3O_2^-{}_{(aq)}$

❷ $K_a = \dfrac{[H^+]\,[C_2H_3O_2^-]}{[HC_2H_3O_2]}$

Practice

Write a K_a expression for chloracetic acid, $HC_2H_2O_2Cl_{(aq)}$

Learning Checkpoint

- A weak acid is one that does not react (ionize) completely.
  - forms an equilibrium with its conjugate base
- The equilibrium law constant for an acid ionization is called the acid dissociation constant, K_a.
- For the general acid ionization reaction

$HA \rightleftharpoons H^+ + A^-$, $\quad K_a = \dfrac{[H^+][A^-]}{[HA]}$