

Acid-Base Equilibria

Chapter 16

Chemistry, The Central Science
11th edition



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Acids and Bases

- Acids and bases are found in many common substances and are important in life processes.
- A range of chemical processes from industrial to biological ones, from reactions in the lab to those occurring in the environment, involve either an acid or a base.
- As such, a massive amount of chemistry can be understood in terms of acid-base reactions.
- There are several models for what constitutes an acid or a base - three models to be discussed in this chapter.

Acids and Bases

Acid	Base
Tastes sour	Tastes bitter
Stings skin	Feels slippery
Turns litmus red	Turns litmus blue
Turns phenolphthalein colourless	Turns phenolphthalein pink
Corrosive to metals	Fouling on metal surfaces
<u>Examples:</u> <ul style="list-style-type: none">- Ascorbic acid (vitamin C)- Citric acid (lemon and orange)	<u>Examples:</u> <ul style="list-style-type: none">- CaCO_3 is used in antacid tables- NaOH to unblock drains

Acids and Bases

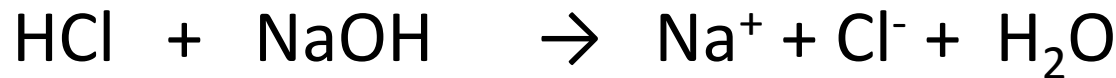
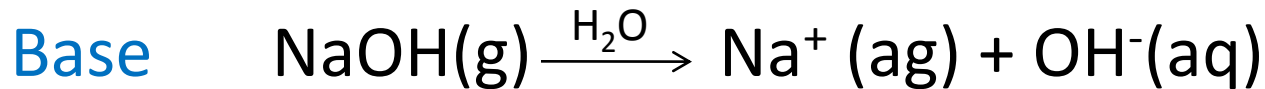
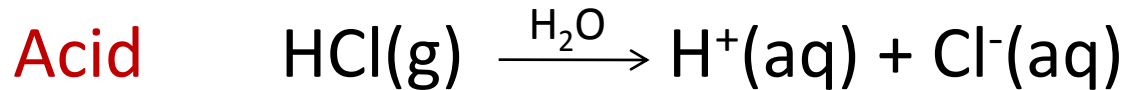
- React together to form a **salt** with loss of the characteristic acid/base properties.

Arrhenius Theory

- **Acid:** substance that produces H^+ in aqueous solution
- **Base:** substance that produces OH^- in aqueous solution

Acids and Bases

Examples of Arrhenius acids and bases:



- The Arrhenius concept is limited to substances that dissolve in water.

Brønsted-Lowry Acids and Bases

In 1923 two chemists, Brønsted and Lowry suggested a more general definition of acids and bases.

- Their concept was based on the fact that acid-base reactions to involve the transfer of a proton, (a proton is an H^+ ion), from an acid to a base.
- This transfer may or may not involve water.
- An acid is a species that _____ a proton.
- A base is a species that _____ a proton.

Brønsted-Lowry Acids and Bases

Examples

- Acids:



- Bases:



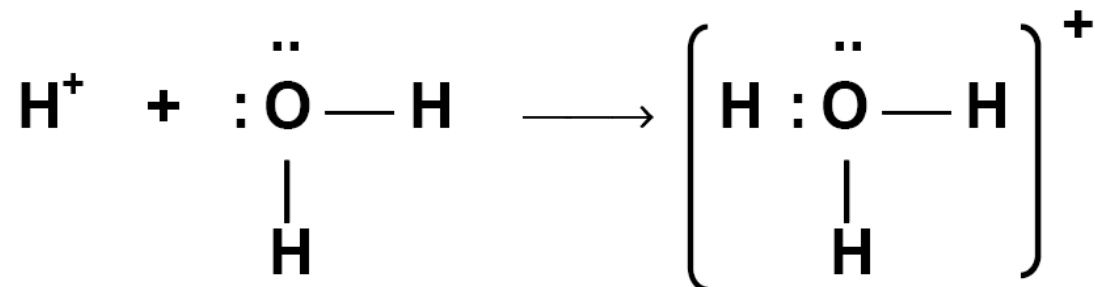
Brønsted-Lowry Acids and Bases

The H⁺ Ion in Water

- If acid-base reactions are proton (H⁺ ion) transfer reactions, we need to know what the H⁺ ion is.
- The hydrogen ion, H⁺ ion, is a very small “bare” proton with no surrounding valence electron.
- This small, positively charged particle interacts with the nonbonding electron pairs (lone pair) of H₂O molecules to form a hydronium ion, H₃O⁺.

Brønsted-Lowry Acids and Bases

The H⁺ Ion in Water



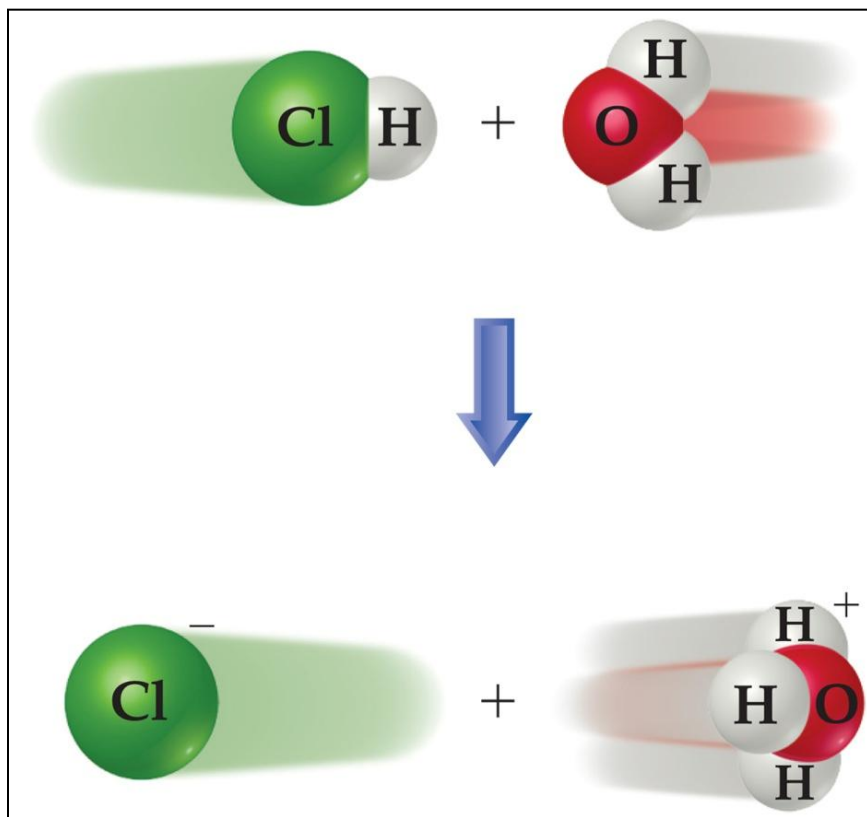
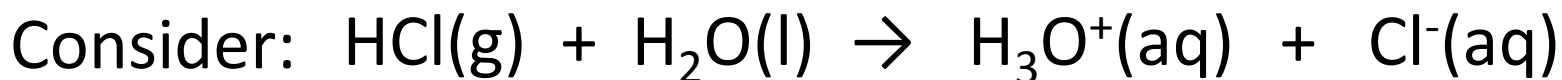
The hydronium ion

- H⁺ and H₃O⁺ represent the same thing – hydrated proton.
- They can be used interchangeably.

Brønsted-Lowry Acids and Bases

Proton-Transfer Reactions

Let us see what happens when an acid dissolves in water...



The polar H₂O molecule promotes the ionization of acids in an aqueous (water) solution by accepting a proton to form H₃O⁺.

Brønsted-Lowry Acids and Bases

Proton-Transfer Reactions

Brønsted and Lowry proposed that acids and bases should be defined as follows:

ACID : An acid is a substance (molecule or ion) that can transfer a proton to another substance.

BASE : A base is a substance that can accept a proton from another substance.

These are termed Brønsted-Lowry acids and bases.

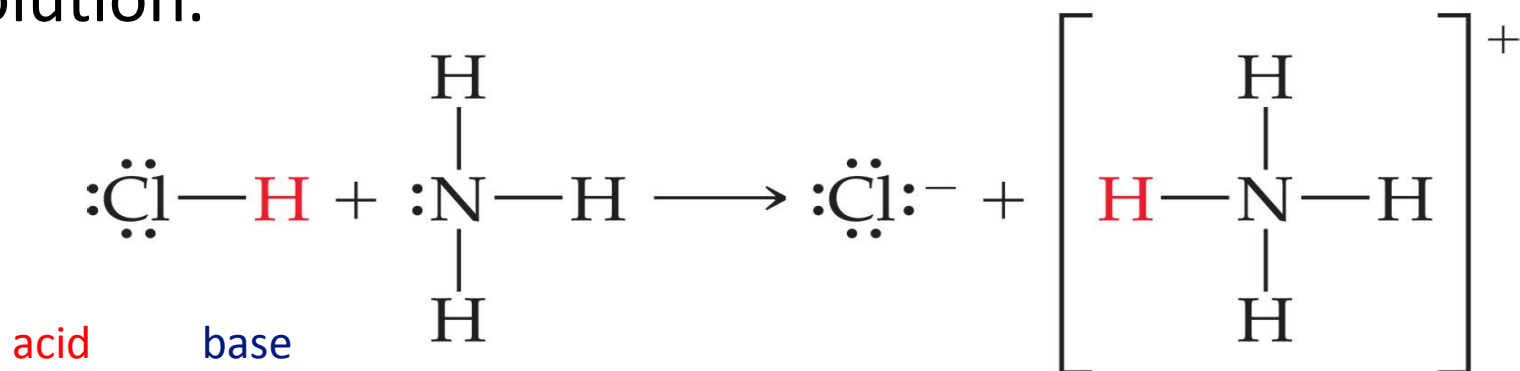
Brønsted-Lowry Acids and Bases

Proton-Transfer Reactions

- Simply put, according to the Brønsted-Lowry definitions:

An acid is a proton donor
A base is a proton acceptor

- The Brønsted-Lowry concept is on proton transfer and also applies to reactions that do not occur in aqueous solution.



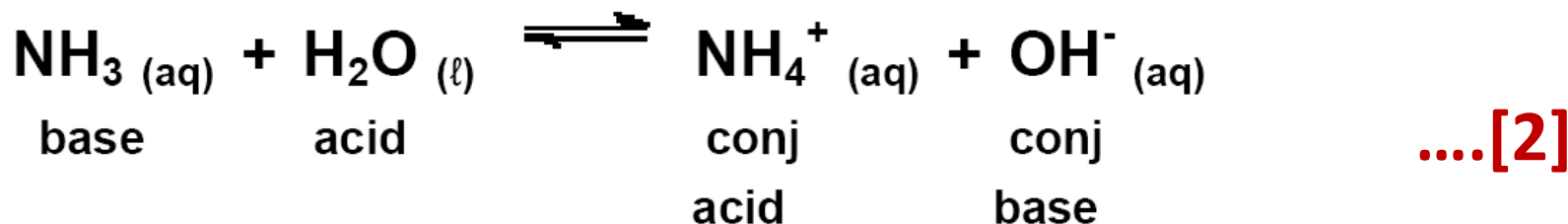
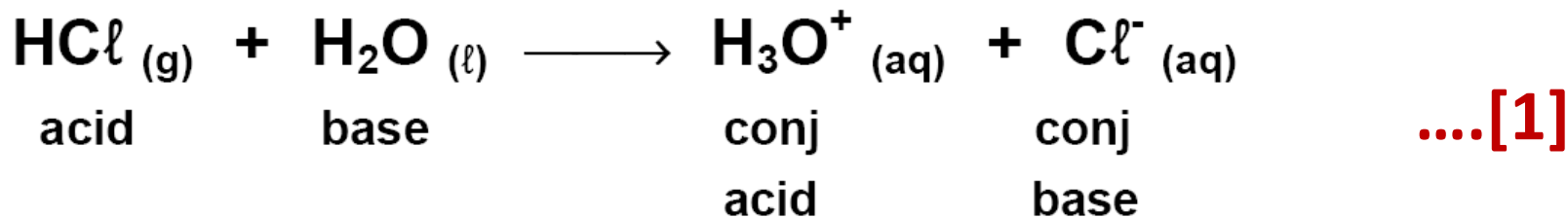
- This reaction occurs in the gas phase

Brønsted-Lowry Acids and Bases

Proton-Transfer Reactions

Amphiprotic substances

- Consider the reaction of H₂O with HCl in [1] and NH₃ in [2].



- In [1] H₂O is a Brønsted-Lowry base it _____ a proton.
- In [2] H₂O is a Brønsted-Lowry acid it _____ a proton.

Brønsted-Lowry Acids and Bases

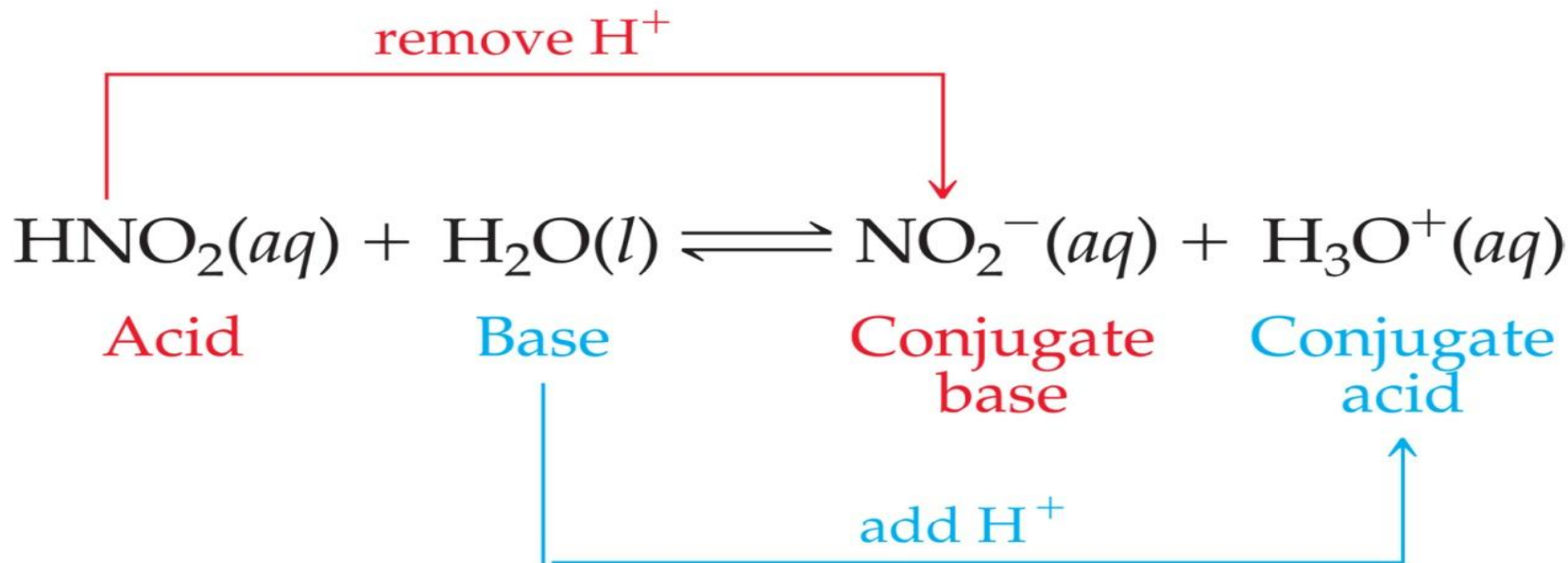
Proton-Transfer Reactions

A substance that is capable of acting as either an acid or a base is called **amphiprotic**.

Brønsted-Lowry Acids and Bases

Conjugate Acid-Base Pairs

- From the Latin word *conjugare*, meaning “to join together.”
- Reactions between acids and bases always yield their conjugate bases and acids.



Brønsted-Lowry Acids and Bases

Conjugate Acid-Base Pairs

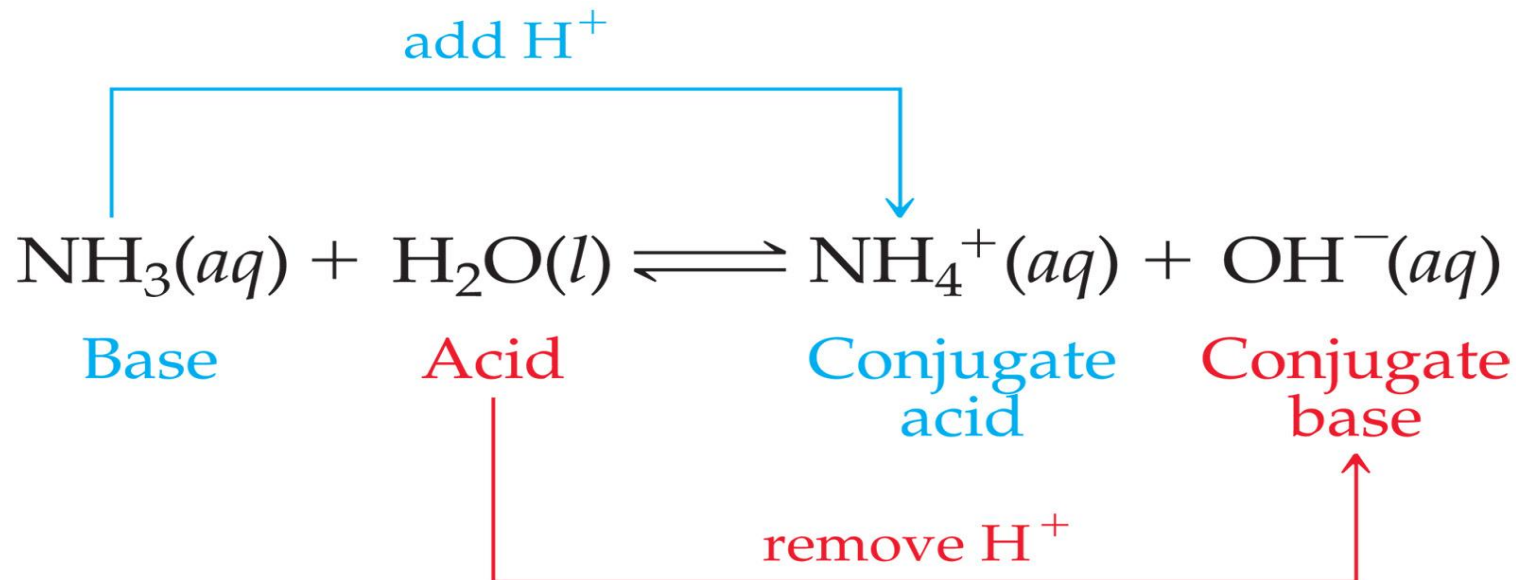
- Consider the reaction of an acid (HA) with water:



- After HA (acid) loses its proton it is converted into A⁻ (base). Therefore HA and A⁻ are **conjugate acid-base pairs**.
- After H₂O (base) gains a proton it is converted into H₃O⁺ (acid). Therefore, H₂O and H₃O⁺ are **conjugate acid-base pairs**.
- Conjugate acid-base pairs differ by only one proton.

Brønsted-Lowry Acids and Bases

Conjugate Acid-Base Pairs

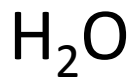


Brønsted-Lowry Acids and Bases

Conjugate Acid-Base Pairs

Examples

conjugate acid



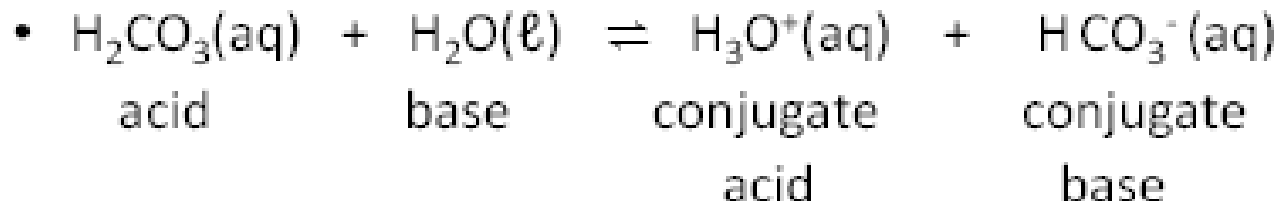
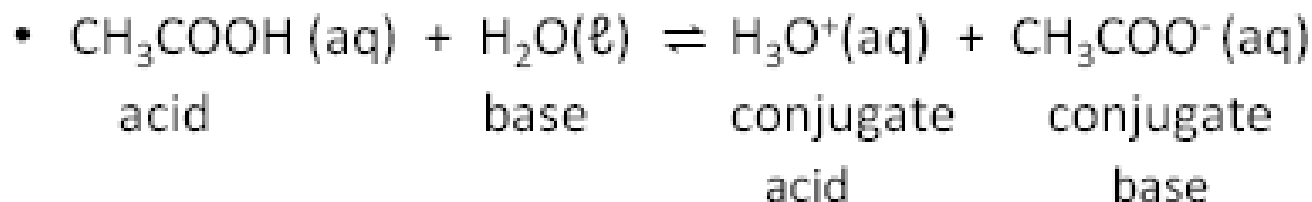
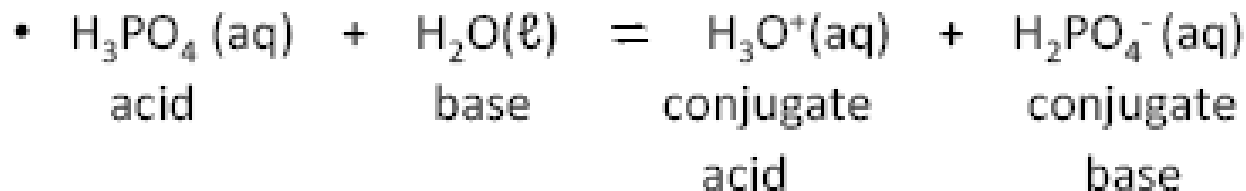
conjugate base



Brønsted-Lowry Acids and Bases

Conjugate Acid-Base Pairs

More examples



Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- **Strong acids** are completely dissociated in water.
 - Their conjugate bases are quite weak.
 - They are good proton donors.
 - examples: HCl, HI, HBr, HNO₃, H₂SO₄, HClO₄
- Example: $\text{HCl(aq)} + \text{H}_2\text{O(l)} \rightarrow \text{H}_3\text{O}^+(\text{aq}) + \text{Cl}^-(\text{aq})$
 1.2 M 1.2 M 1.2 M
- Acid fully dissociates. No equilibrium occurs between the acid and its ions.
- A large acid dissociation constant, K_a , is expected.
- K_a is the equilibrium constant for the dissociation of an acid.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- **Weak acids** only dissociate partially in water.
 - Their conjugate bases are weak bases.
 - When weak acids are in solution, there is an equilibrium between the acid molecules and the constitute ions.
 - Examples: CH_3COOH , HNO_2 , H_2CO_3 , H_3BO_3 , HF .
- **Example:**
$$\text{CH}_3\text{COOH}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{CH}_3\text{COO}^-(\text{aq})$$
- The K_a is very small (1.8×10^{-5}).
- Some of the CH_3COOH remains undissociated.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- **Strong bases** are completely dissociated in water.
 - They have a great attraction to protons.
 - They have a pair of electrons to form a bond with the proton.
 - examples: NaOH, KOH, Ca(OH)₂, O²⁻, H⁻
- **Example:**
$$\text{O}^{2-}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightarrow \text{OH}^{-}(\text{aq}) + \text{OH}^{-}(\text{aq})$$
- Expect a large base dissociation constant, K_b .
- K_b is the equilibrium constant for the dissociation of a base.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- **Weak bases** are molecules or ions that are partially ionized in water to give OH⁻
 - They are poor proton acceptors
 - examples: NH₃, CH₃CO₂⁻
- **Example**
NH₃(aq) + H₂O(l) ⇌ OH⁻(aq) + NH₄⁺(aq)
- The equilibrium constant, K_b, is small, 1.8 x 10⁻⁵.
Therefore the base is weak.
- Some NH₃ molecules remain undissociated.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- Substances with negligible acidity such as CH_4 do not dissociate in water.
 - Their conjugate bases are exceedingly strong and react completely with water.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

Strong and weak vs concentrated and dilute

Do not confuse the meaning of the terms 'strong' and 'weak' with the terms 'concentrated' and 'dilute'.

Examples:

1.0 M HCl = a concentrated solution of a strong acid.

0.0001 M HCl = a dilute solution of a strong acid.

1.0 M HOAc = a concentrated solution of a weak acid.

0.001 M HOAc = a dilute solution of a weak acid.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- A strong acid has a weak conjugate base.
- The stronger the acid the weaker is its conjugate base.
- Very strong acids have conjugate bases with negligible base strength.
- The more easily an acid gives up a proton, the less easily its conjugate base accepts a proton.
- A strong base has a weak conjugate acid.
- The stronger the base the weaker is its conjugate acid.
- Very strong bases have conjugate acids with negligible acid strength.
- The more readily a base accepts a proton, the less easily its conjugate acid gives up a proton.

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

- A weak acid has a weak conjugate base.
- A weak acid only partially dissociates in aqueous solution.
- The conjugate base of a weak acid shows a slight ability to remove protons from water.

Relative strengths of some conjugate-acid base pairs.

	ACID	BASE		
100% ionized in H ₂ O	Strong	HCl	Cl ⁻	Negligible
		H ₂ SO ₄	HSO ₄ ⁻	
		HNO ₃	NO ₃ ⁻	
	H ₃ O ⁺ (aq)	H ₂ O		
Acid strength increases ↑	Weak	HSO ₄ ⁻	SO ₄ ²⁻	Weak
		H ₃ PO ₄	H ₂ PO ₄ ⁻	
		HF	F ⁻	
		HC ₂ H ₃ O ₂	C ₂ H ₃ O ₂ ⁻	
		H ₂ CO ₃	HCO ₃ ⁻	
		H ₂ S	HS ⁻	
		H ₂ PO ₄ ⁻	HPO ₄ ²⁻	
		NH ₄ ⁺	NH ₃	
		HCO ₃ ⁻	CO ₃ ²⁻	
		HPO ₄ ²⁻	PO ₄ ³⁻	
	H ₂ O	OH ⁻		
Negligible	Strong	OH ⁻	O ²⁻	100% protonated in H ₂ O
		H ₂	H ⁻	
		CH ₄	CH ₃ ⁻	

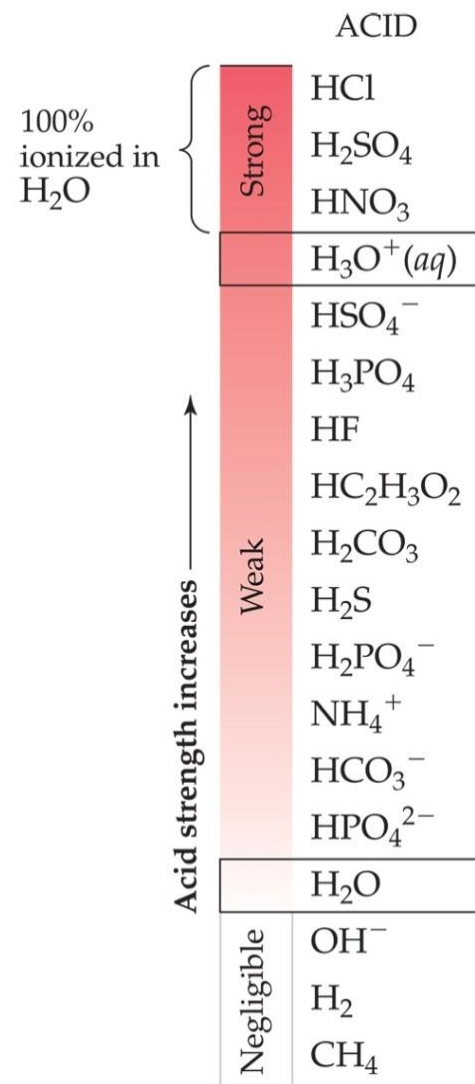
↓ Base strength increases

Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

Which base is the weakest?

- F^-
- NH_3
- OH^-
- SO_4^{2-}
- $C_2H_3O_2^-$



Brønsted-Lowry Acids and Bases

Relative Strengths of Acids and Bases

Which acid is the strongest?

- H_2O
- H_3O^+
- HF
- $\text{HC}_2\text{H}_3\text{O}_2$
- NH_4^+

BASE	
Cl^-	Negligible
HSO_4^-	
NO_3^-	
H_2O	
SO_4^{2-}	Weak
H_2PO_4^-	
F^-	
$\text{C}_2\text{H}_3\text{O}_2^-$	
HCO_3^-	
HS^-	
HPO_4^{2-}	
NH_3	
CO_3^{2-}	
PO_4^{3-}	
OH^-	
O^{2-}	Strong
H^-	
CH_3^-	

Base strength increases ↓

100% protonated in H_2O

The Autoionization of Water

The Ion Product Constant

- The autoionization of water is an equilibrium process: $\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$
- The equilibrium-constant expression for this process is:

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

- This special equilibrium constant is referred to as the **ion-product constant** for water, K_w .
- At 25 C,

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

The Autoionization of Water

The Ion Product Constant

In NEUTRAL solutions $[H^+] \equiv [OH^-]$

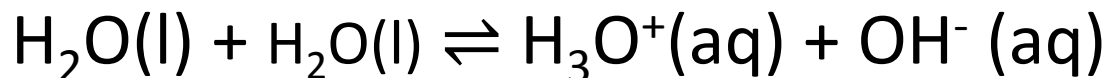
In ACIDIC solutions $[H^+] > [OH^-]$

In BASIC solutions $[OH^-] > [H^+]$

Acids and Bases

Summary so far

- Acids – proton donors
- Bases – proton acceptors
- Weak acids and weak bases – partly dissociate
- Strong acids and strong bases – fully dissociate
- Conjugate acid-base pair – one H⁺ difference
- Autoionisation of water



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

The pH Scale

- The concentration of H_3O^+ or OH^- in solution can vary over a wide range, but these concentrations are usually small.
- We define

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

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

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

The pH Scale

	Strongly acidic	Neutral	Strongly alkaline
$[\text{H}_3\text{O}^+]$	10^{-1} M	10^{-7} M	10^{-14} M
pH	1	7	14

- In neutral water at 25 °C, $\text{pH} = \text{pOH} = 7.00$.
- In acidic solutions, $[\text{H}^+] > 1.0 \times 10^{-7}$, so $\text{pH} < 7.00$.
- In basic solutions, $[\text{H}^+] < 1.0 \times 10^{-7}$, so $\text{pH} > 7.00$.
- The higher the pH, the lower the pOH, the more basic the solution.

The pH Scale

Solution Type	[H ⁺] (M)	[OH ⁻] (M)	pH Value
Acidic	$>1.0 \times 10^{-7}$	$<1.0 \times 10^{-7}$	<7.00
Neutral	$=1.0 \times 10^{-7}$	$=1.0 \times 10^{-7}$	$=7.00$
Basic	$<1.0 \times 10^{-7}$	$>1.0 \times 10^{-7}$	>7.00

- In pure water,

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

- Because in pure water $[\text{H}_3\text{O}^+] = [\text{OH}^-]$,

$$[\text{H}_3\text{O}^+] = (1.0 \times 10^{-14})^{1/2} = 1.0 \times 10^{-7}$$

pH values of some common solutions

Note that a change in $[H^+]$ by a factor of 10 causes the pH to increase by 1

	$[H^+]$ (M)	pH	pOH	$[OH^-]$ (M)
	$1 (1 \times 10^{-0})$	0.0	14.0	1×10^{-14}
	1×10^{-1}	1.0	13.0	1×10^{-13}
Gastric juice	1×10^{-2}	2.0	12.0	1×10^{-12}
Lemon juice	1×10^{-3}	3.0	11.0	1×10^{-11}
Cola, vinegar	1×10^{-4}	4.0	10.0	1×10^{-10}
Wine	1×10^{-5}	5.0	9.0	1×10^{-9}
Tomatoes	1×10^{-6}	6.0	8.0	1×10^{-8}
Banana	1×10^{-7}	7.0	7.0	1×10^{-7}
Black coffee	1×10^{-8}	8.0	6.0	1×10^{-6}
Rain	1×10^{-9}	9.0	5.0	1×10^{-5}
Saliva	1×10^{-10}	10.0	4.0	1×10^{-4}
Milk	1×10^{-11}	11.0	3.0	1×10^{-3}
Human blood, tears	1×10^{-12}	12.0	2.0	1×10^{-2}
Egg white, seawater	1×10^{-13}	13.0	1.0	1×10^{-1}
Baking soda	1×10^{-14}	14.0	0.0	$1 (1 \times 10^{-0})$
Borax				
Milk of magnesia				
Lime water				
Household ammonia				
Household bleach				
NaOH, 0.1 M				

The pH Scale

pOH and other “p” scales

- The “p” in pH tells us to take the negative log of the quantity (in this case, hydrogen ions).
- We can express concentration of OH^- as pOH

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

$$\text{p}K_w = -\log K_w$$

The pH Scale

pOH and other “p” scales

Because at 25°C,

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

we know that

$$-\log [\text{H}_3\text{O}^+] + -\log [\text{OH}^-] = -\log K_w = 14.00$$

or, in other words,

$$\mathbf{pH + pOH = pK_w = 14.00} \text{ (at 25°C)}$$

The pH Scale

pOH and other “p” scales

Exercise

Indicate whether solutions with each of the following ion concentrations are neutral, acidic, or basic:

(a) $[\text{H}^+] = 4 \times 10^{-9} \text{ M};$

(b) $[\text{OH}^-] = 1 \times 10^{-7} \text{ M};$

(c) $[\text{OH}^-] = 7 \times 10^{-13} \text{ M}.$

The pH Scale

Exercise: What is the pH of an aqueous solution of $[\text{OH}^-] = 1.0 \times 10^{-3} \text{ M}$?

1. 2.22
2. 3.00
3. 7.00
4. 10.78
5. 11.00

$$\text{pOH} = -\log(1.0 \times 10^{-3}) = 3.00$$

$$\text{pH} = 14.00 - \text{pOH}$$

$$= 14.00 - 3.00 = 11.00$$

The pH Scale

Exercise:

If the pOH of a solution is 3.00, what is the pH of the solution? Is the solution acidic or basic?

$$\text{pH} + \text{pOH} = \text{p}K_w = 14.00$$

$$\text{pH} = 14.00 - 3.00 = 11.00$$

$$\text{pH} = 11.00 \quad \text{basic}$$









The pH Scale

Measuring pH

- For less accurate measurements, one can use
 - Litmus paper
 - “Red” paper turns blue above $\sim\text{pH} = 8$
 - “Blue” paper turns red below $\sim\text{pH} = 5$
 - An indicator
 - Certain dyes change color as pH changes. These are indicators.
 - Many indicators do not have a sharp color change as a function of pH.
 - Most indicators tend to be red in more acidic solutions.

The pH Scale

Measuring pH

	pH range for color change									
	0	2	4	6	8	10	12	14		
Methyl violet	Yellow		Violet							
Thymol blue	Red		Yellow		Yellow		Blue			
Methyl orange		Red		Yellow						
Methyl red			Red		Yellow					
Bromthymol blue				Yellow		Blue				
Phenolphthalein						Colorless		Pink		
Alizarin yellow R							Yellow		Red	

The pH Scale

Measuring pH

For more accurate measurements, one uses a pH meter, which measures the voltage in the solution.



Strong Acids and Bases

Strong acids

- Recall that strong acids e.g. HCl, HBr, HI, HNO₃, H₂SO₄, HClO₃, and HClO₄ dissociate completely to form H₃O⁺.
- By definition, these are strong electrolytes and exist totally as ions in aqueous solution.
- Thus for the monoprotic strong acids,

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

$$\text{pH} = -\log[\text{H}_3\text{O}^+]; \text{pH} + \text{pOH} = 14$$

Strong Acids and Bases

Strong acids

Exercise:

Calculate the pH of a solution made of 0.723g perchloric acid in 500mL of solution.

Strong Acids and Bases

Strong bases

- Strong bases also dissociate completely in aqueous solution.
- The pOH (and hence pH) of a strong base is given by the initial molarity of the base.
- In order for a hydroxide to be a base, it must be soluble.
- Bases do not have to contain the OH⁻ ion e.g. O²⁻
- As with strong acids, the calculation of pH for strong bases is simple.

$$[\text{base}] = [\text{OH}^-]$$

$$\text{pOH} = -\log[\text{OH}^-]; \text{pH} + \text{pOH} = 14$$

Strong Acids and Bases

Strong bases

Exercise

If 3.56g of pure NaOH is dissolved in 10 dm³ of water, what is the pH of the solution?

Weak Acids

- Weak acids are only partially ionized in solution.
- There is a mixture of ions and unionized acid in solution.
- Therefore, weak acids are in equilibrium:



$$K_a = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$



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

Weak Acids

Dissociation Constants

- For a generalized acid dissociation,



the equilibrium expression would be




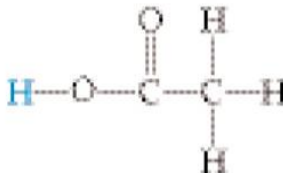

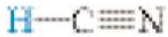
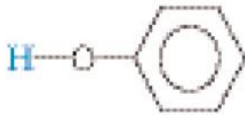
$$K_c = \frac{[\text{H}_3\text{O}^+][\text{A}^-]}{[\text{HA}]}$$

- This equilibrium constant is called the **acid-dissociation constant**, K_a .

Weak Acids

Dissociation Constants

The greater the value of K_a , the stronger the acid.

Acid	Structural Formula	Conjugate Base	Equilibrium Reaction	K_a
Hydrofluoric (HF)		F^-	$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$	6.8×10^{-4}
Nitrous (HNO_2)		NO_2^-	$HNO_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + NO_2^-(aq)$	4.5×10^{-4}
Benzoic ($HC_7H_5O_2$)		$C_7H_5O_2^-$	$HC_7H_5O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_7H_5O_2^-(aq)$	6.3×10^{-5}
Acetic ($HC_2H_3O_2$)		$C_2H_3O_2^-$	$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq)$	1.8×10^{-5}
Hypochlorous (HClO)		ClO^-	$HClO(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + ClO^-(aq)$	3.0×10^{-8}
Hydrocyanic (HCN)		CN^-	$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$	4.9×10^{-10}
Phenol (HC_6H_5O)		$C_6H_5O^-$	$HC_6H_5O(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_6H_5O^-(aq)$	1.3×10^{-10}

*The proton that ionizes is shown in blue.

- If $K_a \gg 1$, then the acid is completely ionized and the acid is a strong acid.

Weak Acids

Calculating K_a from the pH

- Calculate H_3O^+ and hence pH from K_a
 - K_a is usually small therefore use a approximation to simplify the problem.
 - Test the assumption made.
 - Use quadratic equation if “x” is out by more the 5% of the original acid concentration.
- Calculate K_a from pH

Weak Acids

Calculating K_a from the pH

Exercise:

The pH of a 0.10 M solution of formic acid, HCOOH, at 25 °C is 2.38. Calculate K_a for formic acid at this temperature.

- $\text{HCOOH}(\text{aq}) + \text{H}_2\text{O}(\text{aq}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{COO}^-(\text{aq})$
- To calculate K_a , we need the **equilibrium concentrations** of all three species.
- We can find $[\text{H}_3\text{O}^+]$, which is the same as $[\text{HCOO}^-]$, from the pH.

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

Weak Acids

Calculating K_a from the pH

Weak acids are simply equilibrium calculations.

- The pH gives the equilibrium concentration of H^+ .
- Using K_a , the concentration of H^+ (and hence the pH) can be calculated.
 - Write the balanced chemical equation clearly showing the equilibrium.
 - Write the equilibrium expression. Find the value for K_a .
 - Write down the initial and equilibrium concentrations for everything except pure water.

Weak Acids

Calculating K_a from the pH

$$K_a = \frac{[\text{H}_3\text{O}^+][\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$\text{pH} = -\log [\text{H}_3\text{O}^+] = 2.38$$

$$-2.38 = \log [\text{H}_3\text{O}^+]$$

$$[\text{H}_3\text{O}^+] = 10^{-2.38}$$

$$[\text{H}_3\text{O}^+] = [\text{HCOO}^-] = 4.2 \times 10^{-3} \text{ M}$$

Now we can set up a table...

Weak Acids

Calculating K_a from the pH



	[HCOOH], M	[H ₃ O ⁺], M	[HCOO ⁻], M
Initially	0.10	0	0
Change	-4.2×10^{-3}	$+4.2 \times 10^{-3}$	$+4.2 \times 10^{-3}$
At Equilibrium	$0.10 - 4.2 \times 10^{-3}$ $= 0.0958 \sim 0.10$	4.2×10^{-3}	4.2×10^{-3}

$$K_a = \frac{[4.2 \times 10^{-3}] [4.2 \times 10^{-3}]}{[0.10]} = 1.8 \times 10^{-4}$$

Weak Acids

Calculating $[\text{H}_3\text{O}^+]$ and pH from K_a

Example

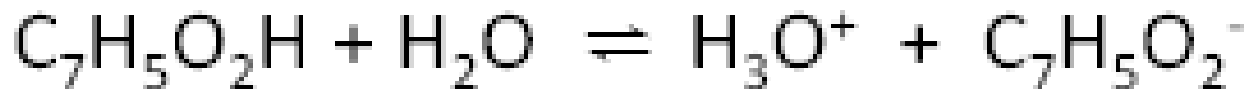
What is the $[\text{H}_3\text{O}^+]$ in a 0.20 *M* solution of benzoic acid? ($K_a = 6.3 \times 10^{-5}$)

Solution

- Benzoic acid, $\text{C}_7\text{H}_5\text{O}_2\text{H}$
- We have a weak acid – does not fully dissociate
- We construct a table

Weak Acids

Calculating $[H_3O^+]$ and pH from K_a



Initially	0.20 M	0 M	0 M
Change	- x	+ x	+ x
Equilibrium	0.20 - x M	x M	x M

$$K_a = \frac{[C_7H_5O_2^-][H_3O^+]}{[C_7H_5O_2H]} = 6.3 \times 10^{-5}$$

$$\frac{(x)(x)}{(0.20 - x)} = 6.3 \times 10^{-5}$$

Weak Acids

Calculating $[\text{H}_3\text{O}^+]$ and pH from K_a

This could be solved using a quadratic equation. But since this is a weak acid, we know that “x” is very small, in fact much smaller than 0.20 M.

Therefore it is assumed that $(0.20 - x) \cong 0.20$

Weak Acids

Calculating $[H_3O^+]$ and pH from K_a

$$\text{If } \frac{(x)(x)}{(0.20 - x)} = 6.3 \times 10^{-5}$$

and we assume $(0.20 - x) \approx 0.20$

$$\text{then, } \frac{x^2}{0.20} = 6.3 \times 10^{-5}$$

$$x = \sqrt{(0.20) \times (6.3 \times 10^{-5})}$$
$$= 3.5 \times 10^{-3} \text{ M}$$

Therefore, $[H_3O^+] = 3.5 \times 10^{-3} \text{ M}$

$$\text{pH} = 2.46$$

Weak Acids

Calculating $[\text{H}_3\text{O}^+]$ and pH from K_a

Test the assumption

$$\begin{aligned}(0.20 - X) &= 0.20 - (3.5 \times 10^{-3}) \\ &= 0.196 \sim 0.20 \text{ therefore assumption is valid}\end{aligned}$$

or

$$\frac{3.5 \times 10^{-3} \text{ M}}{0.20 \text{ M}} \times 100 = 1.8\%$$

Weak Acids

Calculating $[\text{H}_3\text{O}^+]$ and pH from K_a

If the *assumption is not valid*, one must use the quadratic equation:

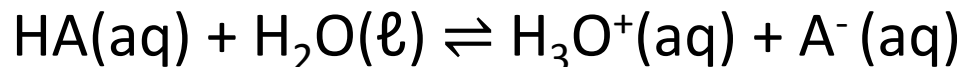
$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

Weak Acids

Calculating $[H_3O^+]$ and pH from K_a

Guidelines on how to tackle the previous exercise

- 1) Write the ionization equilibrium of benzoic acid. If you don't know the actual formula of the acid use the general formula:



- 2) Write the equilibrium constant expression using the given value of K_a .
- 3) Substitute the equilibrium concentrations into the equilibrium constant expression.
- 4) Simplify the equation by making an assumption.
- 5) Check the assumption.
- 6) If the quantity "x" is more than about 5% of the concentration of the acid, use the quadratic equation.

Weak Acids

Calculating $[\text{H}_3\text{O}^+]$ and pH from K_a

Exercise:

Calculate the pH of 0.10 mol L^{-1} HF solution ($K_a = 6.8 \times 10^{-4}$).

Weak Acids

Percent Ionization

- Magnitude of K_a indicates the strength of a weak acid.
- Likewise percent ionization is a measure of acid strength.

$$\text{Percent Ionization} = \frac{[\text{H}^+]_{\text{equilibrium}}}{[\text{HA}]_{\text{initial}}} \times 100\%$$

- In the previous example

$$[\text{H}^+]_{\text{eq}} = 4.2 \times 10^{-3} \text{ M}$$

$$[\text{HCOOH}]_{\text{initial}} = 0.10 \text{ M}$$

$$\begin{aligned} \text{Percent Ionization} &= \frac{4.2 \times 10^{-3}}{0.10} \times 100 \\ &= 4.2\% \end{aligned}$$

Weak Acids

Percent Ionization



(a)



(b)

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(a) The flask on the left contains $1\text{ M HC}_2\text{H}_3\text{O}_2$; the one on the right contains 1 M HCl . Each balloon contains the same amount of magnesium metal. (b) When the Mg metal is dropped into the acid, H_2 gas is formed. The rate of H_2 formation is higher for the 1 M HCl solution on the right as evidenced by more gas in the balloon.

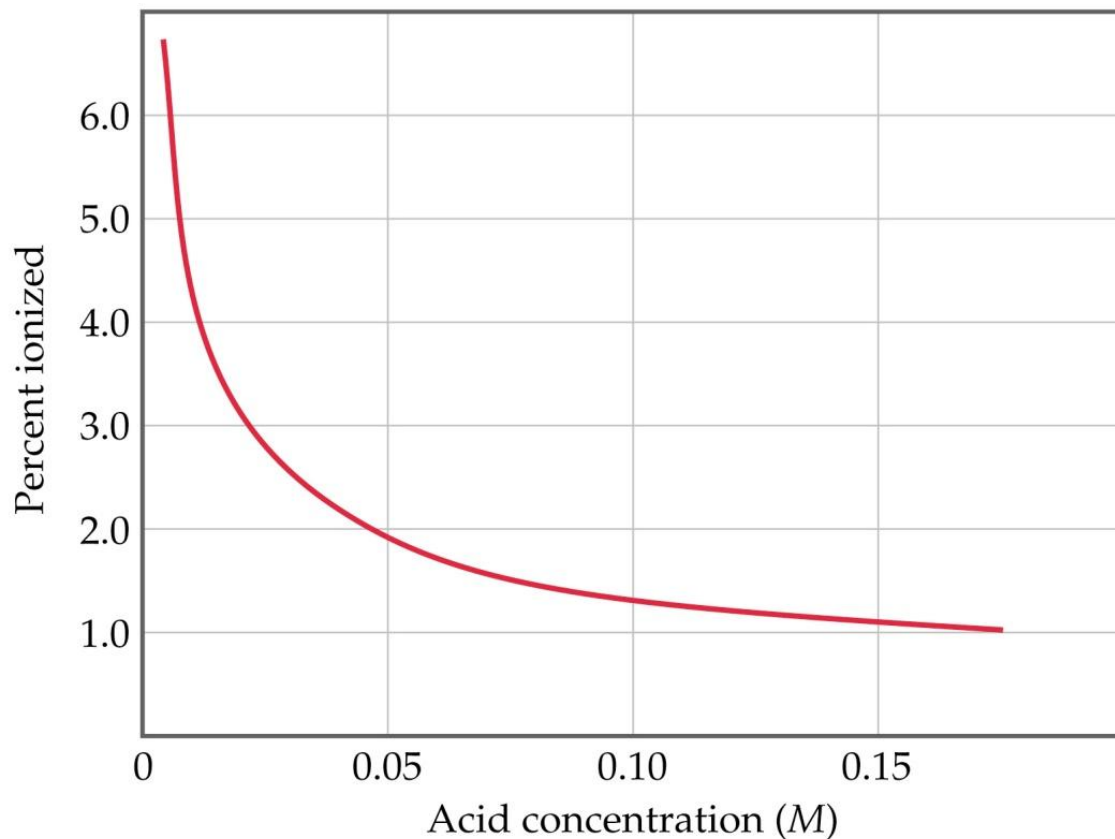
Weak Acids

Percent Ionization

- Percent ionization relates the equilibrium H^+ concentration, $[H^+]_{eqm}$, to the initial HA concentration, $[HA]_0$.
- **The higher percent ionization, the stronger the acid.**
- Percent ionization of a weak acid decreases as the molarity of the solution increases.

Weak Acids

Percent Ionization



For acetic acid 0.05 M solution is 2.0 % ionized whereas a 0.15 M solution is 1.0 % ionized.

Weak Acids

Polyprotic Acids

- Polyprotic acids have more than one ionizable H atom.
- The protons (H atoms) are removed in steps. For example:



- It is always easier to remove the first proton in a polyprotic acid than the second.
- Therefore, $K_{a1} > K_{a2} > K_{a3}$ etc.

Weak Acids

Polyprotic Acids

- If the difference between the K_a for the first dissociation and subsequent K_a values is 10^3 or more, the pH generally depends *only* on the **first dissociation**.

Name	Formula	K_{a1}	K_{a2}	K_{a3}
Ascorbic	$\text{H}_2\text{C}_6\text{H}_6\text{O}_6$	8.0×10^{-5}	1.6×10^{-12}	
Carbonic	H_2CO_3	4.3×10^{-7}	5.6×10^{-11}	
Citric	$\text{H}_3\text{C}_6\text{H}_5\text{O}_7$	7.4×10^{-4}	1.7×10^{-5}	4.0×10^{-7}
Oxalic	$\text{H}_2\text{C}_2\text{O}_4$	5.9×10^{-2}	6.4×10^{-5}	
Phosphoric	H_3PO_4	7.5×10^{-3}	6.2×10^{-8}	4.2×10^{-13}
Sulfurous	H_2SO_3	1.7×10^{-2}	6.4×10^{-8}	
Sulfuric	H_2SO_4	Large	1.2×10^{-2}	
Tartaric	$\text{H}_2\text{C}_4\text{H}_4\text{O}_6$	1.0×10^{-3}	4.6×10^{-5}	

Weak Acids

Summary so far

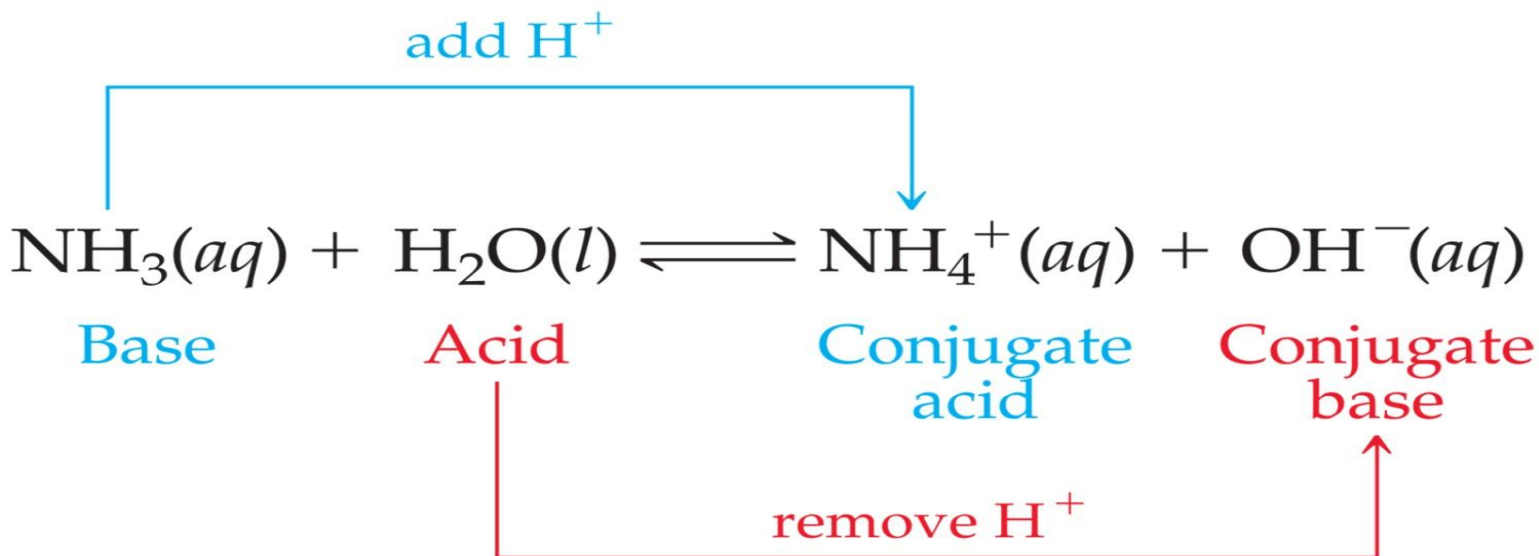
Calculate pH $\begin{cases} \rightarrow \text{strong acid} \rightarrow [\text{H}_3\text{O}^+] \rightarrow \text{pH} \\ \rightarrow \text{strong base} \rightarrow [\text{OH}^-] \rightarrow \text{pOH} \rightarrow \text{pH} \end{cases}$

Calculate pH – weak acid \rightarrow Equilibrium equation
 $\rightarrow K_a \rightarrow \text{pH}$
 $\left[\begin{array}{l} \rightarrow \text{Assumption} \\ \rightarrow \text{If not valid, use} \\ \quad \text{quadratic equation.} \end{array} \right.$
 \rightarrow Calculate K_a from pH.

weak bases
(next)

Weak Bases

- Weak bases react with water to remove protons from water.
- There is an equilibrium between the base and the resulting ions:



Weak Bases

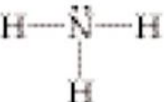

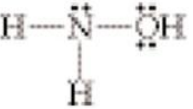
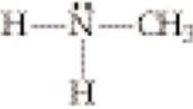
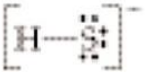
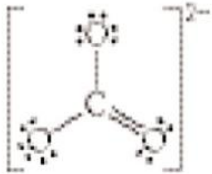

- The equilibrium constant expression for a weak base is

$$K_b = \frac{[\text{HB}] [\text{OH}^-]}{[\text{B}^-]}$$

where K_b is the base-dissociation constant.

K_b can be used to find $[\text{OH}^-]$ and, consequently, pH.

Weak Bases

Base	Lewis Structure	Conjugate Acid	Equilibrium Reaction	K_b
Ammonia (NH ₃)		NH ₄ ⁺	NH ₃ + H ₂ O ⇌ NH ₄ ⁺ + OH ⁻	1.8 × 10 ⁻⁵
Pyridine (C ₅ H ₅ N)		C ₅ H ₅ NH ⁺	C ₅ H ₅ N + H ₂ O ⇌ C ₅ H ₅ NH ⁺ + OH ⁻	1.7 × 10 ⁻⁹
Hydroxylamine (H ₂ NOH)		H ₃ NOH ⁺	H ₂ NOH + H ₂ O ⇌ H ₃ NOH ⁺ + OH ⁻	1.1 × 10 ⁻⁸
Methylamine (NH ₂ CH ₃)		NH ₃ CH ₃ ⁺	NH ₂ CH ₃ + H ₂ O ⇌ NH ₃ CH ₃ ⁺ + OH ⁻	4.4 × 10 ⁻⁴
Hydrosulfide ion (HS ⁻)		H ₂ S	HS ⁻ + H ₂ O ⇌ H ₂ S + OH ⁻	1.8 × 10 ⁻⁷
Carbonate ion (CO ₃ ²⁻)		HCO ₃ ⁻	CO ₃ ²⁻ + H ₂ O ⇌ HCO ₃ ⁻ + OH ⁻	1.8 × 10 ⁻⁴
Hypochlorite ion (ClO ⁻)		HClO	ClO ⁻ + H ₂ O ⇌ HClO + OH ⁻	3.3 × 10 ⁻⁷

Weak Bases

Types of Weak Bases

1. *Neutral (uncharged) substances that have an atom with a nonbonding pair of electrons that can accept a proton.*

Examples:

- Ammonia - :NH_3
- Amines

$\text{CH}_3\ddot{\text{N}}\text{H}_2$	primary amine
$\text{CH}_3\ddot{\text{N}}\text{HCH}_3$	secondary amine
$\begin{array}{c} \text{CH}_3-\ddot{\text{N}}-\text{CH}_3 \\ \\ \text{CH}_3 \end{array}$	tertiary amine

Weak Bases

Types of Weak Bases

2. *The conjugate bases (anions) of weak acids.*

- Acetate ion, CH_3COO^- – conjugate base of acetic acid.

Example: in sodium acetate, CH_3COONa .

- Hypochlorite ion, ClO^- – conjugate base of hypochlorous acid.

Example: in calcium hypochlorite, $\text{Ca}(\text{ClO})_2$

- Carbonates, bicarbonates, phosphates.

3. *Alkaloids.*

Naturally occurring basic substances

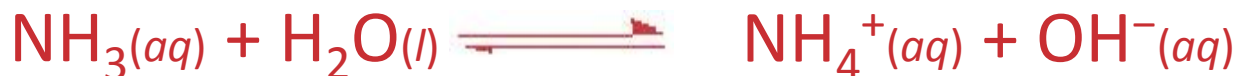
e.g. cocaine, ephedrine, quinine, mescaline, morphine

Weak Bases

pH of Basic Solutions

Example: What is the pH of a 0.15 M solution of NH_3 ?

Solution:



$$K_b = \frac{[\text{NH}_4^+][\text{OH}^-]}{[\text{NH}_3]} = 1.8 \times 10^{-5}$$

Tabulate the data...

Weak Bases

pH of Basic Solutions

	$[\text{NH}_3], M$	$[\text{NH}_4^+], M$	$[\text{OH}^-], M$
Initially	0.15	0	0
Change	- x	+ x	+ x
At Equilibrium	$0.15 - x \approx 0.15$	x	x

Weak Bases

pH of Basic Solutions

$$1.8 \times 10^{-5} = \frac{(x)^2}{(0.15)}$$

$$(1.8 \times 10^{-5}) (0.15) = x^2$$

$$2.7 \times 10^{-6} = x^2, \quad x = 1.6 \times 10^{-3}$$

Test the assumption:

$$\begin{aligned} \text{Percent ionization} &= (0.0016/0.15) \times 100 \% \\ &= 1.07\% \end{aligned}$$

$$\text{Or } 0.15 - x = 0.15 - 0.0016 = 0.1484$$

Weak Bases

pH of Basic Solutions

Therefore,

$$[\text{OH}^-] = 1.6 \times 10^{-3} \text{ M}$$

$$\text{pOH} = -\log (1.6 \times 10^{-3})$$

$$\text{pOH} = 2.80$$

$$\text{pH} = 14.00 - 2.80$$

$$\text{pH} = 11.20$$

Relationship between K_a and K_b

- When two reactions are added to give a third, the equilibrium constant for the third reaction is the product of the equilibrium constants for the first two:

Reaction 1 + Reaction 2 = Reaction 3

has

$$K_3 = K_1 \times K_2$$

Relationship between K_a and K_b

- K_a and K_b are related in this way:

$$K_a \times K_b = K_w$$

- Therefore, if you know one of them, you can calculate the other.

- For a conjugate acid-base pair $K_w = K_a \times K_b$
- The larger the K_a , the smaller the K_b . That is, the stronger the acid, the weaker the conjugate base.
- In terms of pK_a and pK_b

$$pK_w = pK_a + pK_b = 14 \quad (\text{at } 25^\circ\text{C})$$

Relationship between K_a and K_b

Acid	K_a	Base	K_b
HNO ₃	(Strong acid)	NO ₃ ⁻	(Negligible basicity)
HF	6.8×10^{-4}	F ⁻	1.5×10^{-11}
HC ₂ H ₃ O ₂	1.8×10^{-5}	C ₂ H ₃ O ₂ ⁻	5.6×10^{-10}
H ₂ CO ₃	4.3×10^{-7}	HCO ₃ ⁻	2.3×10^{-8}
NH ₄ ⁺	5.6×10^{-10}	NH ₃	1.8×10^{-5}
HCO ₃ ⁻	5.6×10^{-11}	CO ₃ ²⁻	1.8×10^{-4}
OH ⁻	(Negligible acidity)	O ²⁻	(Strong base)

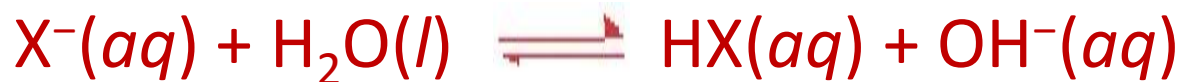
Acid-Base Properties of Solutions

- Nearly all salts are strong electrolytes.
- Therefore, salts exist entirely of ions in solution.
- Acid-base properties of salts are a consequence of the reaction of their *ions* in solution.
- The reaction in which ions produce H^+ or OH^- in water is called hydrolysis.
- Anions from weak acids are basic.
- Anions from strong acids are neutral.

Acid-Base Properties of Solutions

Reactions of Anions with Water

- In general, anions, X^- , are bases in solution i.e. can be considered as the conjugate base of an acid.
- E.g. Cl^- is the conjugate base of HCl , CH_3COO^- is the conjugate base of CH_3COOH .
- As such, anions can react with water in a **hydrolysis reaction** to form OH^- and the conjugate acid:



Acid-Base properties of Solutions

Reactions of Cations with Water

- Cations with acidic protons (like NH_4^+) will lower the pH of a solution.
- They can be considered as the conjugate acids of weak bases.
- $\text{NH}_4^+(\text{aq}) + \text{H}_2\text{O} (\text{l}) \rightleftharpoons \text{NH}_3(\text{aq}) + \text{H}_3\text{O}^+(\text{aq})$
- Most metal cations that are hydrated in solution also lower the pH of the solution.

Acid-Base Properties of Solutions

Combined Effect of Cation and Anion in Solution



1. An anion that is the conjugate base of a strong acid will not affect the pH. Example, Br^-
2. An anion that is the conjugate base of a weak acid will increase the pH. Example, CN^-
3. A cation that is the conjugate acid of a weak base will decrease the pH. Example, CH_3NH_3^+

Acid-Base Properties of Solutions

Combined Effect of Cation and Anion in Solution

4. Cations of the strong Arrhenius bases will not affect the pH.
5. Other metal ions will cause a decrease in pH.
6. When a solution contains both the conjugate base of a weak acid and the conjugate acid of a weak base, the affect on pH depends on the K_a and K_b values.



(a)



(b)



(c)

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Acid-Base Behaviour and Chemical Structure

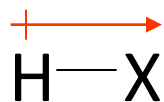
- In this section we discuss briefly the effects of chemical structure on acid base behavior.
- An acid must contain at least one H bonded to an X atom, i.e. H-X
- Three factors affect the strengths of acids

Acid-Base Behaviour and Chemical Structure

Factors that Affect Acid Strength

(i) Polarity of the H-X

- H-X bond must be polar with $H\delta^+$ and $X\delta^-$



- The more polar the bond, the stronger the acid.
- If X is a metal then the bond polarity is $H\delta^-$, $X\delta^+$ and the substance is a base, the H-X bond must be weak enough to be broken, the conjugate base, X^- , must be stable.

Acid-Base Behaviour and Chemical Structure

Factors that Affect Acid Strength

(ii) The Strength of the bond

- Release of a proton implies the breaking of the X-H bond.
- The weaker the bond stronger the acid.
- The H-F bond is the most polar. However it has the highest bond strength amongst the hydrogen halides, hence it is a *weak acid*.

Acid-Base Behaviour and Chemical Structure

Factors that Affect Acid Strength

(iii) Stability of the conjugate base, X^-

- The conjugate base, X^- , must be stable.
- Any feature that lowers the energy of the conjugate base in comparison with the acid increases its stability.
- The greater the stability of the conjugate base, the stronger the acid.

The strength of an acid is often a combination of the polarity of the H-X bond, the strength of the H-X bond, and the stability of the conjugate base, X^- .

Acid-Base Behaviour and Chemical Structure

Binary Acids

- Binary acids contain H and just one other element. E.g. HF, H₂S, HI, HCl, etc.
- The more polar the H-X bond and/or the weaker the H-X bond, the more acidic the compound.

	GROUP			
	4A	5A	6A	7A
Period 2	CH ₄ No acid or base properties	NH ₃ Weak base	H ₂ O ---	HF Weak acid
Period 3	SiH ₄ No acid or base properties	PH ₃ Weak base	H ₂ S Weak acid	HCl Strong acid

Increasing acid strength

Increasing base strength

Increasing acid strength

Increasing base strength

Acid-Base Behaviour and Chemical Structure

Binary Acids

Which acid is the strongest?

- H_2S
- HF
- HCl
- HBr
- HI

3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
5 B	6 C	7 N	8 O	9 F	10 Ne
13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
113	114	115	116		

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Acid-Base Behaviour and Chemical Structure

Binary Acids

Which base is the strongest?

- ClO_4^-
- BrO_3^-
- BrO_4^-
- IO_3^-
- IO_4^-

	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
	5 B	6 C	7 N	8 O	9 F	10 Ne
	13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
	113	114	115	116		

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Acid-Base Behaviour and Chemical Structure

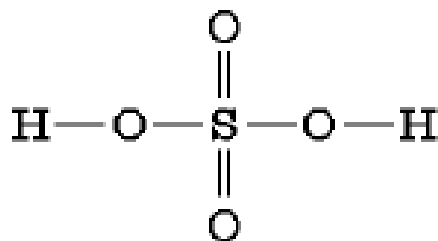
Binary Acids

- Acid strength increases across a period and down a group.
- Conversely, base strength decreases across a period and down a group.
- HF is a weak acid because the bond energy is high.
- The electronegativity difference between C and H is so small that the C-H bond is non-polar and CH₄ is neither an acid nor a base.

Acid-Base Behaviour and Chemical Structure

Oxyacids

- Oxyacids contain O-H bonds (and often also =O groups).
- Example: sulphuric acid.



- All oxyacids have the general structure Y-O-H.
- The strength of the acid depends on Y and the atoms attached to Y. The OH group/s in these molecules can ionise as follows:



Acid-Base Behaviour and Chemical Structure

Oxyacids

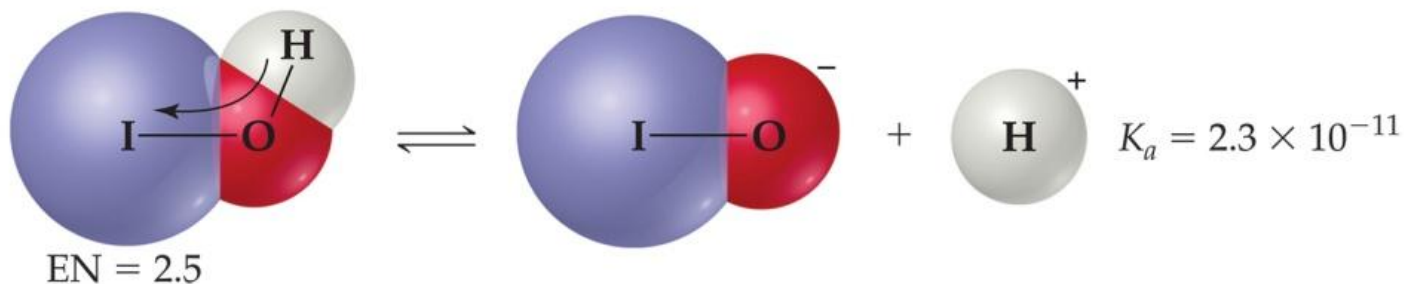
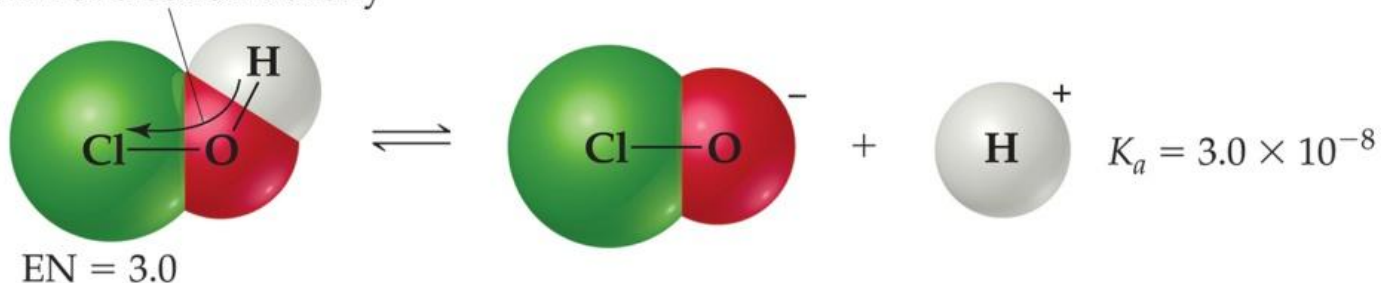
- This only happens in practice if the atom Y is a strongly electronegative atom, e.g., P, S, N, etc. Atom Y can then attract electron density towards itself, thereby polarising the O-H bond and facilitating ionisation:
- If Y is a metal (low electronegativity), then the substances are bases.
- If Y has intermediate electronegativity (e.g. I, EN = 2.5), the electrons are between Y and O and the substance is a weak oxyacid.

Acid-Base Behaviour and Chemical Structure

Oxyacids

- If Y has a large electronegativity (e.g. Cl, EN = 3.0), the electrons are located closer to Y than O and the O-H bond is polarized to lose H⁺.

Shift of electron density



Acid-Base Behaviour and Chemical Structure

Oxyacids

Summary

1. For oxyacids that have the same number of OH groups and the same number of O atoms, acid strength increases with increasing electronegativity of the central atom.

H-O-Y

Acid	EN of Y	K_a
HClO	3.0	3.0×10^{-8}
HBrO	2.8	2.5×10^{-9}
HIO	2.5	2.3×10^{-11}

Acid-Base Behaviour and Chemical Structure

Oxyacids

Summary

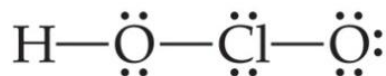
2. For a series of oxyacids that have the same central atom Y, acidity increases with the number of oxygen atoms attached to Y increases.

Hypochlorous



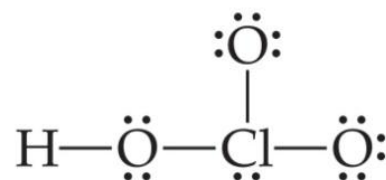
$$K_a = 3.0 \times 10^{-8}$$

Chlorous



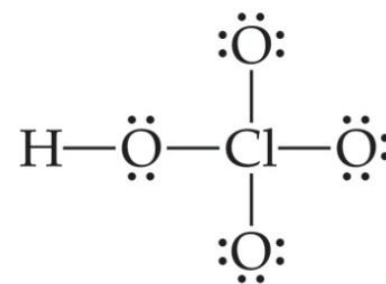
$$K_a = 1.1 \times 10^{-2}$$

Chloric



Strong acid

Perchloric



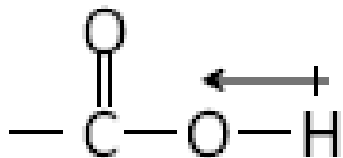
Strong acid

Increasing acid strength 

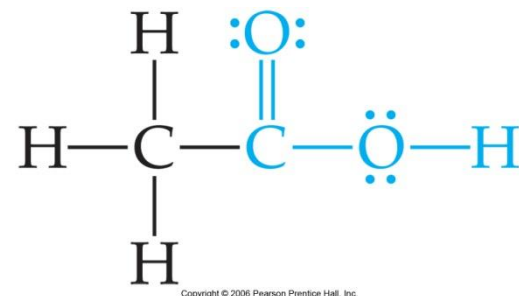
Acid-Base Behaviour and Chemical Structure

Carboxylic Acids

- Carboxylic acids are organic acids that contain the carboxyl group:

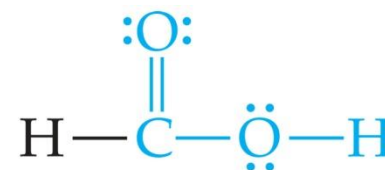


- Often written as —COOH or as $\text{—CO}_2\text{H}$.
- All carboxylic acids are weak acids.
- When the carboxylic acid loses a proton, it generates the carboxylate anion, COO^- .

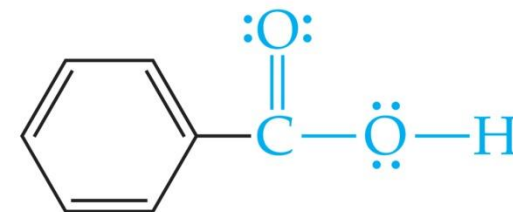


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Acetic acid



Formic acid



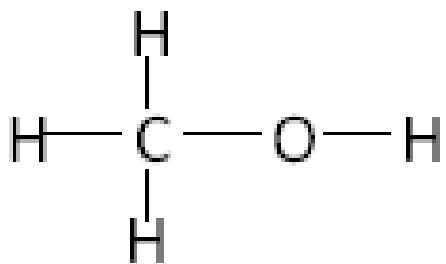
Benzoic acid

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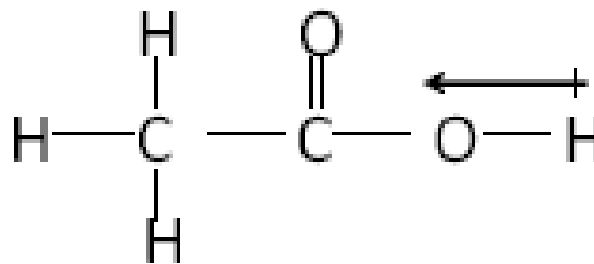
Acid-Base Behaviour and Chemical Structure

Carboxylic Acids

- It is of interest to note that methanol, CH_3OH , does not have acidic properties in water, while acetic acid [CH_3COOH] does. Why?
- Compare the molecular structures:



methanol



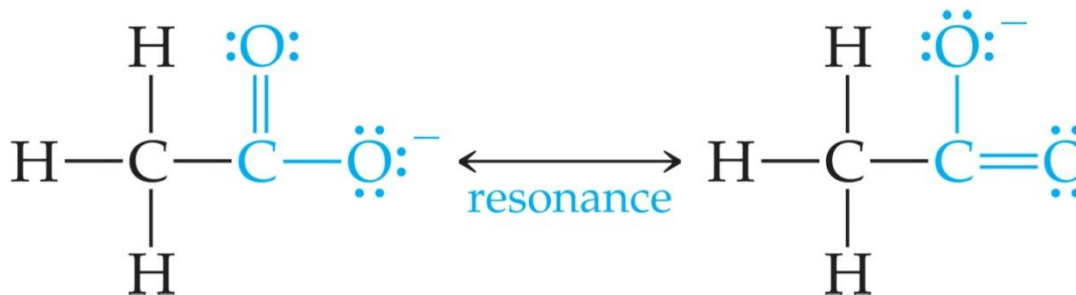
acetic acid

Acid-Base Behaviour and Chemical Structure

Carboxylic Acids

Two factors contribute to acidic behaviour of carboxylic acids:

1. Additional O atom attached to the carbon attracts electron density from the O-H bond increasing its polarity and stabilising the conjugate base
2. Resonance in the conjugate bases of carboxylic acids stabilizes the base and makes the conjugate acid more acidic.



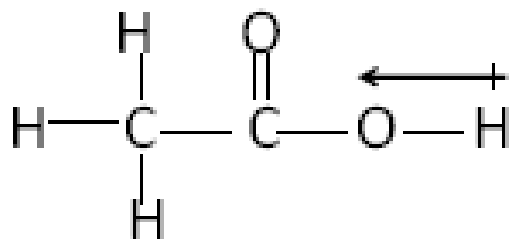
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Acid-Base Behaviour and Chemical Structure

Carboxylic Acids

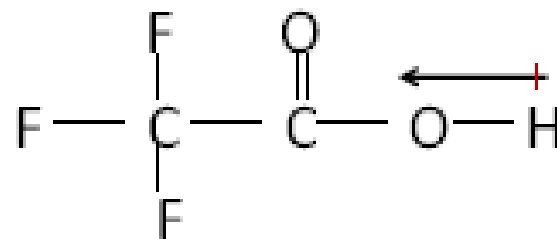
- The acid strength of carboxylic acids also increases as the number of electronegative atoms in the molecule increases.

- Example:



acetic acid

$$K_a = 1.8 \times 10^{-5}$$



Trifluoroacetic acid

$$K_a = 0.50$$

Lewis Acids and Bases

- In previous sections we discussed Brønsted-Lowry acids and bases.
- In the Brønsted-Lowry view, an acid is a proton donor. When a proton is donated to another molecule, where does it actually go on the acceptor molecule ?
- It invariably goes onto a lone pair of electrons situated on some atom (e.g., a N or O-) belonging to the base.

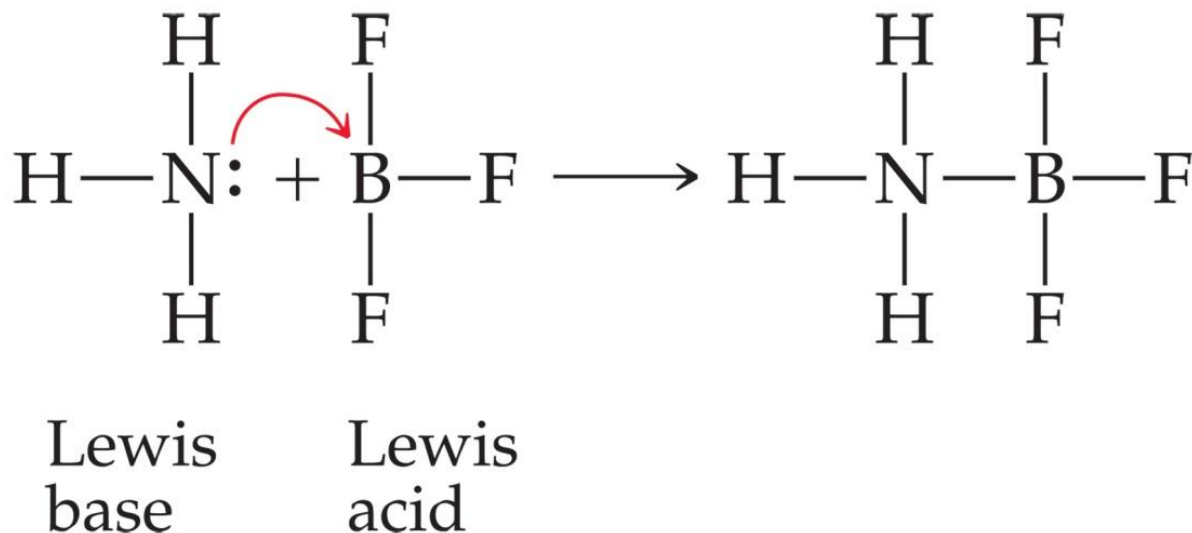
Lewis Acids and Bases

- G.N. Lewis proposed that what was characteristic of acids and bases wasn't so much that a proton is always exchanged between an acid and a base, but that a lone pair of electrons is always involved.
- This led to the definition of Lewis acids and bases.
- According to the Lewis view:

A Lewis acid is an electron pair acceptor.
A Lewis base is electron pair donor.

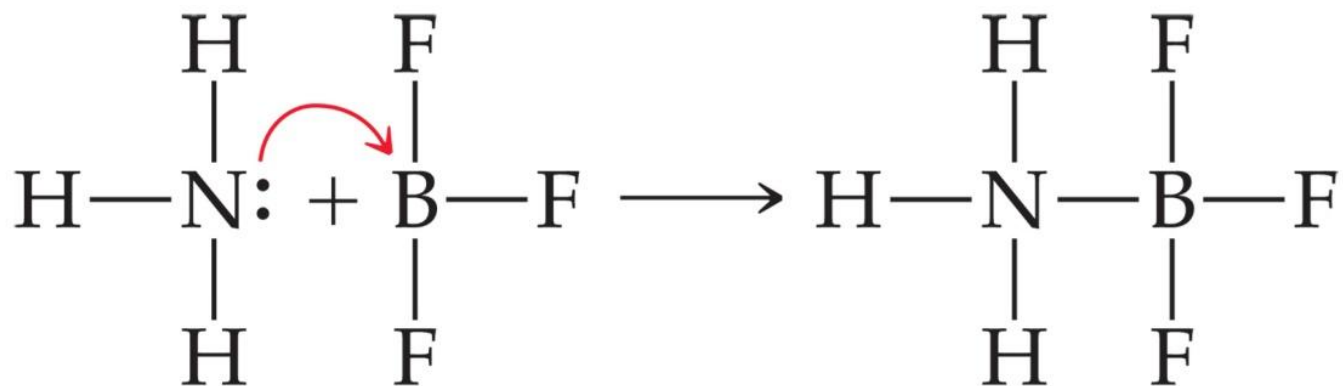
- Note: Lewis acids and bases do not need to contain protons. Therefore, the Lewis definition is the most general definition of acids and bases.

Lewis Acids



- Lewis acids are defined as electron-pair acceptors.
- Atoms with an empty valence orbital can be Lewis acids.
- Lewis acids generally have an incomplete octet e.g. BF_3
- Compounds with π -bonds can act as Lewis acids.

Lewis Bases



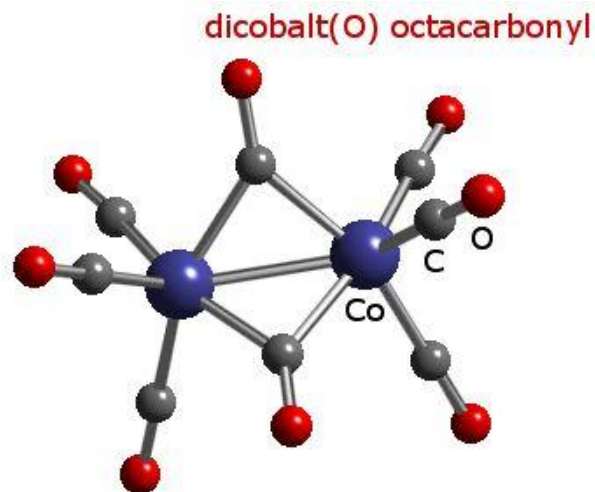
Lewis
base

Lewis
acid

- Lewis bases are defined as electron-pair donors.
- Anything that could be a Brønsted–Lowry base is a Lewis base.
- Lewis bases can interact with things other than protons, however.

Metal Complexes

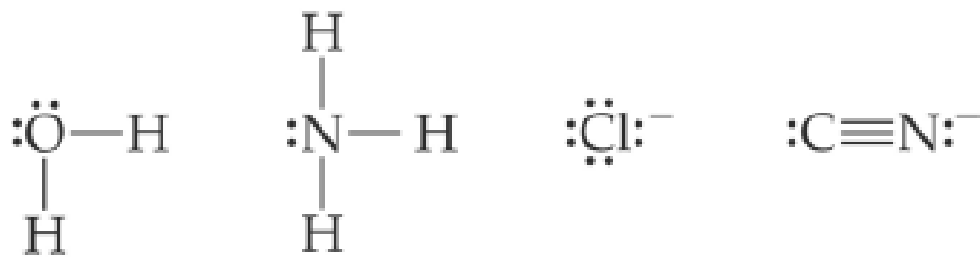
- A central metal atom bonded to a group of molecules or ions is a **metal complex** e.g. $[\text{Ag}(\text{NH}_3)_2]^+$.
- If the complex bears a charge, it is a **complex ion**.
- Compounds containing complexes are **coordination compounds**.



Metal Complexes

The metal-ligand bond

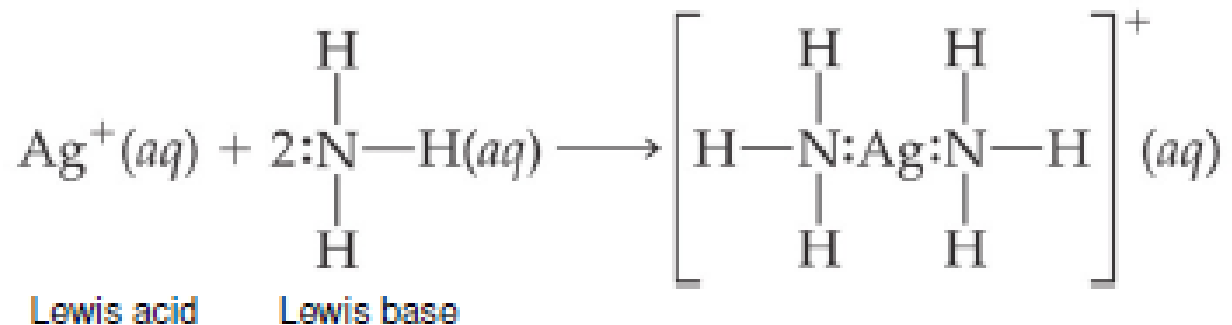
- The molecules or ions that bond to the metal ion in a complex are known as **ligands**.
- They are usually anions or polar molecules. They have at least one unshared pair of valence electrons.
- Examples of ligands:



Metal Complexes

The metal-ligand bond

- The bond between a ligand and a metal ion is an example of an interaction between a Lewis base and acid.
- The ligands have nonbonding electrons and are Lewis bases.
- The metal has empty orbitals and is a Lewis acid.
- A **metal-ligand bond** is formed between the Lewis acid (metal) and the Lewis base (ligand).



End of Chapter 16



Acid-Base Equilibria