



Acids and Bases

Acids

chua

họ cam quýt

Have a **sour taste**. **Vinegar** owes its taste to acetic acid. **Citrus** fruits contain citric acid.

React with certain **metals** to produce **hydrogen gas**.

H₂

React with **carbonates** and **bicarbonates** to produce **carbon dioxide gas**.

CO₂

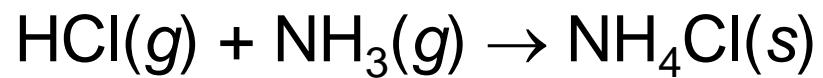
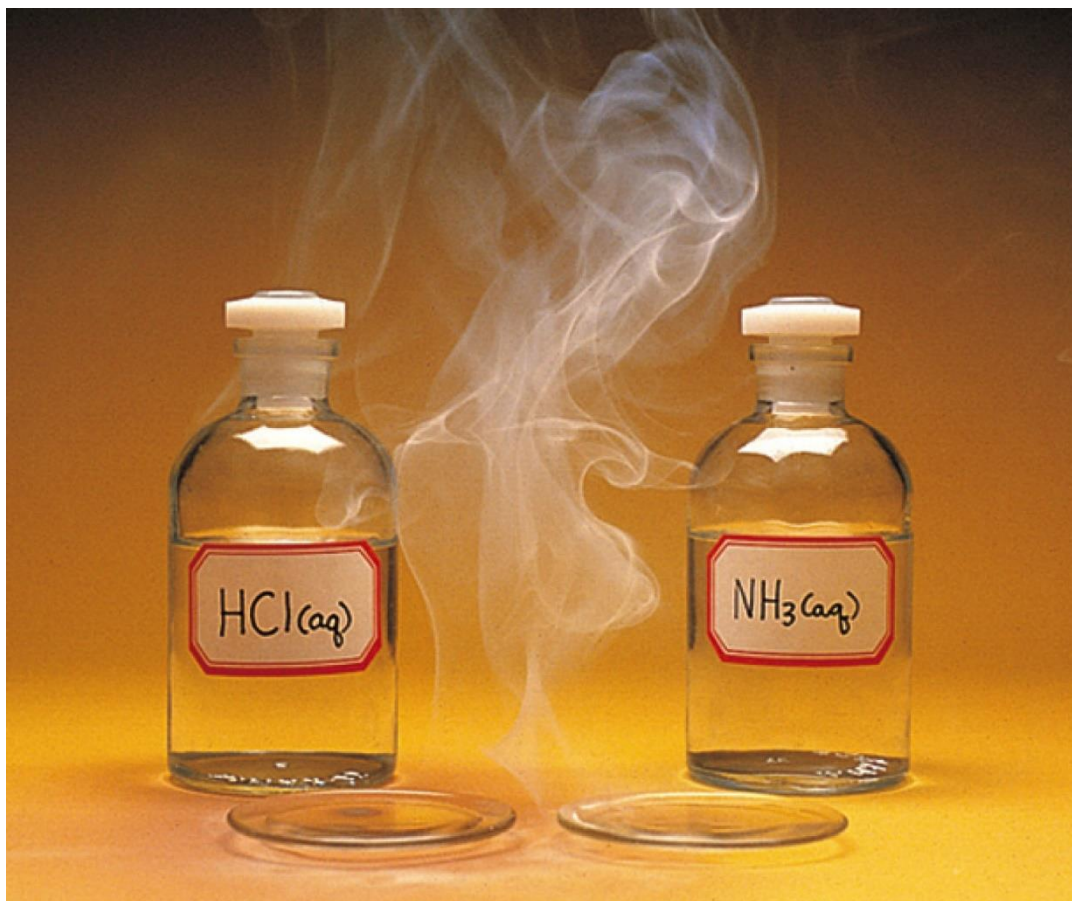
Bases

đắng

Have a **bitter taste**.

Feel **slippery**. Many **soaps** contain bases.

Acid - Base



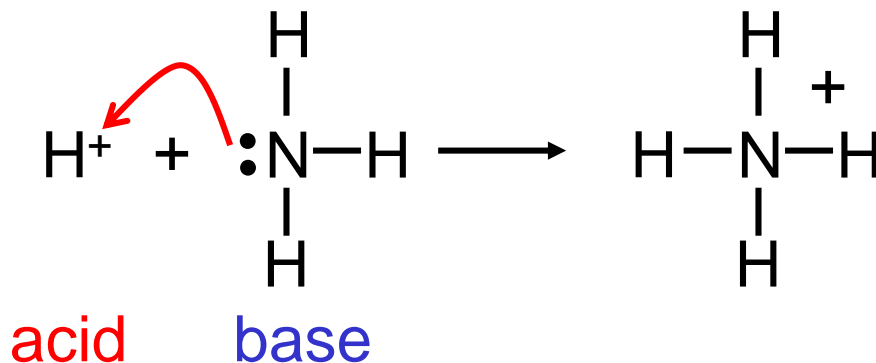
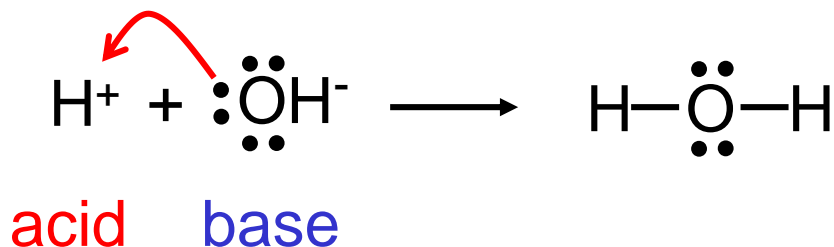
Definition of An Acid

Arrhenius acid is a substance that produces H^+ (H_3O^+) in water

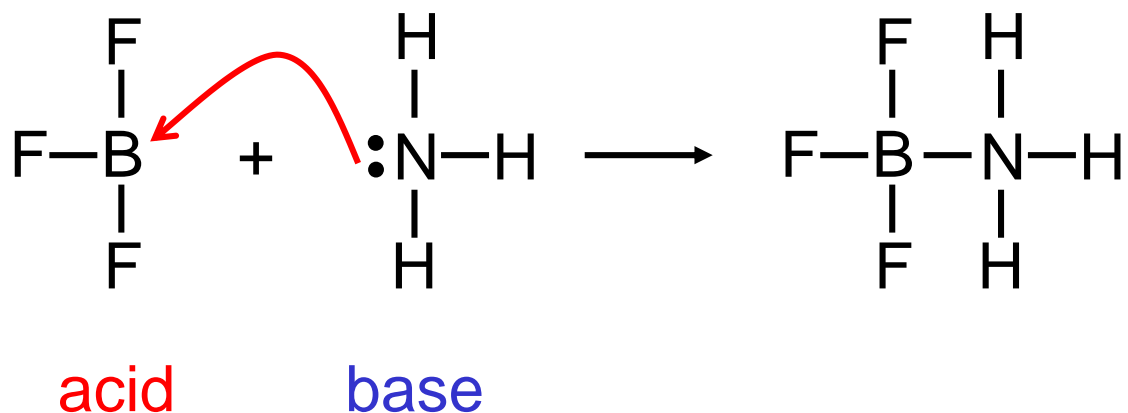
A **Brønsted acid** is a proton donor

A **Lewis acid** is a substance that can accept a pair of electrons

A **Lewis base** is a substance that can donate a pair of electrons



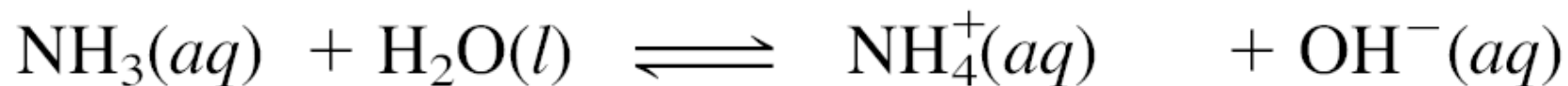
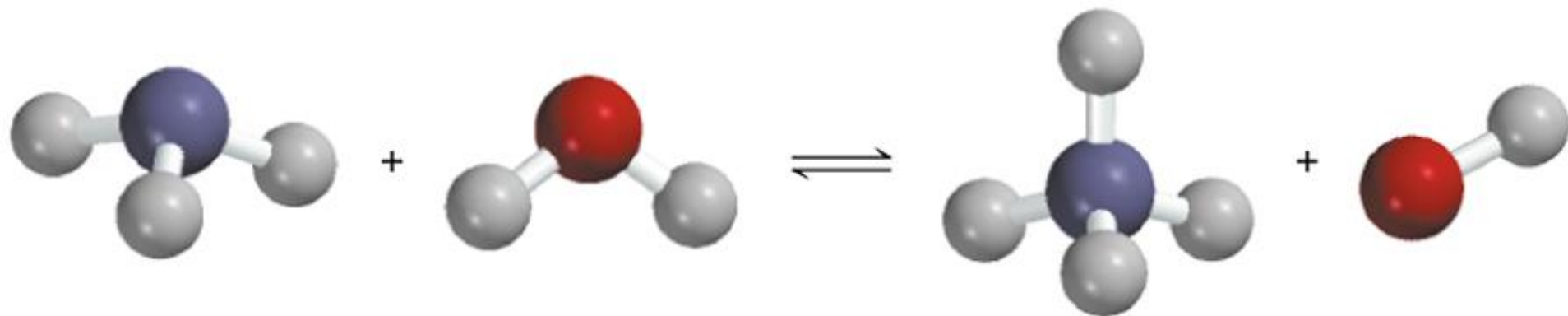
Lewis Acids and Bases



No protons donated or accepted!

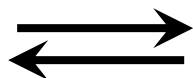
A Brønsted **acid** is a proton donor

A Brønsted **base** is a proton acceptor



base

acid



acid

base

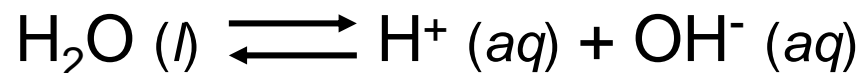
base

acid

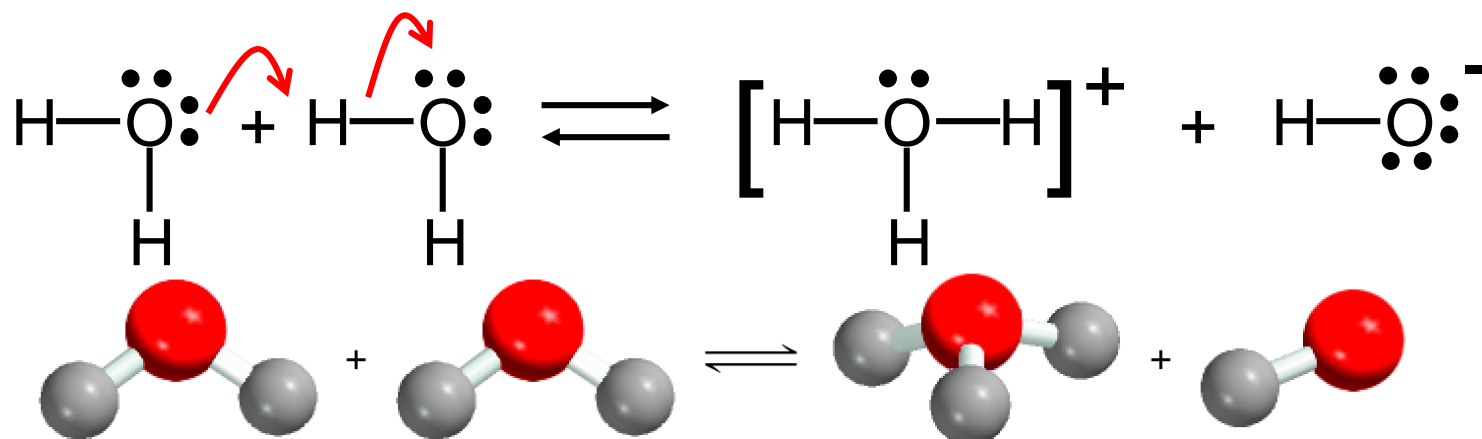
conjugate
acid

conjugate
base

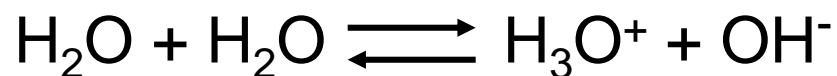
Acid-Base Properties of Water



autoionization of water



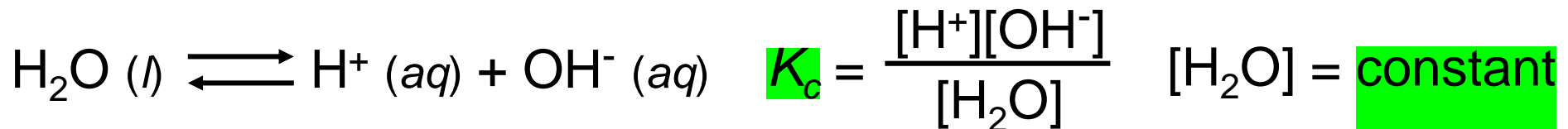
base conjugate
acid



acid

conjugate
base

The Ion Product of Water



$$K_c[\text{H}_2\text{O}] = K_w = [\text{H}^+][\text{OH}^-]$$

The **ion-product constant** (K_w) is the product of the molar concentrations of H^+ and OH^- ions **at a particular temperature**.

Solution Is

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

neutral

$$[\text{H}^+] > [\text{OH}^-]$$

acidic

$$[\text{H}^+] < [\text{OH}^-]$$

basic

At 25 °C

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

What is the **concentration** of OH⁻ ions in a HCl solution whose hydrogen ion concentration is 1.3 *M*?

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

$$[\text{H}^+] = 1.3 \text{ } M$$

$$[\text{OH}^-] = \frac{K_w}{[\text{H}^+]} = \frac{1 \times 10^{-14}}{1.3} = 7.7 \times 10^{-15} \text{ } M$$

pH – A Measure of Acidity

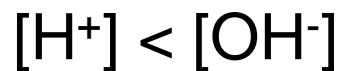
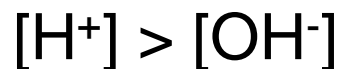
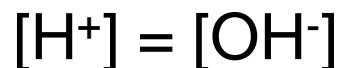
$$\text{pH} = -\log [\text{H}^+]$$

Solution is

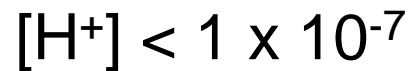
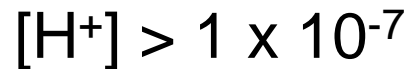
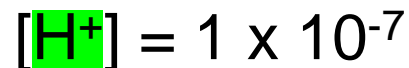
neutral

acidic

basic



At 25 °C



$$\text{pH} = 7$$

$$\text{pH} < 7$$

$$\text{pH} > 7$$

pH ↑

[H⁺] ↓

TABLE 15.1**The pHs of Some Common Fluids**

Sample	pH Value
Gastric juice in the stomach	1.0–2.0
Lemon juice	2.4
Vinegar	3.0
Grapefruit juice	3.2
Orange juice	3.5
Urine	4.8–7.5
Water exposed to air*	5.5
Saliva	6.4–6.9
Milk	6.5
Pure water	7.0
Blood	7.35–7.45
Tears	7.4
Milk of magnesia	10.6
Household ammonia	11.5

*Water exposed to air for a long period of time absorbs atmospheric CO_2 to form carbonic acid, H_2CO_3 .

Other important relationships

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

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

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

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

Measure of pH



pH meter



pH paper



Red cabbage



Litmus paper

The pH of rainwater collected in a certain region of the northeastern United States on a particular day was 4.82. What is the H^+ ion concentration of the rainwater?

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

$$[\text{H}^+] = 10^{-\text{pH}} = 10^{-4.82} = 1.5 \times 10^{-5} \text{ M}$$

The OH^- ion concentration of a blood sample is $2.5 \times 10^{-7} \text{ M}$. What is the pH of the blood?

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

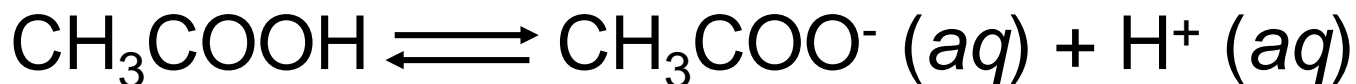
$$\text{pOH} = -\log [\text{OH}^-] = -\log (2.5 \times 10^{-7}) = 6.60$$

$$\text{pH} = 14.00 - \text{pOH} = 14.00 - 6.60 = 7.40$$

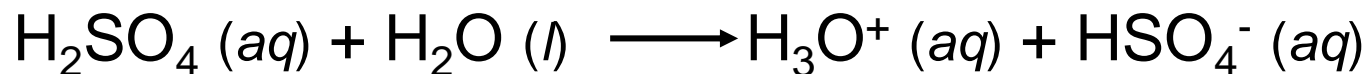
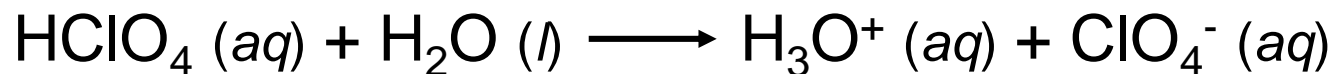
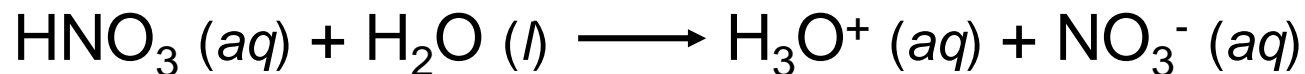
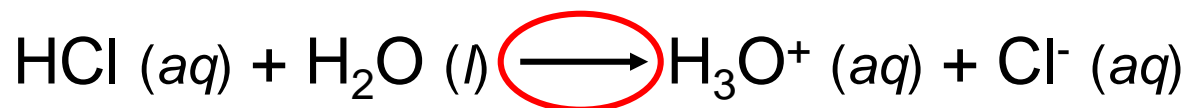
Strong Electrolyte – **100%** dissociation



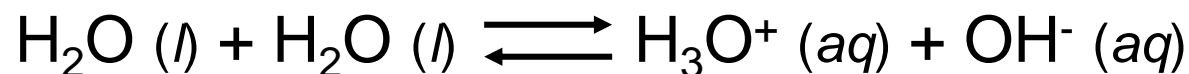
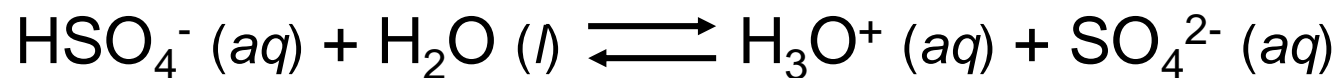
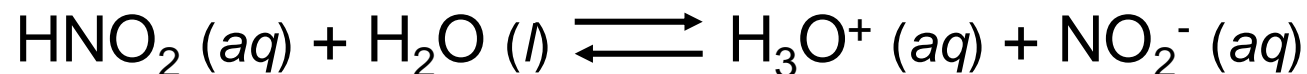
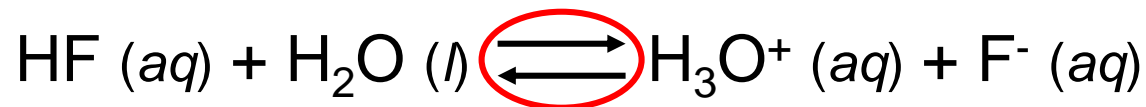
Weak Electrolyte – **not completely** dissociated



Strong Acids are **strong** electrolytes



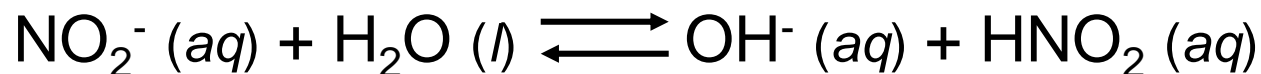
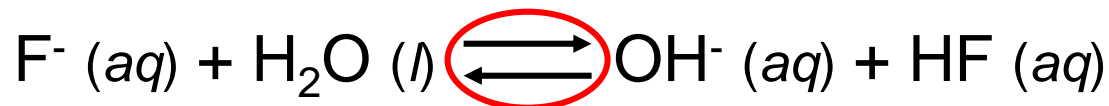
Weak Acids are **weak** electrolytes



Strong Bases are **strong** electrolytes



Weak Bases are **weak** electrolytes



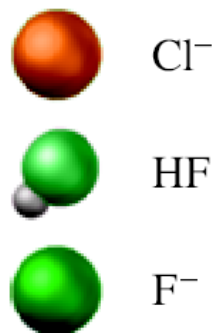
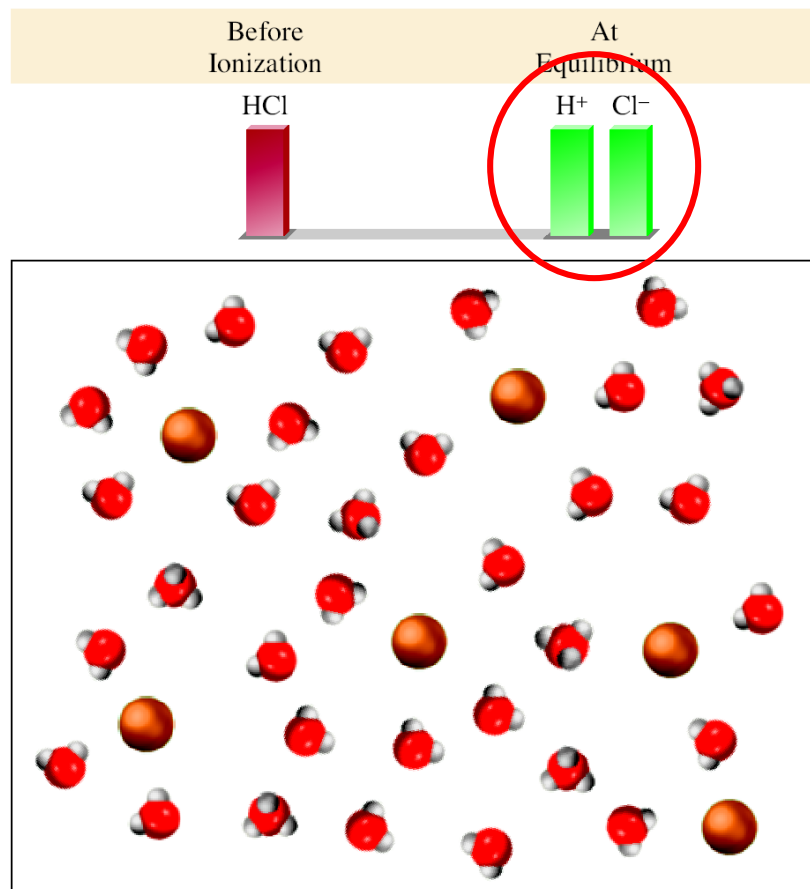
Conjugate acid-base pairs:

- The **conjugate base** of a **strong acid** has **no measurable strength**.
- H_3O^+ is the **strongest** acid that can exist in **aqueous solution**.
- The OH^- ion is the **strongest** base that can exist in **aqueous solution**.

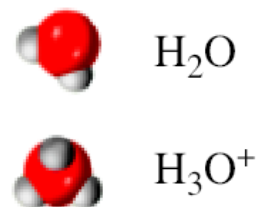
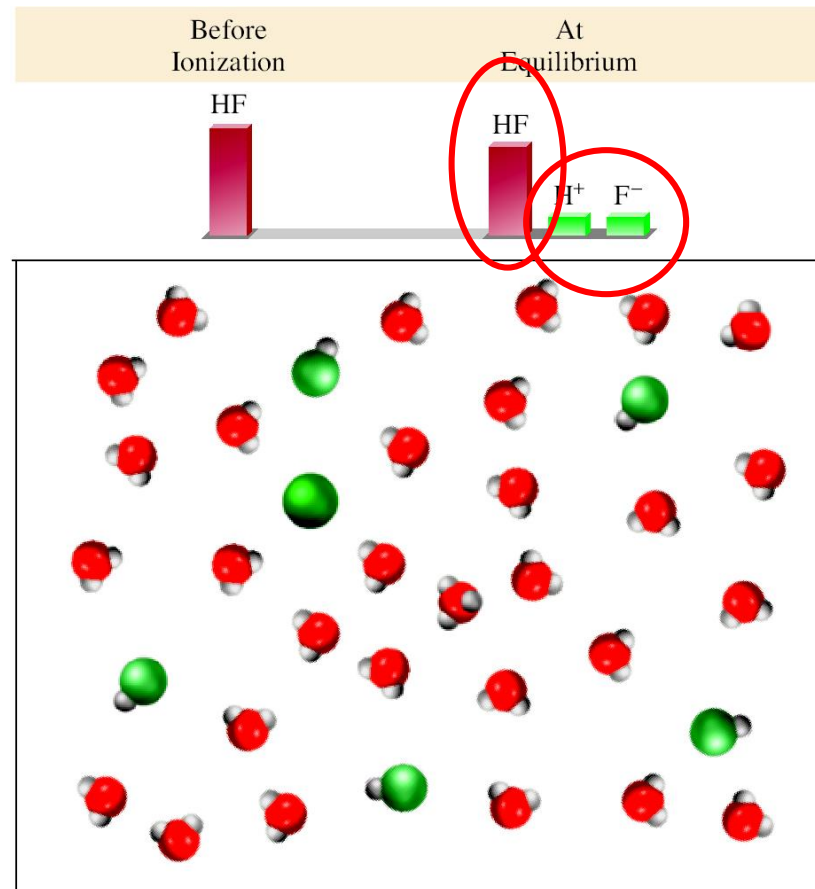
TABLE 15.2 Relative Strengths of Conjugate Acid-Base Pairs

		Acid	Conjugate Base		
Acid strength increases ↑	Strong acids	HClO ₄ (perchloric acid)	ClO ₄ ⁻ (perchlorate ion)	↓ Base strength increases	
		HI (hydroiodic acid)	I ⁻ (iodide ion)		
		HBr (hydrobromic acid)	Br ⁻ (bromide ion)		
		HCl (hydrochloric acid)	Cl ⁻ (chloride ion)		
		H ₂ SO ₄ (sulfuric acid)	HSO ₄ ⁻ (hydrogen sulfate ion)		
		HNO ₃ (nitric acid)	NO ₃ ⁻ (nitrate ion)		
		H ₃ O ⁺ (hydronium ion)	H ₂ O (water)		
	Weak acids	HSO ₄ ⁻ (hydrogen sulfate ion)	SO ₄ ²⁻ (sulfate ion)		
		HF (hydrofluoric acid)	F ⁻ (fluoride ion)		
		HNO ₂ (nitrous acid)	NO ₂ ⁻ (nitrite ion)		
		HCOOH (formic acid)	HCOO ⁻ (formate ion)		
		CH ₃ COOH (acetic acid)	CH ₃ COO ⁻ (acetate ion)		
		NH ₄ ⁺ (ammonium ion)	NH ₃ (ammonia)		
		HCN (hydrocyanic acid)	CN ⁻ (cyanide ion)		
		H ₂ O (water)	OH ⁻ (hydroxide ion)		
		NH ₃ (ammonia)	NH ₂ ⁻ (amide ion)		

Strong Acid (HCl)

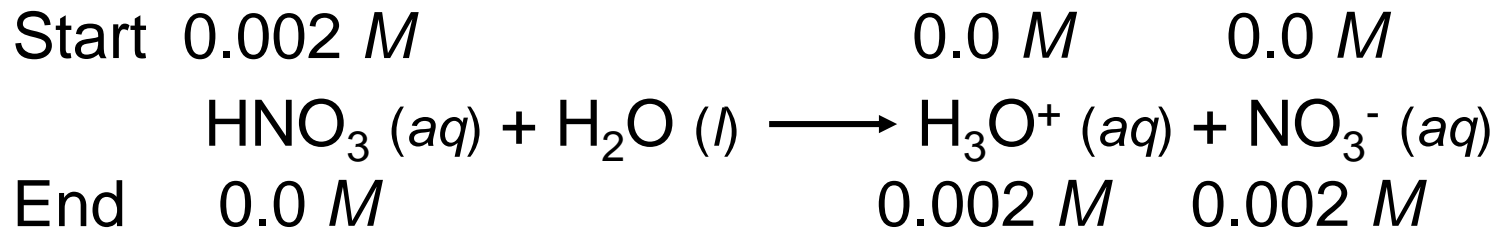


Weak Acid (HF)



What is the pH of a $2 \times 10^{-3} \text{ M}$ HNO_3 solution?

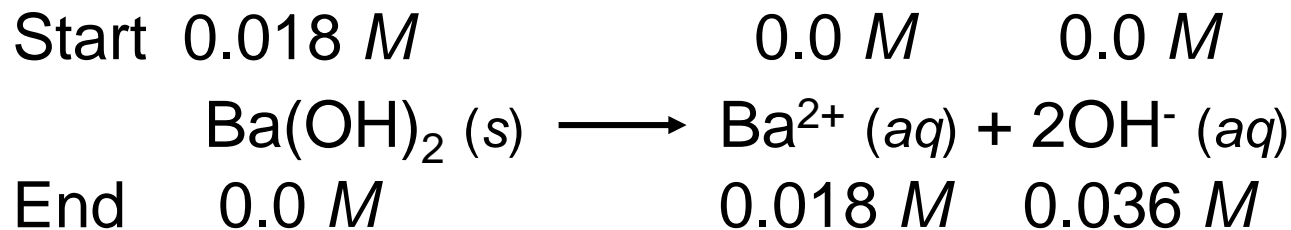
HNO_3 is a strong acid – 100% dissociation.



$$\text{pH} = -\log [\text{H}^+] = -\log [\text{H}_3\text{O}^+] = -\log(0.002) = 2.7$$

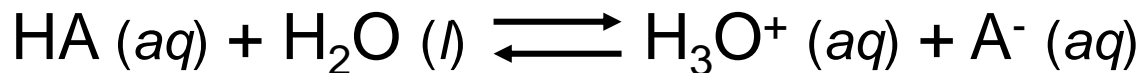
What is the pH of a $1.8 \times 10^{-2} \text{ M}$ $\text{Ba}(\text{OH})_2$ solution?

$\text{Ba}(\text{OH})_2$ is a strong base – 100% dissociation.



$$\text{pH} = 14.00 - \text{pOH} = 14.00 + \log(0.036) = 12.6$$

Weak Acids (HA) and Acid Ionization Constants



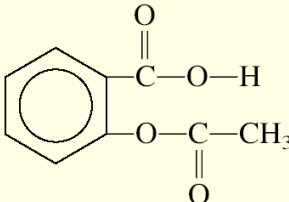
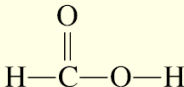
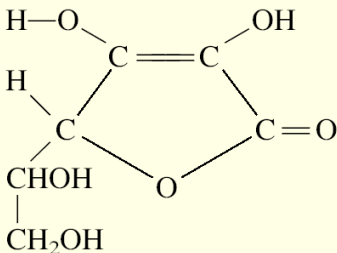
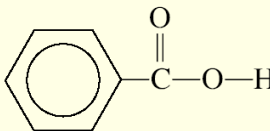
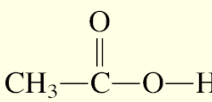
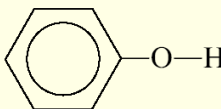
$$K_a = \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]}$$

K_a is the ***acid ionization constant***

$K_a \uparrow$

weak acid
strength \uparrow

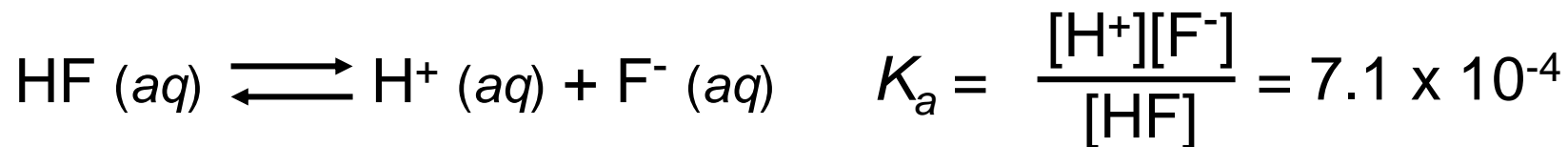
TABLE 15.3 Ionization Constants of Some Weak Acids and Their Conjugate Bases at 25°C

Name of Acid	Formula	Structure	K_a	Conjugate Base	K_b^\dagger
Hydrofluoric acid	HF	H—F	7.1×10^{-4}	F^-	1.4×10^{-11}
Nitrous acid	HNO_2	O=N—O—H	4.5×10^{-4}	NO_2^-	2.2×10^{-11}
Acetylsalicylic acid (aspirin)	$C_9H_8O_4$		3.0×10^{-4}	$C_9H_7O_4^-$	3.3×10^{-11}
Formic acid	HCOOH		1.7×10^{-4}	$HCOO^-$	5.9×10^{-11}
Ascorbic acid*	$C_6H_8O_6$		8.0×10^{-5}	$C_6H_7O_6^-$	1.3×10^{-10}
Benzoic acid	C_6H_5COOH		6.5×10^{-5}	$C_6H_5COO^-$	1.5×10^{-10}
Acetic acid	CH_3COOH		1.8×10^{-5}	CH_3COO^-	5.6×10^{-10}
Hydrocyanic acid	HCN	H—C≡N	4.9×10^{-10}	CN^-	2.0×10^{-5}
Phenol	C_6H_5OH		1.3×10^{-10}	$C_6H_5O^-$	7.7×10^{-5}

*For ascorbic acid it is the upper left hydroxyl group that is associated with this ionization constant.

[†]The base ionization constant K_b is discussed in Section 15.6.

What is the pH of a 0.5 M HF solution (at 25 °C)?



Initial (M)	0.50	0.00	0.00
-------------	------	------	------

Change (M)	-x	+x	+x
------------	----	----	----

Equilibrium (M)	0.50 - x	x	x
-----------------	----------	---	---

$$K_a = \frac{x^2}{0.50 - x} = 7.1 \times 10^{-4} \quad K_a \ll 1 \quad 0.50 - x \approx 0.50$$

$$K_a \approx \frac{x^2}{0.50} = 7.1 \times 10^{-4} \quad x^2 = 3.55 \times 10^{-4} \quad x = 0.019 \text{ M}$$

$$[\text{H}^+] = [\text{F}^-] = 0.019 \text{ M}$$

$$\text{pH} = -\log [\text{H}^+] = 1.72$$

$$[\text{HF}] = 0.50 - x = 0.48 \text{ M}$$

When can I use the approximation?

$$K_a \ll 1 \quad 0.50 - x \approx 0.50$$

When x is less than 5% of the value from which it is subtracted.

$$x = 0.019 \quad \frac{0.019 \text{ M}}{0.50 \text{ M}} \times 100\% = 3.8\% \quad \begin{array}{l} \text{Less than 5\%} \\ \text{Approximation ok.} \end{array}$$

What is the pH of a 0.05 M HF solution (at 25 °C)?

$$K_a \approx \frac{x^2}{0.05} = 7.1 \times 10^{-4} \quad x = 0.006 \text{ M}$$

$$\frac{0.006 \text{ M}}{0.05 \text{ M}} \times 100\% = 12\% \quad \begin{array}{l} \text{More than 5\%} \\ \text{Approximation **not** ok.} \end{array}$$

Must solve for x exactly using quadratic equation or method of successive approximations.

Solving **weak acid** ionization problems:

1. Identify the major species that can affect the pH.
 - In most cases, you can ignore the autoionization of water.
 - Ignore $[\text{OH}^-]$ because it is determined by $[\text{H}^+]$.
2. Use ICE to express the equilibrium concentrations in terms of single unknown x .
3. Write K_a in terms of equilibrium concentrations. Solve for x by the approximation method. If approximation is not valid, solve for x exactly.
4. Calculate concentrations of all species and/or pH of the solution.

What is the pH of a 0.122 *M* monoprotic acid whose K_a is 5.7×10^{-4} ?



Initial (<i>M</i>)	0.122	0.00	0.00
----------------------	-------	------	------

Change (<i>M</i>)	- <i>x</i>	+ <i>x</i>	+ <i>x</i>
---------------------	------------	------------	------------

Equilibrium (<i>M</i>)	0.122 - <i>x</i>	<i>x</i>	<i>x</i>
--------------------------	------------------	----------	----------

$$K_a = \frac{x^2}{0.122 - x} = 5.7 \times 10^{-4}$$

$$K_a \ll 1 \quad 0.122 - x \approx 0.122$$

$$K_a \approx \frac{x^2}{0.122} = 5.7 \times 10^{-4}$$

$$x^2 = 6.95 \times 10^{-5} \quad x = 0.0083 \text{ M}$$

$$\frac{0.0083 \text{ M}}{0.122 \text{ M}} \times 100\% = 6.8\%$$

More than 5%
Approximation **not** ok.

$$K_a = \frac{x^2}{0.122 - x} = 5.7 \times 10^{-4} \quad x^2 + 0.00057x - 6.95 \times 10^{-5} = 0$$

$$ax^2 + bx + c = 0$$

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

$$x = 0.0081$$

~~$$x = -0.0081$$~~



Initial (M)	0.122	0.00	0.00
-------------	-------	------	------

Change (M)	-x	+x	+x
------------	----	----	----

Equilibrium (M)	0.122 - x	x	x
-----------------	-----------	---	---

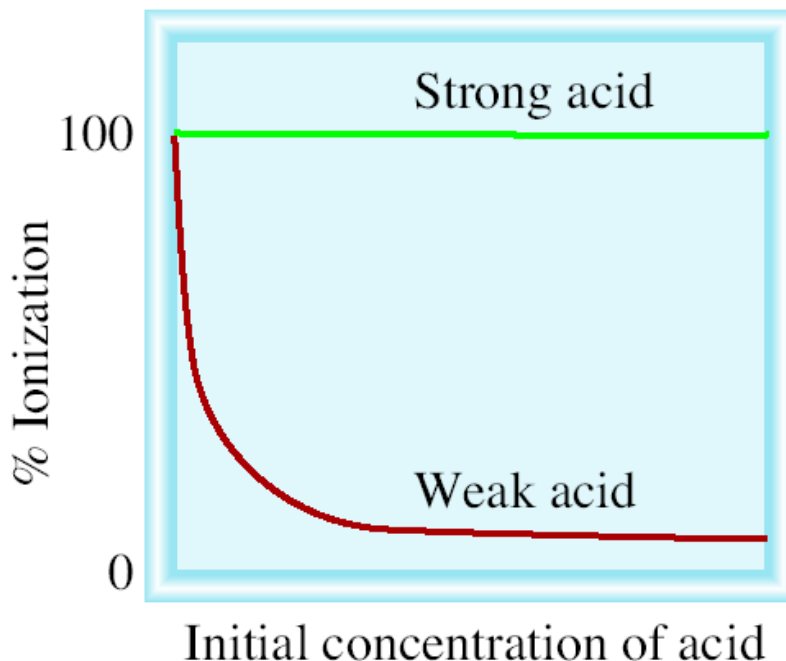
$$[\text{H}^+] = x = 0.0081 \text{ M}$$

$$\text{pH} = -\log[\text{H}^+] = 2.09$$

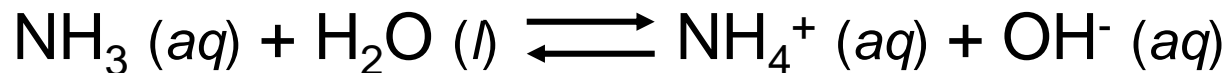
$$\textbf{percent ionization} = \frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$$

For a monoprotic acid HA

$$\text{Percent ionization} = \frac{[\text{H}^+]}{[\text{HA}]_0} \times 100\% \quad [\text{HA}]_0 = \text{initial concentration}$$



Weak Bases and Base Ionization Constants



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

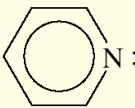
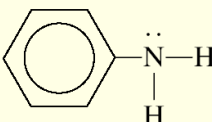
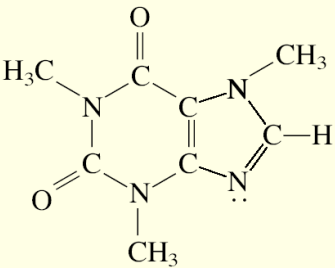
K_b is the ***base ionization constant***

K_b ↑

weak base
strength ↑

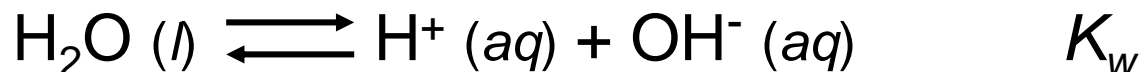
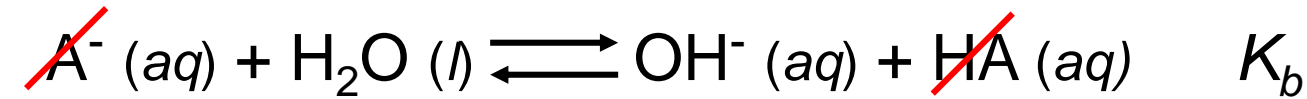
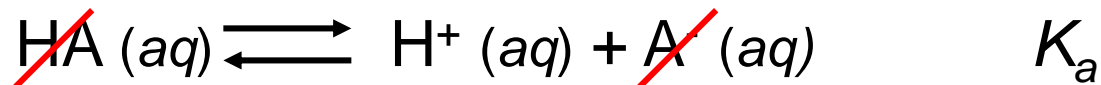
Solve weak base problems like weak acids ***except*** solve for $[\text{OH}^-]$ instead of $[\text{H}^+]$.

TABLE 15.4 Ionization Constants of Some Weak Bases and Their Conjugate Acids at 25°C

Name of Base	Formula	Structure	K_b^*	Conjugate Acid	K_a
Ethylamine	$C_2H_5NH_2$	$ \begin{array}{c} \text{CH}_3\text{---CH}_2\text{---}\ddot{\text{N}}\text{---H} \\ \\ \text{H} \end{array} $	5.6×10^{-4}	$C_2H_5NH_3^+$	1.8×10^{-11}
Methylamine	CH_3NH_2	$ \begin{array}{c} \text{CH}_3\text{---}\ddot{\text{N}}\text{---H} \\ \\ \text{H} \end{array} $	4.4×10^{-4}	$CH_3NH_3^+$	2.3×10^{-11}
Ammonia	NH_3	$ \begin{array}{c} \ddot{\text{N}}\text{---H} \\ \\ \text{H} \end{array} $	1.8×10^{-5}	NH_4^+	5.6×10^{-10}
Pyridine	C_5H_5N		1.7×10^{-9}	$C_5H_5NH^+$	5.9×10^{-6}
Aniline	$C_6H_5NH_2$		3.8×10^{-10}	$C_6H_5NH_3^+$	2.6×10^{-5}
Caffeine	$C_8H_{10}N_4O_2$		5.3×10^{-14}	$C_8H_{11}N_4O_2^+$	0.19
Urea	$(NH_2)_2CO$	$ \begin{array}{c} \text{O} \\ \\ \text{H---}\ddot{\text{N}}\text{---C---}\ddot{\text{N}}\text{---H} \\ \quad \quad \\ \text{H} \quad \quad \text{H} \end{array} $	1.5×10^{-14}	$H_2NCONH_3^+$	0.67

*The nitrogen atom with the lone pair accounts for each compound's basicity. In the case of urea, K_b can be associated with either nitrogen atom.

Ionization Constants of Conjugate Acid-Base Pairs



$$K_a K_b = K_w$$

Weak Acid and Its Conjugate Base

$$K_a = \frac{K_w}{K_b}$$

$$K_b = \frac{K_w}{K_a}$$

Diprotic and Triprotic Acids

- May yield more than one hydrogen ion per molecule.
- Ionize in a stepwise manner; that is, they lose one proton at a time.
- An ionization constant expression can be written for each ionization stage.
- Consequently, two or more equilibrium constant expressions must often be used to calculate the concentrations of species in the acid solution.

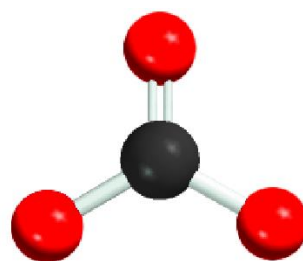
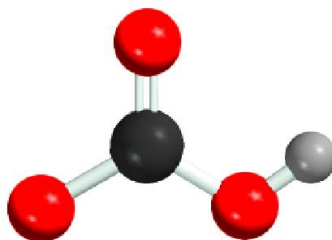
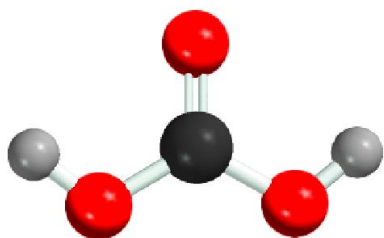
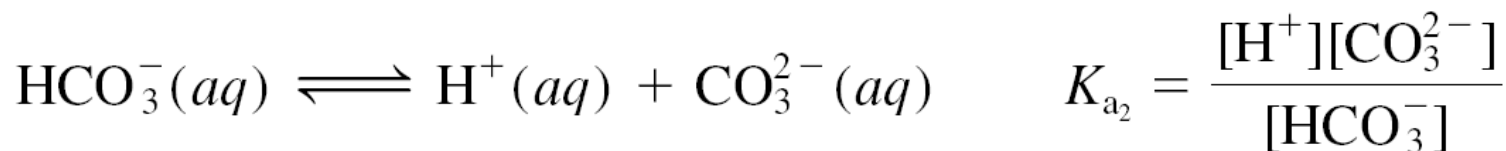
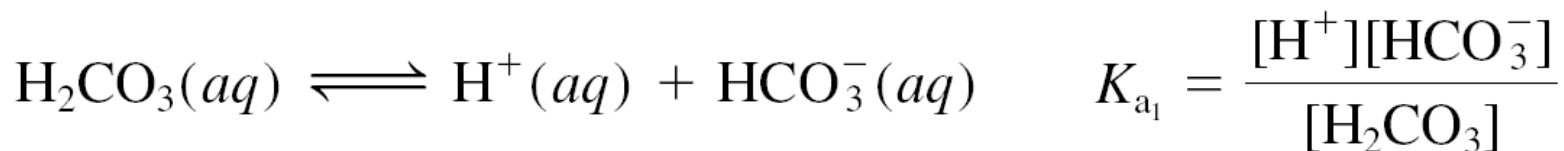


TABLE 15.5

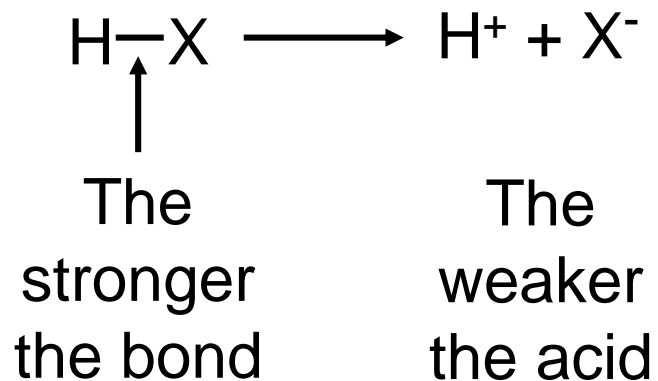
Ionization Constants of Some Diprotic Acids and a Polyprotic Acid and Their Conjugate Bases at 25°C

Name of Acid	Formula	Structure	K_a	Conjugate Base	K_b
Sulfuric acid	H_2SO_4	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{S}-\text{O}-\text{H} \\ \parallel \\ \text{O} \end{array}$	very large	HSO_4^-	very small
Hydrogen sulfate ion	HSO_4^-	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{S}-\text{O}^- \\ \parallel \\ \text{O} \end{array}$	1.3×10^{-2}	SO_4^{2-}	7.7×10^{-13}
Oxalic acid	$\text{H}_2\text{C}_2\text{O}_4$	$\begin{array}{c} \text{O} \quad \text{O} \\ \parallel \quad \parallel \\ \text{H}-\text{O}-\text{C}-\text{C}-\text{O}-\text{H} \end{array}$	6.5×10^{-2}	HC_2O_4^-	1.5×10^{-13}
Hydrogen oxalate ion	HC_2O_4^-	$\begin{array}{c} \text{O} \quad \text{O} \\ \parallel \quad \parallel \\ \text{H}-\text{O}-\text{C}-\text{C}-\text{O}^- \end{array}$	6.1×10^{-5}	$\text{C}_2\text{O}_4^{2-}$	1.6×10^{-10}
Sulfurous acid*	H_2SO_3	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{S}-\text{O}-\text{H} \end{array}$	1.3×10^{-2}	HSO_3^-	7.7×10^{-13}
Hydrogen sulfite ion	HSO_3^-	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{S}-\text{O}^- \end{array}$	6.3×10^{-8}	SO_3^{2-}	1.6×10^{-7}
Carbonic acid	H_2CO_3	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{C}-\text{O}-\text{H} \end{array}$	4.2×10^{-7}	HCO_3^-	2.4×10^{-8}
Hydrogen carbonate ion	HCO_3^-	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{C}-\text{O}^- \end{array}$	4.8×10^{-11}	CO_3^{2-}	2.1×10^{-4}
Hydrosulfuric acid	H_2S	$\text{H}-\text{S}-\text{H}$	9.5×10^{-8}	HS^-	1.1×10^{-7}
Hydrogen sulfide ion [†]	HS^-	$\text{H}-\text{S}^-$	1×10^{-19}	S^{2-}	1×10^5
Phosphoric acid	H_3PO_4	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{P}-\text{O}-\text{H} \\ \\ \text{O} \\ \\ \text{H} \end{array}$	7.5×10^{-3}	H_2PO_4^-	1.3×10^{-12}
Dihydrogen phosphate ion	H_2PO_4^-	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{P}-\text{O}^- \\ \\ \text{O} \\ \\ \text{H} \end{array}$	6.2×10^{-8}	HPO_4^{2-}	1.6×10^{-7}
Hydrogen phosphate ion	HPO_4^{2-}	$\begin{array}{c} \text{O} \\ \parallel \\ \text{H}-\text{O}-\text{P}-\text{O}^- \\ \\ \text{O}^- \end{array}$	4.8×10^{-13}	PO_4^{3-}	2.1×10^{-2}

* H_2SO_3 has never been isolated and exists in only minute concentration in aqueous solution of SO_2 . The K_a value here refers to the process $\text{SO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}^+(\text{aq}) + \text{HSO}_3^-(\text{aq})$.

[†]The ionization constant of HS^- is very low and difficult to measure. The value listed here is only an estimate.

Molecular Structure and Acid Strength

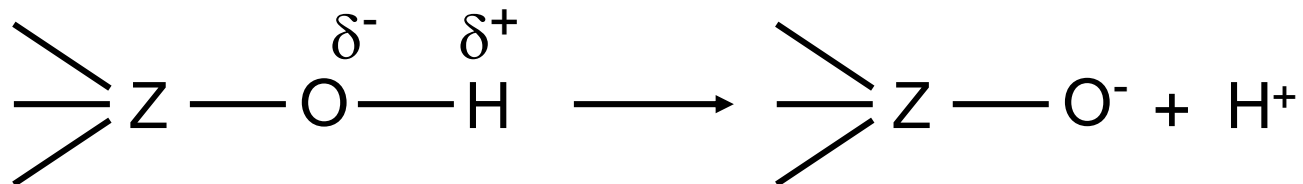


The diagram shows a simplified periodic table with columns labeled 1A, 2A, 3A, 4A, 5A, 6A, 7A, and 8A. The elements Fluorine (F), Chlorine (Cl), Bromine (Br), and Iodine (I) are highlighted in the 7A column. A red arrow points downwards from F to I, with the text "acidity increases" next to it.

TABLE 15.6**Bond Enthalpies for Hydrogen Halides and Acid Strengths for Hydrohalic Acids**

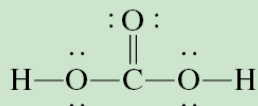
Bond	Bond Enthalpy (kJ/mol)	Acid Strength
H—F	568.2	weak
H—Cl	431.9	strong
H—Br	366.1	strong
H—I	298.3	strong

Molecular Structure and Oxoacid Strength

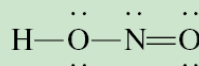


The O-H bond will be more polar and easier to break if:

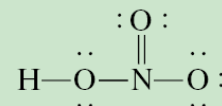
- Z is very electronegative or
- Z is in a high oxidation state



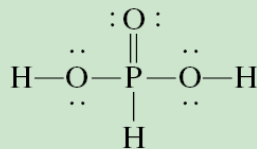
Carbonic acid



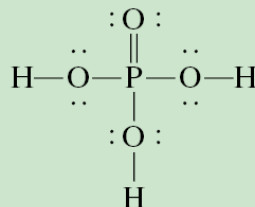
Nitrous acid



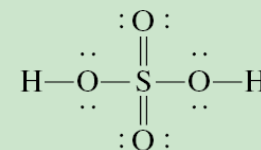
Nitric acid



Phosphorous acid



Phosphoric acid

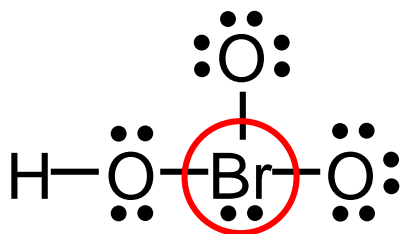
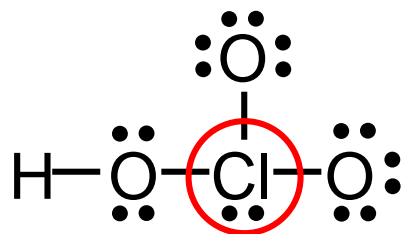


Sulfuric acid

Molecular Structure and Oxoacid Strength

1. Oxoacids having different central atoms (Z) that **are from the same group** and that have the **same oxidation number**.

Acid strength increases with increasing electronegativity of Z



Cl is more electronegative than Br

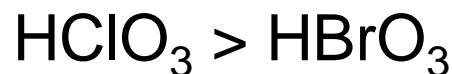
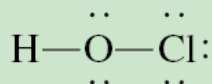


Diagram illustrating the trend of increasing acidity across the periodic table. The diagram shows a simplified periodic table with groups labeled 1A, 2A, 3A, 4A, 5A, 6A, 7A, and 8A. The elements Cl, Br, and I are highlighted in the 7A group. A red arrow points upwards from the bottom right towards the top right, with the text "acidity increases" next to it.

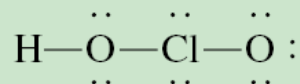
Molecular Structure and Acid Strength

2. Oxoacids having the same central atom (Z) but different numbers of attached groups.

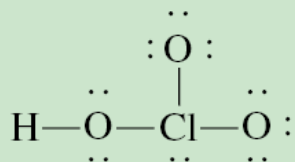
Acid strength increases as the oxidation number of Z increases.



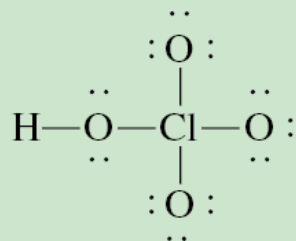
Hypochlorous acid (+1)



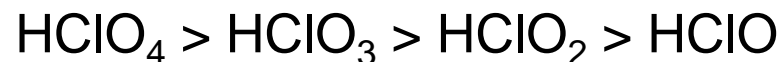
Chlorous acid (+3)



Chloric acid (+5)



Perchloric acid (+7)



Acid-Base Properties of Salts

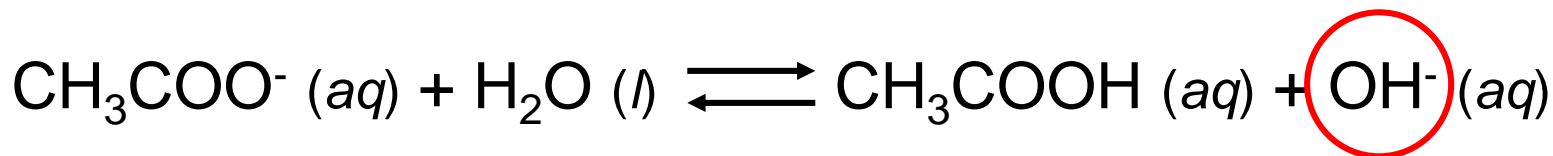
Neutral Solutions:

Salts containing an alkali metal or alkaline earth metal ion (except Be^{2+}) **and** the conjugate base of a **strong** acid (e.g. Cl^- , Br^- , and NO_3^-).



Basic Solutions:

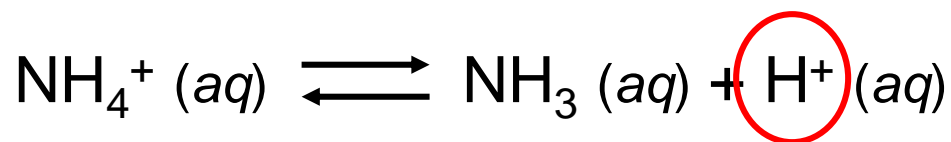
Salts derived from a strong base **and** a **weak** acid.



Acid-Base Properties of Salts

Acid Solutions:

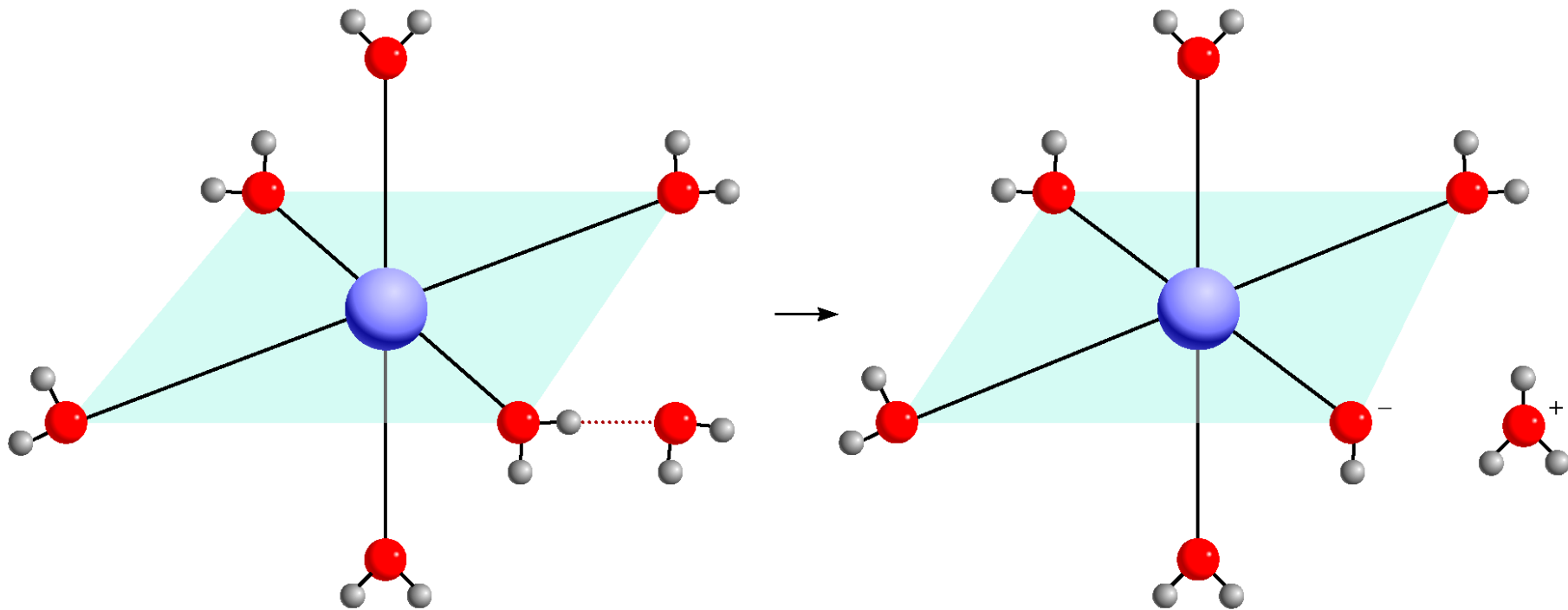
Salts derived from a strong acid and a weak base.



Salts with small, highly charged metal cations (e.g. Al^{3+} , Cr^{3+} , and Be^{2+}) and the conjugate base of a strong acid.



Acid Hydrolysis of Al^{3+}



Acid-Base Properties of Salts

Solutions in which both the cation and the anion hydrolyze:

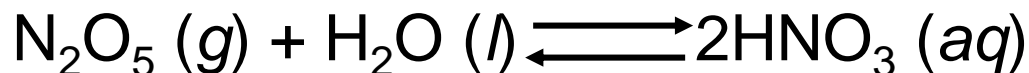
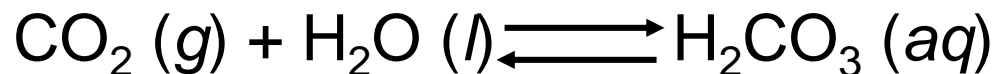
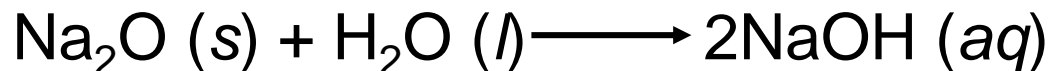
- K_b for the anion $> K_a$ for the cation, solution will be basic
- K_b for the anion $< K_a$ for the cation, solution will be acidic
- K_b for the anion $\approx K_a$ for the cation, solution will be neutral

TABLE 15.7 Acid-Base Properties of Salts

Type of Salt	Examples	Ions That Undergo Hydrolysis	pH of Solution
Cation from strong base; anion from strong acid	NaCl, KI, KNO ₃ , RbBr, BaCl ₂	None	≈ 7
Cation from strong base; anion from weak acid	CH ₃ COONa, KNO ₂	Anion	> 7
Cation from weak base; anion from strong acid	NH ₄ Cl, NH ₄ NO ₃	Cation	< 7
Cation from weak base; anion from weak acid	NH ₄ NO ₂ , CH ₃ COONH ₄ , NH ₄ CN	Anion and cation	< 7 if $K_b < K_a$ ≈ 7 if $K_b \approx K_a$ > 7 if $K_b > K_a$
Small, highly charged cation; anion from strong acid	AlCl ₃ , Fe(NO ₃) ₃	Hydrated cation	< 7

Oxides of the Representative Elements In Their Highest Oxidation States

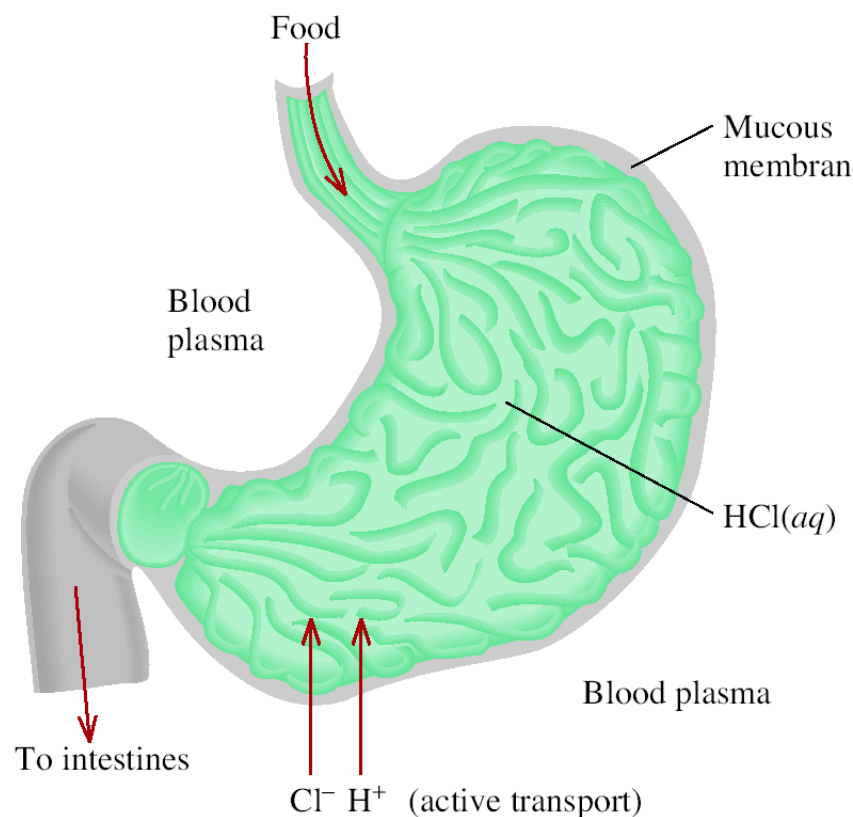
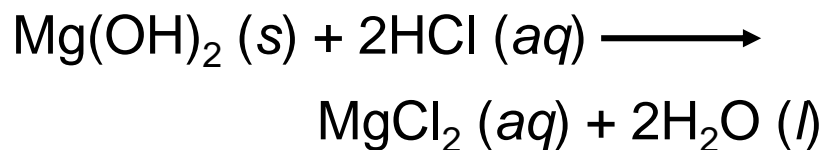
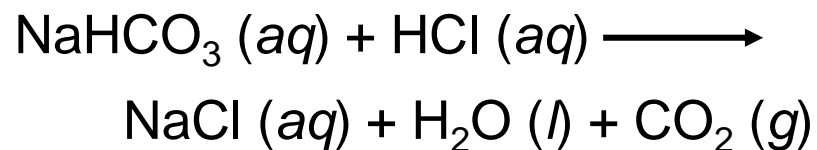
1 1A	2 2A											13 3A	14 4A	15 5A	16 6A	17 7A	18 8A
Li ₂ O	BeO											B ₂ O ₃	CO ₂	N ₂ O ₅		OF ₂	
Na ₂ O	MgO	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B	Al ₂ O ₃	SiO ₂	P ₄ O ₁₀	SO ₃	Cl ₂ O ₇	
K ₂ O	CaO											Ga ₂ O ₃	GeO ₂	As ₂ O ₅	SeO ₃	Br ₂ O ₇	
Rb ₂ O	SrO											In ₂ O ₃	SnO ₂	Sb ₂ O ₅	TeO ₃	I ₂ O ₇	
Cs ₂ O	BaO											Tl ₂ O ₃	PbO ₂	Bi ₂ O ₅	PoO ₃	At ₂ O ₇	



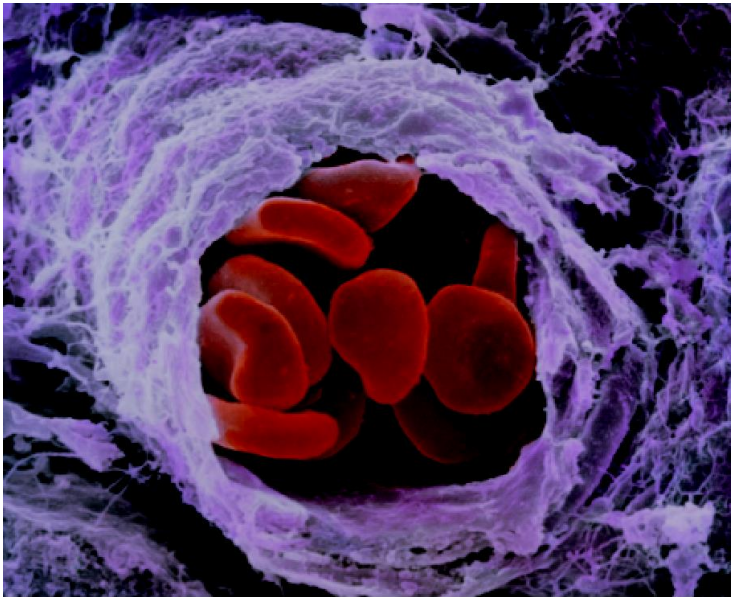
Chemistry In Action: Antacids and the Stomach pH Balance

Some Common Commercial Antacid Preparations

Commercial Name	Active Ingredients
Alka-2	Calcium carbonate
Alka-Seltzer	Aspirin, sodium bicarbonate, citric acid
Bufferin	Aspirin, magnesium carbonate, aluminum glycinate
Buffered aspirin	Aspirin, magnesium carbonate, aluminum hydroxide-glycine
Milk of magnesia	Magnesium hydroxide
Rolaids	Dihydroxy aluminum sodium carbonate
Tums	Calcium carbonate



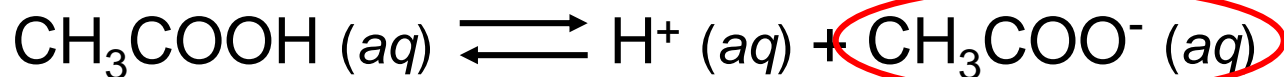
Acid-Base Equilibria and Solubility Equilibria



The ***common ion effect*** is the shift in equilibrium caused by the addition of a compound having an ion in common with the dissolved substance.

The presence of a common ion **suppresses** the ionization of a weak acid or a weak base.

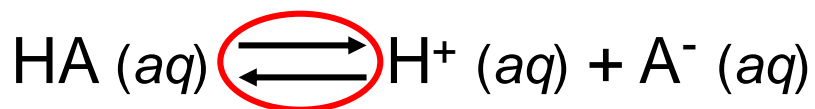
Consider mixture of CH_3COONa (strong electrolyte) and CH_3COOH (weak acid).



common
ion



Consider mixture of salt NaA and weak acid HA.



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

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

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

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

$$\text{pH} = \text{p}K_a + \log \frac{[\text{A}^-]}{[\text{HA}]}$$

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

Henderson-Hasselbalch
equation

$$\text{pH} = \text{p}K_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$$

What is the pH of a solution containing 0.30 M HCOOH and 0.52 M HCOOK?

Mixture of weak acid and conjugate base!



Initial (M)	0.30	0.00	0.52
Change (M)	-x	+x	+x
Equilibrium (M)	0.30 - x	x	0.52 + x

Common ion effect

$$0.30 - x \approx 0.30$$

$$0.52 + x \approx 0.52$$

$$\text{HCOOH } pK_a = 3.77$$

$$\text{pH} = pK_a + \log \frac{[\text{HCOO}^-]}{[\text{HCOOH}]}$$

$$\text{pH} = 3.77 + \log \frac{[0.52]}{[0.30]} = 4.01$$

A **buffer solution** is a solution of:

1. A weak acid or a weak base **and**
2. The salt of the weak acid or weak base

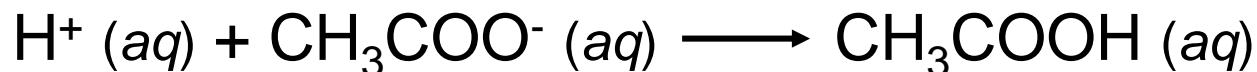
Both must be present!



A buffer solution has the ability to resist changes in pH upon the addition of small amounts of either acid or base.

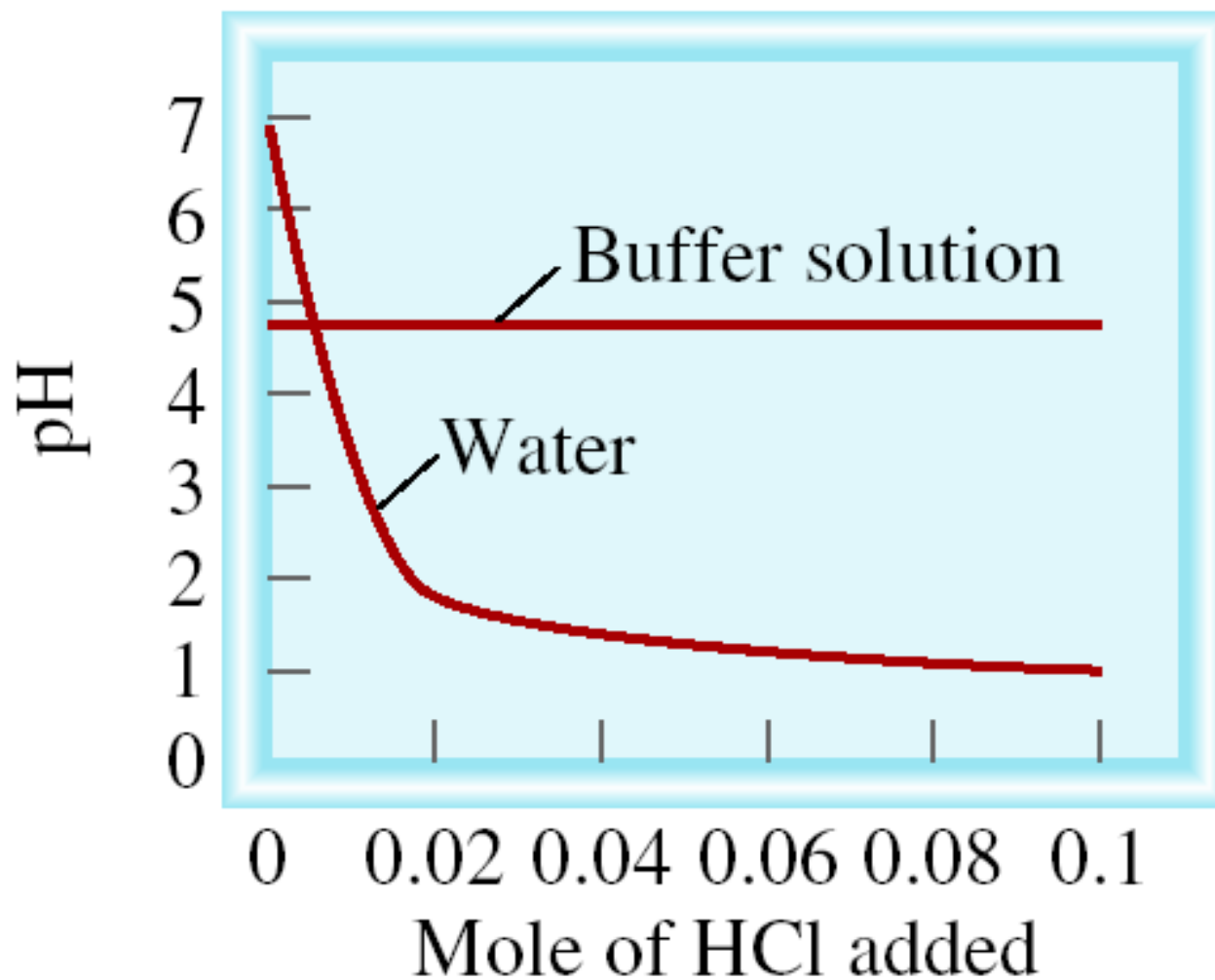
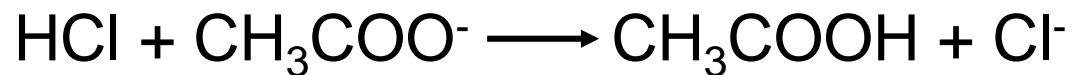
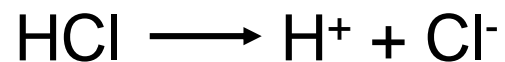
Consider an equal molar mixture of CH_3COOH and CH_3COONa

Add strong acid



Add strong base





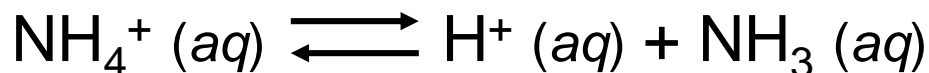
Which of the following are buffer systems? (a) KF/HF
(b) KBr/HBr, (c) $\text{Na}_2\text{CO}_3/\text{NaHCO}_3$

(a) KF is a **weak** acid and F^- is its **conjugate** base
buffer solution

(b) HBr is a **strong** acid
not a buffer solution

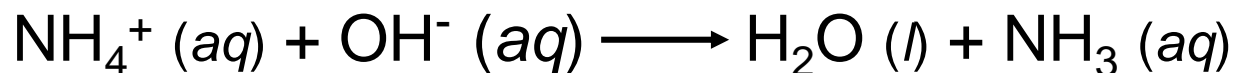
(c) CO_3^{2-} is a **weak** base and HCO_3^- is its **conjugate** acid
buffer solution

Calculate the pH of the 0.30 M NH₃/0.36 M NH₄Cl buffer system. What is the pH after the addition of 20.0 mL of 0.050 M NaOH to 80.0 mL of the buffer solution?



$$\text{pH} = \text{p}K_a + \log \frac{[\text{NH}_3]}{[\text{NH}_4^+]} \quad \text{p}K_a = 9.25 \quad \text{pH} = 9.25 + \log \frac{[0.30]}{[0.36]} = 9.17$$

start (moles)	0.029	0.001	0.024
---------------	-------	-------	-------

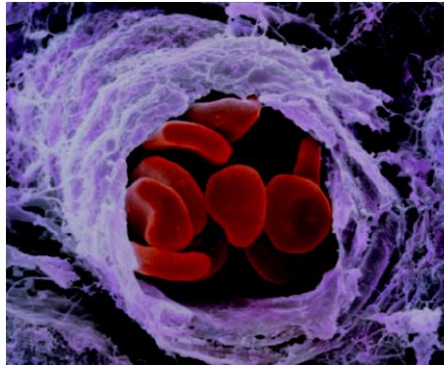


end (moles)	0.028	0.0	0.025
-------------	-------	-----	-------

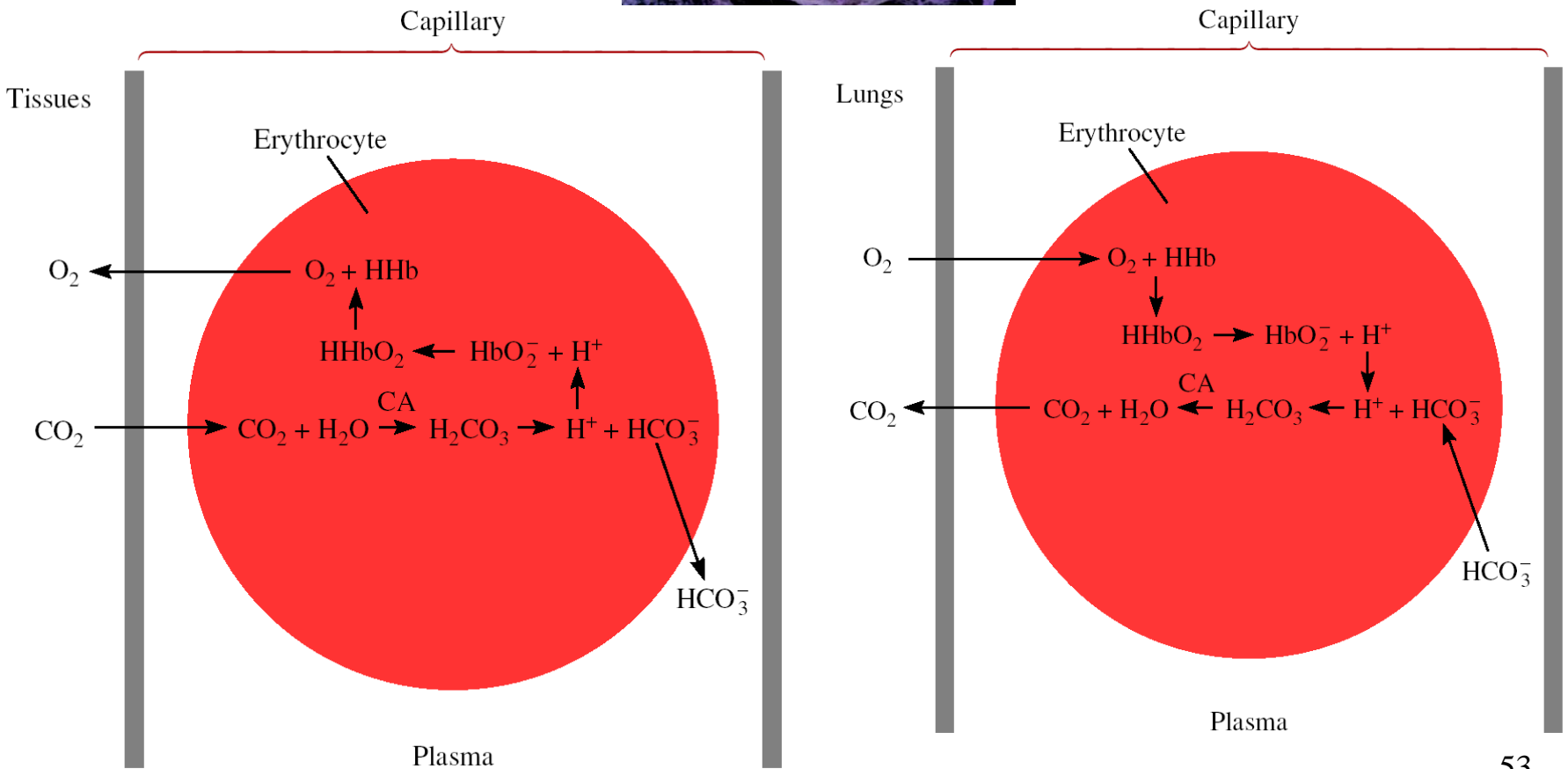
final volume = 80.0 mL + 20.0 mL = 100 mL

$$[\text{NH}_4^+] = \frac{0.028}{0.10} \quad [\text{NH}_3] = \frac{0.025}{0.10} \quad \text{pH} = 9.25 + \log \frac{[0.25]}{[0.28]} = 9.20$$

Chemistry In Action: Maintaining the pH of Blood



Red blood cells in a capillary

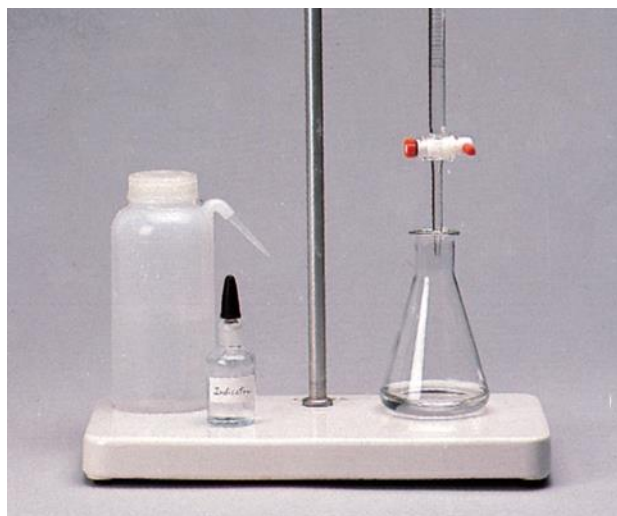


Titration (Review)

In a ***titration*** a solution of accurately known concentration is added gradually to another solution of unknown concentration until the chemical reaction between the two solutions is complete.

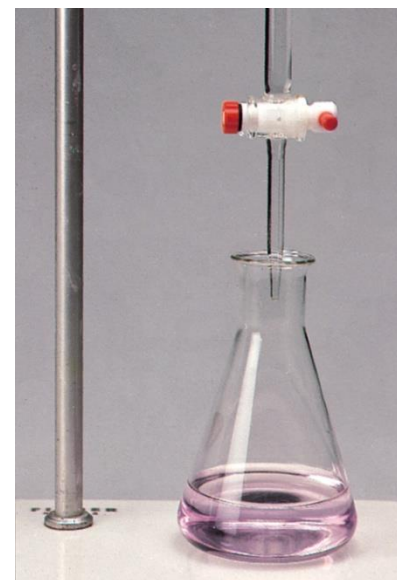
Equivalence point – the point at which the reaction is complete

Indicator – substance that changes color at (or near) the equivalence point

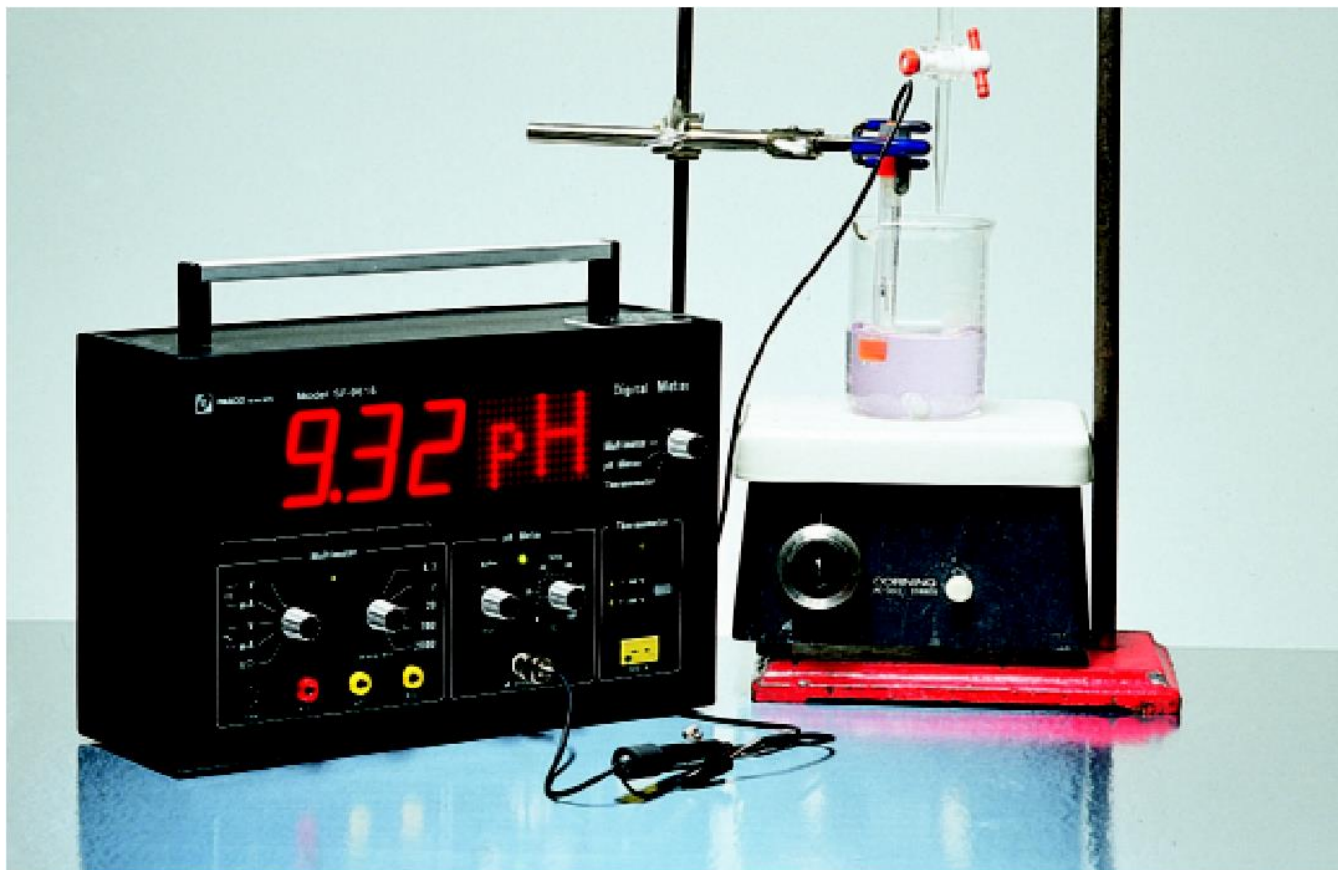


Slowly add base
to unknown acid
UNTIL

The indicator
changes color
(pink)

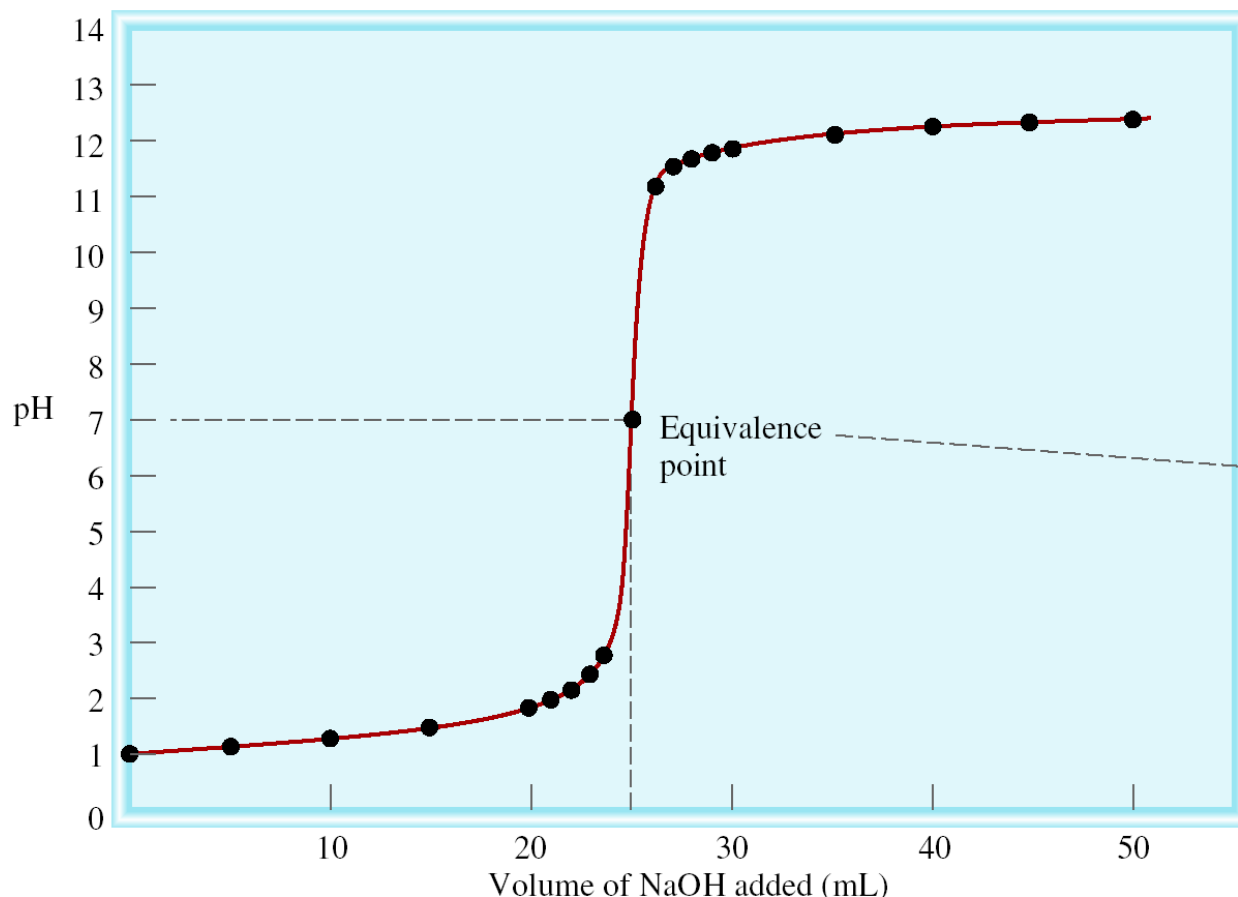
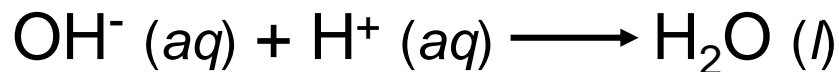


Alternative Method of Equivalence Point Detection



monitor pH

Strong Acid-Strong Base Titrations

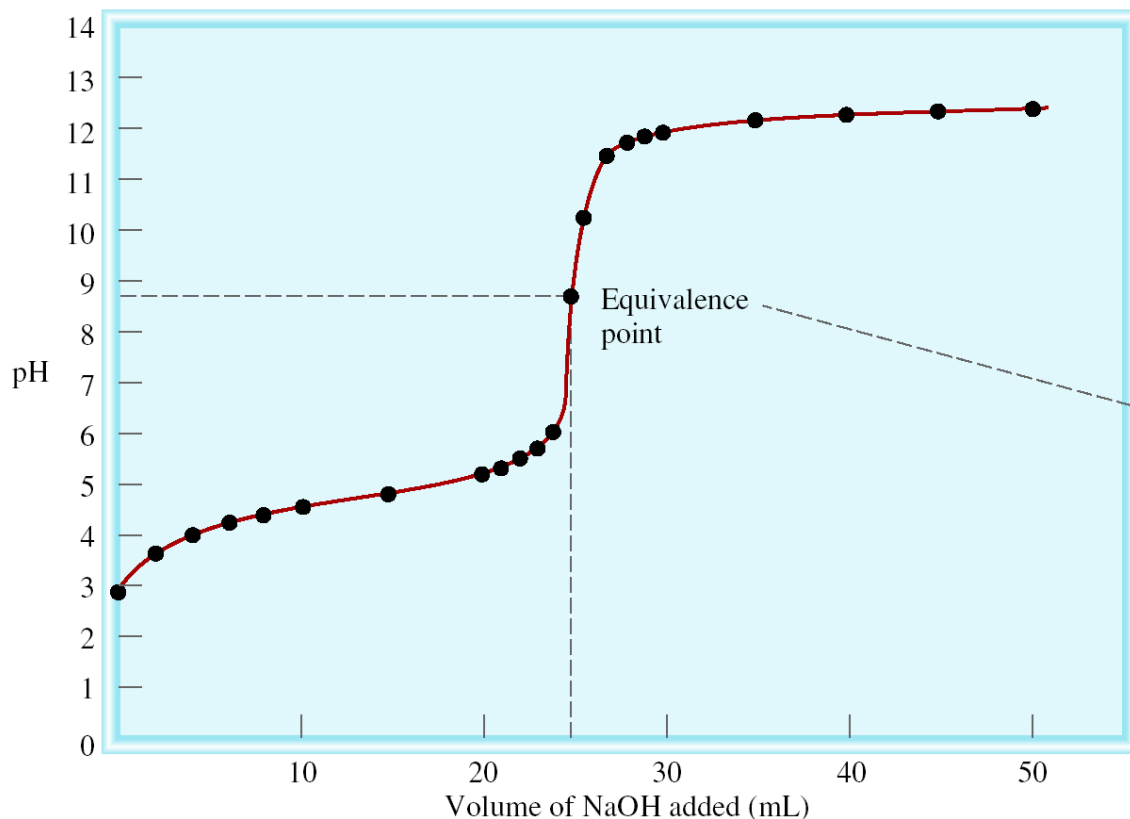
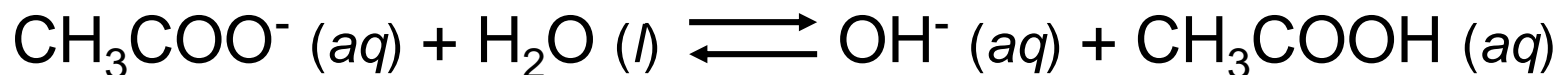


Volume NaOH added (mL)	pH
0.0	1.00
5.0	1.18
10.0	1.37
15.0	1.60
20.0	1.95
22.0	2.20
24.0	2.69
25.0	7.00
26.0	11.29
28.0	11.75
30.0	11.96
35.0	12.22
40.0	12.36
45.0	12.46
50.0	12.52

Weak Acid-Strong Base Titrations

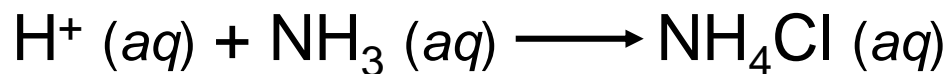
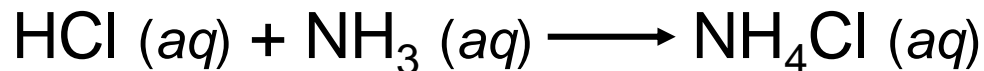


At equivalence point ($\text{pH} > 7$):

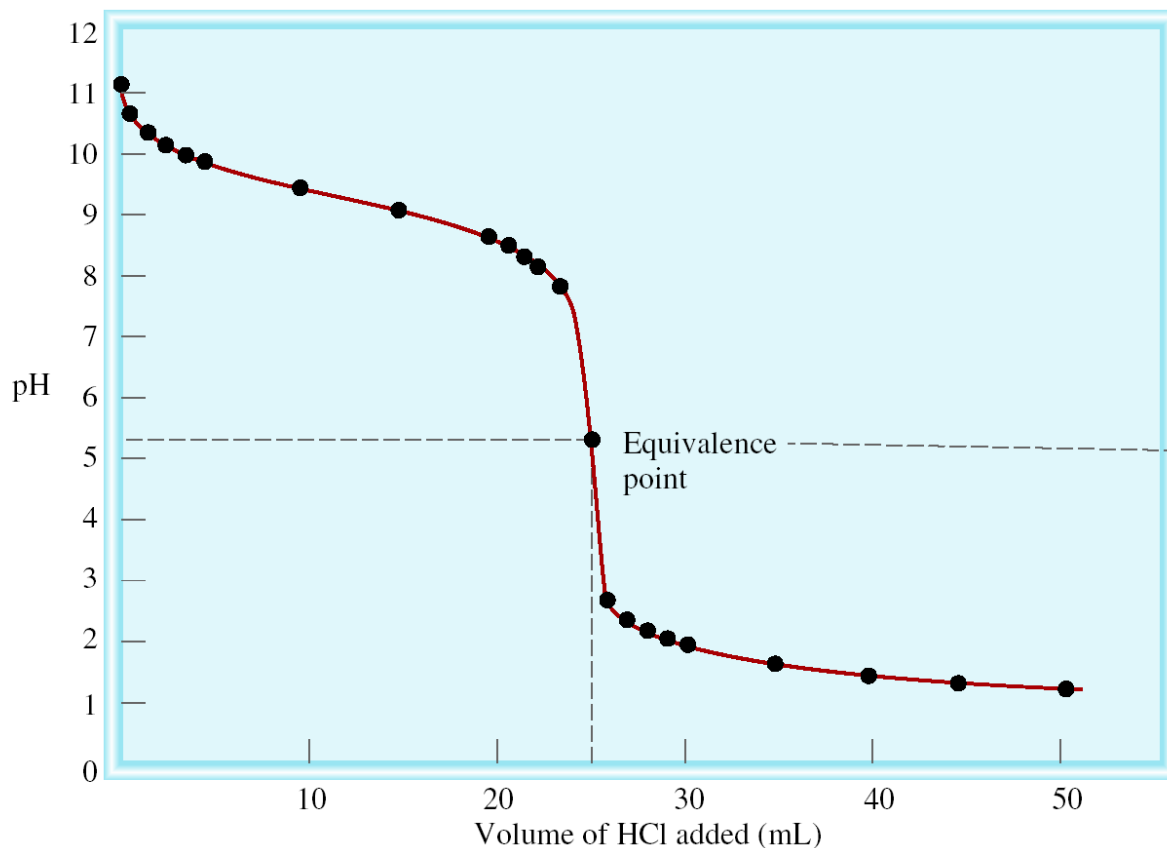


Volume NaOH added (mL)	pH
0.0	2.87
5.0	4.14
10.0	4.57
15.0	4.92
20.0	5.35
22.0	5.61
24.0	6.12
25.0	8.72
26.0	10.29
28.0	11.75
30.0	11.96
35.0	12.22
40.0	12.36
45.0	12.46
50.0	12.52

Strong Acid-Weak Base Titrations

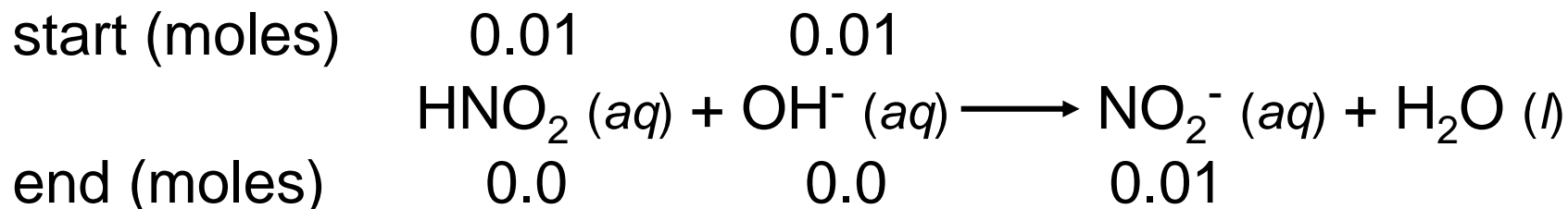


At equivalence point ($\text{pH} < 7$):



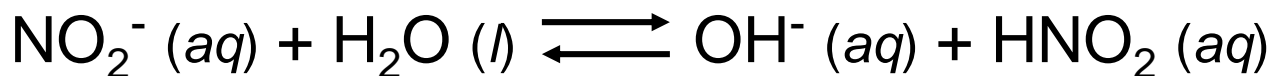
Volume HCl added (mL)	pH
0.0	11.13
5.0	9.86
10.0	9.44
15.0	9.08
20.0	8.66
22.0	8.39
24.0	7.88
25.0	5.28
26.0	2.70
28.0	2.22
30.0	2.00
35.0	1.70
40.0	1.52
45.0	1.40
50.0	1.30

Exactly 100 mL of 0.10 *M* HNO₂ are titrated with a 0.10 *M* NaOH solution. What is the pH at the equivalence point ?



Final volume = 200 mL

$$[\text{NO}_2^-] = \frac{0.01}{0.200} = 0.05 \text{ M}$$



Initial (<i>M</i>)	0.05	0.00	0.00
----------------------	------	------	------

Change (<i>M</i>)	- <i>x</i>	+ <i>x</i>	+ <i>x</i>
---------------------	------------	------------	------------

Equilibrium (<i>M</i>)	0.05 - <i>x</i>	<i>x</i>	<i>x</i>
--------------------------	-----------------	----------	----------

$$K_b = \frac{[\text{OH}^-][\text{HNO}_2]}{[\text{NO}_2^-]} = \frac{x^2}{0.05-x} = 2.2 \times 10^{-11} \quad \text{pOH} = 5.98$$

0.05 - *x* ≈ 0.05 $x \approx 1.05 \times 10^{-6} = [\text{OH}^-]$ pH = 14 - pOH = 8.02

Acid-Base Indicators



$\frac{[\text{HIn}]}{[\text{In}^-]} \geq 10$ Color of acid (HIn) predominates

$\frac{[\text{HIn}]}{[\text{In}^-]} \leq 10$ Color of conjugate base (In^-) predominates

TABLE 16.1 Some Common Acid-Base Indicators

Indicator	Color		pH Range*
	In Acid	In Base	
Thymol blue	Red	Yellow	1.2–2.8
Bromophenol blue	Yellow	Bluish purple	3.0–4.6
Methyl orange	Orange	Yellow	3.1–4.4
Methyl red	Red	Yellow	4.2–6.3
Chlorophenol blue	Yellow	Red	4.8–6.4
Bromothymol blue	Yellow	Blue	6.0–7.6
Cresol red	Yellow	Red	7.2–8.8
Phenolphthalein	Colorless	Reddish pink	8.3–10.0

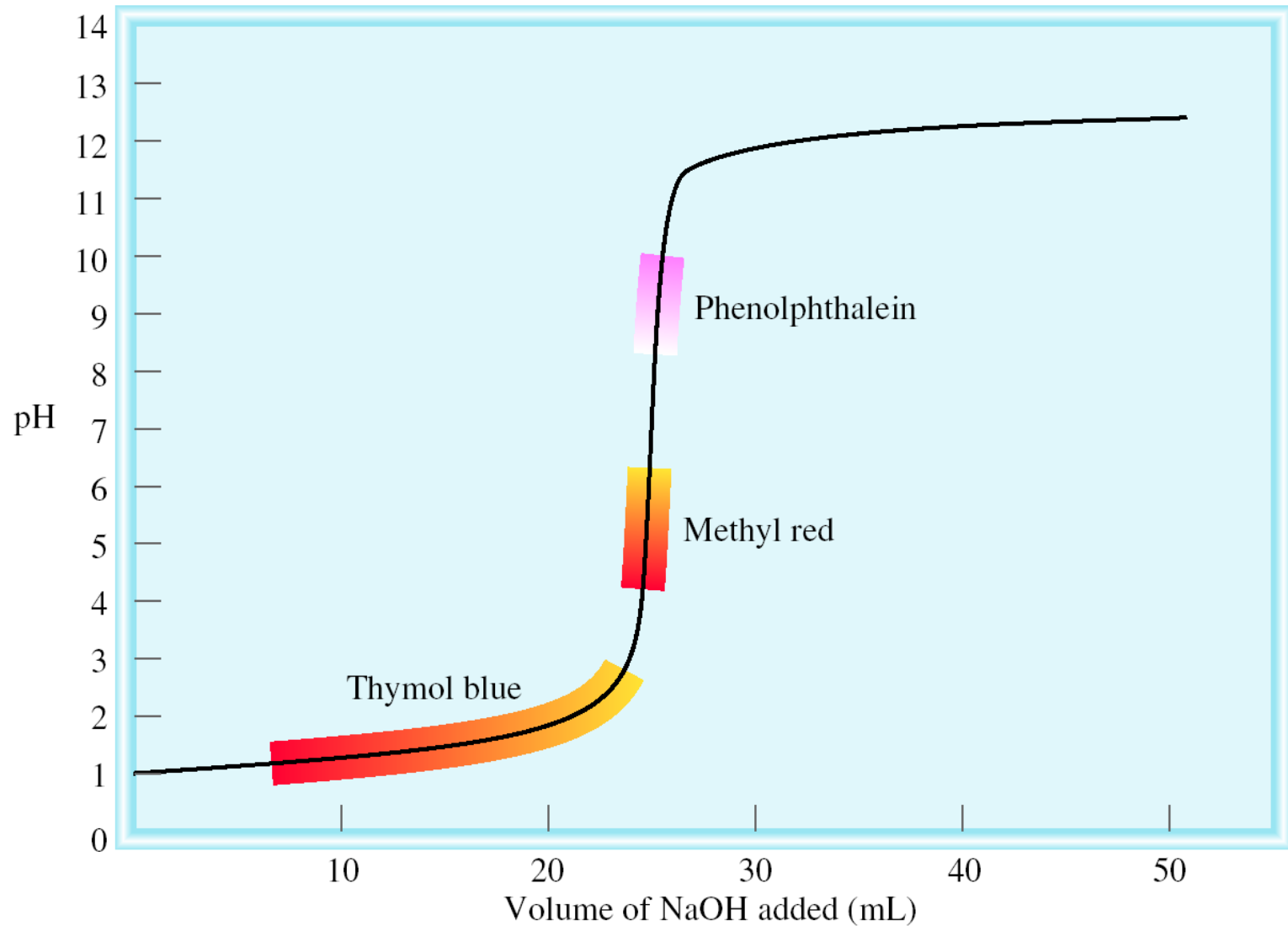
*The pH range is defined as the range over which the indicator changes from the acid color to the base color.

Solutions of Red Cabbage Extract



———— pH —————→

The titration curve of a strong acid with a strong base.



Which indicator(s) would you use for a titration of HNO_2 with KOH ?

Weak acid titrated with strong base.

At equivalence point, will have conjugate base of weak acid.

At equivalence point, $\text{pH} > 7$

Use cresol red or phenolphthalein

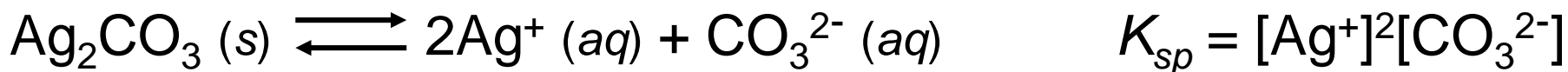
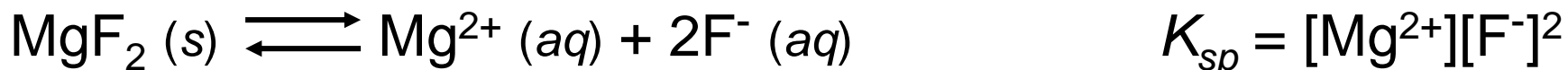
TABLE 16.1 Some Common Acid-Base Indicators			
Indicator	Color		pH Range*
	In Acid	In Base	
Thymol blue	Red	Yellow	1.2–2.8
Bromophenol blue	Yellow	Bluish purple	3.0–4.6
Methyl orange	Orange	Yellow	3.1–4.4
Methyl red	Red	Yellow	4.2–6.3
Chlorophenol blue	Yellow	Red	4.8–6.4
Bromothymol blue	Yellow	Blue	6.0–7.6
Cresol red	Yellow	Red	7.2–8.8
Phenolphthalein	Colorless	Reddish pink	8.3–10.0

*The pH range is defined as the range over which the indicator changes from the acid color to the base color.

Solubility Equilibria



$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$ K_{sp} is the ***solubility product constant***



Dissolution of an ionic solid in aqueous solution:

$Q < K_{sp}$ Unsaturated solution No precipitate

$Q = K_{sp}$ Saturated solution

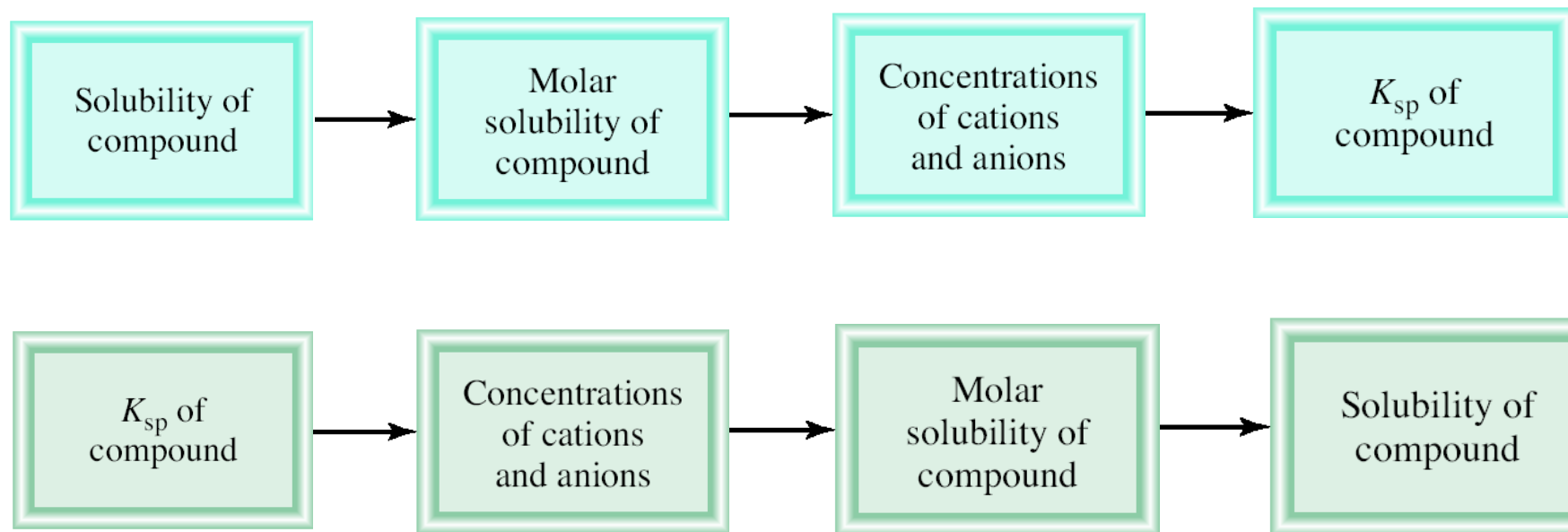
$Q > K_{sp}$ Supersaturated solution Precipitate will form

TABLE 16.2 Solubility Products of Some Slightly Soluble Ionic Compounds at 25°C

Compound	K_{sp}	Compound	K_{sp}
Aluminum hydroxide [Al(OH) ₃]	1.8×10^{-33}	Lead(II) chromate (PbCrO ₄)	2.0×10^{-14}
Barium carbonate (BaCO ₃)	8.1×10^{-9}	Lead(II) fluoride (PbF ₂)	4.1×10^{-8}
Barium fluoride (BaF ₂)	1.7×10^{-6}	Lead(II) iodide (PbI ₂)	1.4×10^{-8}
Barium sulfate (BaSO ₄)	1.1×10^{-10}	Lead(II) sulfide (PbS)	3.4×10^{-28}
Bismuth sulfide (Bi ₂ S ₃)	1.6×10^{-72}	Magnesium carbonate (MgCO ₃)	4.0×10^{-5}
Cadmium sulfide (CdS)	8.0×10^{-28}	Magnesium hydroxide [Mg(OH) ₂]	1.2×10^{-11}
Calcium carbonate (CaCO ₃)	8.7×10^{-9}	Manganese(II) sulfide (MnS)	3.0×10^{-14}
Calcium fluoride (CaF ₂)	4.0×10^{-11}	Mercury(I) chloride (Hg ₂ Cl ₂)	3.5×10^{-18}
Calcium hydroxide [Ca(OH) ₂]	8.0×10^{-6}	Mercury(II) sulfide (HgS)	4.0×10^{-54}
Calcium phosphate [Ca ₃ (PO ₄) ₂]	1.2×10^{-26}	Nickel(II) sulfide (NiS)	1.4×10^{-24}
Chromium(III) hydroxide [Cr(OH) ₃]	3.0×10^{-29}	Silver bromide (AgBr)	7.7×10^{-13}
Cobalt(II) sulfide (CoS)	4.0×10^{-21}	Silver carbonate (Ag ₂ CO ₃)	8.1×10^{-12}
Copper(I) bromide (CuBr)	4.2×10^{-8}	Silver chloride (AgCl)	1.6×10^{-10}
Copper(I) iodide (CuI)	5.1×10^{-12}	Silver iodide (AgI)	8.3×10^{-17}
Copper(II) hydroxide [Cu(OH) ₂]	2.2×10^{-20}	Silver sulfate (Ag ₂ SO ₄)	1.4×10^{-5}
Copper(II) sulfide (CuS)	6.0×10^{-37}	Silver sulfide (Ag ₂ S)	6.0×10^{-51}
Iron(II) hydroxide [Fe(OH) ₂]	1.6×10^{-14}	Strontium carbonate (SrCO ₃)	1.6×10^{-9}
Iron(III) hydroxide [Fe(OH) ₃]	1.1×10^{-36}	Strontium sulfate (SrSO ₄)	3.8×10^{-7}
Iron(II) sulfide (FeS)	6.0×10^{-19}	Tin(II) sulfide (SnS)	1.0×10^{-26}
Lead(II) carbonate (PbCO ₃)	3.3×10^{-14}	Zinc hydroxide [Zn(OH) ₂]	1.8×10^{-14}
Lead(II) chloride (PbCl ₂)	2.4×10^{-4}	Zinc sulfide (ZnS)	3.0×10^{-23}

Molar solubility (mol/L) is the number of moles of solute dissolved in 1 L of a saturated solution.

Solubility (g/L) is the number of grams of solute dissolved in 1 L of a saturated solution.



What is the solubility of silver chloride in g/L ?



Initial (<i>M</i>)	0.00	0.00	$K_{sp} = [\text{Ag}^+][\text{Cl}^-]$
----------------------	------	------	---------------------------------------

Change (<i>M</i>)	+ <i>s</i>	+ <i>s</i>	$K_{sp} = s^2$
---------------------	------------	------------	----------------

Equilibrium (<i>M</i>)	<i>s</i>	<i>s</i>	$s = \sqrt{K_{sp}}$
--------------------------	----------	----------	---------------------

$$s = 1.3 \times 10^{-5}$$

$$[\text{Ag}^+] = 1.3 \times 10^{-5} \text{ M} \quad [\text{Cl}^-] = 1.3 \times 10^{-5} \text{ M}$$

$$\text{Solubility of AgCl} = \frac{1.3 \times 10^{-5} \cancel{\text{mol AgCl}}}{1 \text{ L soln}} \times \frac{143.35 \text{ g AgCl}}{1 \cancel{\text{mol AgCl}}} = 1.9 \times 10^{-3} \text{ g/L}$$

TABLE 16.3 Relationship Between K_{sp} and Molar Solubility (s)

Compound	K_{sp} Expression	Cation	Anion	Relation Between K_{sp} and s
AgCl	$[\text{Ag}^+][\text{Cl}^-]$	s	s	$K_{\text{sp}} = s^2; s = (K_{\text{sp}})^{\frac{1}{2}}$
BaSO ₄	$[\text{Ba}^{2+}][\text{SO}_4^{2-}]$	s	s	$K_{\text{sp}} = s^2; s = (K_{\text{sp}})^{\frac{1}{2}}$
Ag ₂ CO ₃	$[\text{Ag}^+]^2[\text{CO}_3^{2-}]$	$2s$	s	$K_{\text{sp}} = 4s^3; s = \left(\frac{K_{\text{sp}}}{4}\right)^{\frac{1}{3}}$
PbF ₂	$[\text{Pb}^{2+}][\text{F}^-]^2$	s	$2s$	$K_{\text{sp}} = 4s^3; s = \left(\frac{K_{\text{sp}}}{4}\right)^{\frac{1}{3}}$
Al(OH) ₃	$[\text{Al}^{3+}][\text{OH}^-]^3$	s	$3s$	$K_{\text{sp}} = 27s^4; s = \left(\frac{K_{\text{sp}}}{27}\right)^{\frac{1}{4}}$
Ca ₃ (PO ₄) ₂	$[\text{Ca}^{2+}]^3[\text{PO}_4^{3-}]^2$	$3s$	$2s$	$K_{\text{sp}} = 108s^5; s = \left(\frac{K_{\text{sp}}}{108}\right)^{\frac{1}{5}}$

If 2.00 mL of 0.200 M NaOH are added to 1.00 L of 0.100 M CaCl_2 , will a precipitate form?

The ions present in solution are Na^+ , OH^- , Ca^{2+} , Cl^- .

Only possible precipitate is $\text{Ca}(\text{OH})_2$ (solubility rules).

Is $Q > K_{sp}$ for $\text{Ca}(\text{OH})_2$?

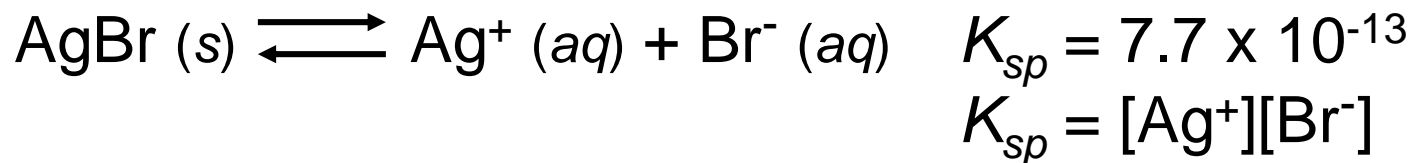
$$[\text{Ca}^{2+}]_0 = 0.100 \text{ M} \quad [\text{OH}^-]_0 = 4.0 \times 10^{-4} \text{ M}$$

$$Q = [\text{Ca}^{2+}]_0 [\text{OH}^-]_0^2 = 0.10 \times (4.0 \times 10^{-4})^2 = 1.6 \times 10^{-8}$$

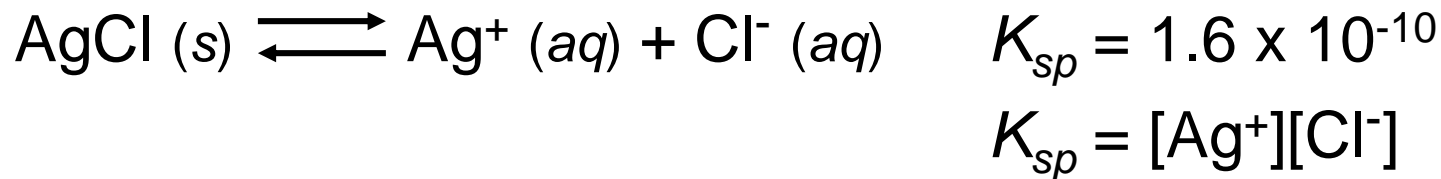
$$K_{sp} = [\text{Ca}^{2+}] [\text{OH}^-]^2 = 8.0 \times 10^{-6}$$

$$Q < K_{sp} \quad \text{No precipitate will form}$$

What concentration of Ag is required to precipitate ONLY AgBr in a solution that contains both Br⁻ and Cl⁻ at a concentration of 0.02 M?

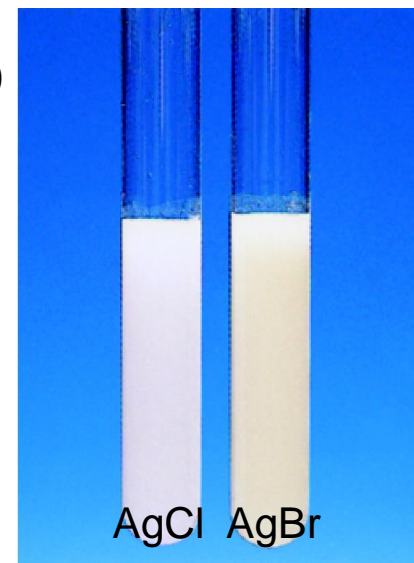


$$[\text{Ag}^+] = \frac{K_{sp}}{[\text{Br}^-]} = \frac{7.7 \times 10^{-13}}{0.020} = 3.9 \times 10^{-11} \text{ M}$$



$$[\text{Ag}^+] = \frac{K_{sp}}{[\text{Cl}^-]} = \frac{1.6 \times 10^{-10}}{0.020} = 8.0 \times 10^{-9} \text{ M}$$

$$3.9 \times 10^{-11} \text{ M} < [\text{Ag}^+] < 8.0 \times 10^{-9} \text{ M}$$



The Common Ion Effect and Solubility

The presence of a common ion **decreases** the solubility of the salt.

What is the molar solubility of AgBr in (a) pure water and (b) 0.0010 M NaBr?



$$K_{sp} = 7.7 \times 10^{-13}$$

$$s^2 = K_{sp}$$

$$s = 8.8 \times 10^{-7}$$



$$[\text{Br}^-] = 0.0010 \text{ M}$$



$$[\text{Ag}^+] = s$$

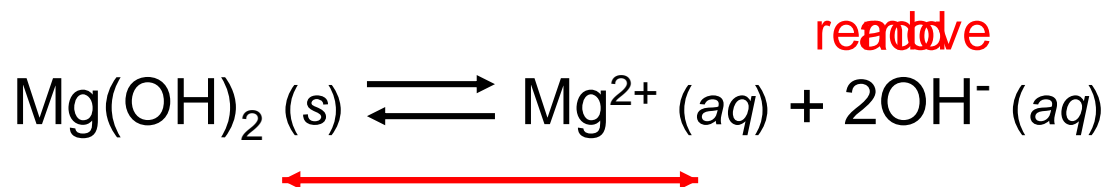
$$[\text{Br}^-] = 0.0010 + s \approx 0.0010$$

$$K_{sp} = 0.0010 \times s$$

$$s = 7.7 \times 10^{-10}$$

pH and Solubility

- The presence of a common ion **decreases** the solubility.
- Insoluble bases dissolve in acidic solutions
- Insoluble acids dissolve in basic solutions



At pH less than 10.45

$$K_{sp} = [\text{Mg}^{2+}][\text{OH}^-]^2 = 1.2 \times 10^{-11} \quad \text{Lower } [\text{OH}^-]$$

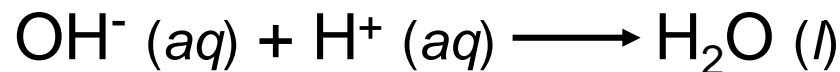
$$K_{sp} = (s)(2s)^2 = 4s^3$$

$$4s^3 = 1.2 \times 10^{-11}$$

$$s = 1.4 \times 10^{-4} \text{ M}$$

$$[\text{OH}^-] = 2s = 2.8 \times 10^{-4} \text{ M}$$

$$\text{pOH} = 3.55 \quad \text{pH} = 10.45$$



Increase solubility of Mg(OH)_2

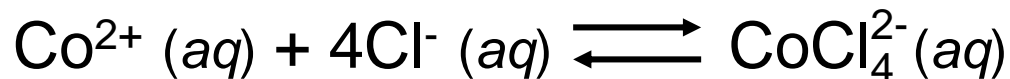
At pH greater than 10.45

Raise $[\text{OH}^-]$

Decrease solubility of Mg(OH)_2

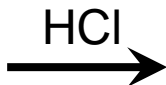
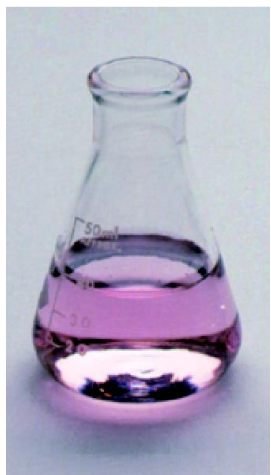
Complex Ion Equilibria and Solubility

A **complex ion** is an ion containing a central metal cation bonded to one or more molecules or ions.



The **formation constant or stability constant (K_f)** is the equilibrium constant for the complex ion formation.

$$K_f = \frac{[\text{CoCl}_4^{2-}]}{[\text{Co}^{2+}][\text{Cl}^{-}]^4}$$



stability of
complex ↑

Effect of Complexation on Solubility

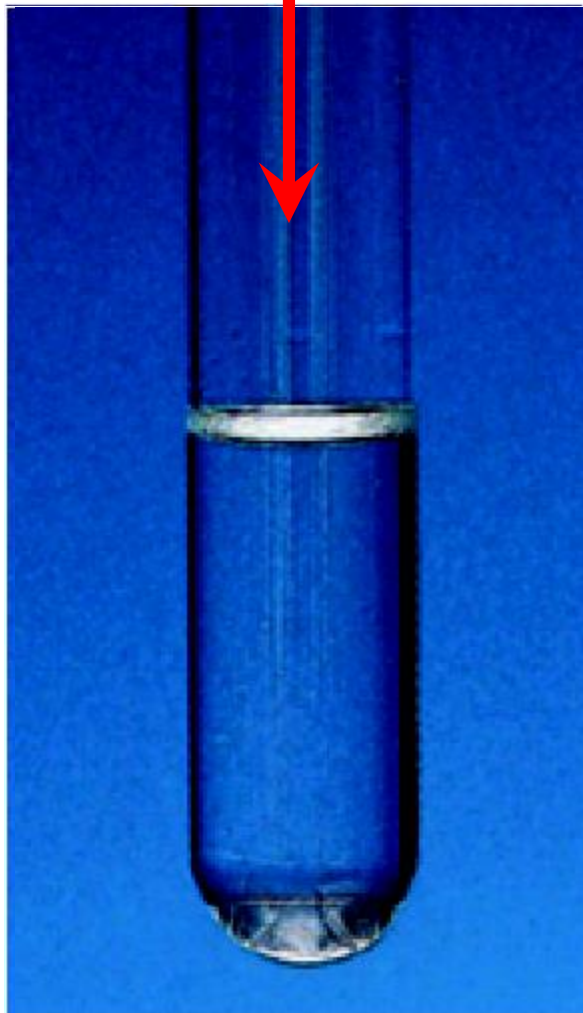
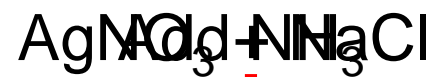


TABLE 16.4 **Formation Constants of Selected Complex Ions in Water at 25°C**

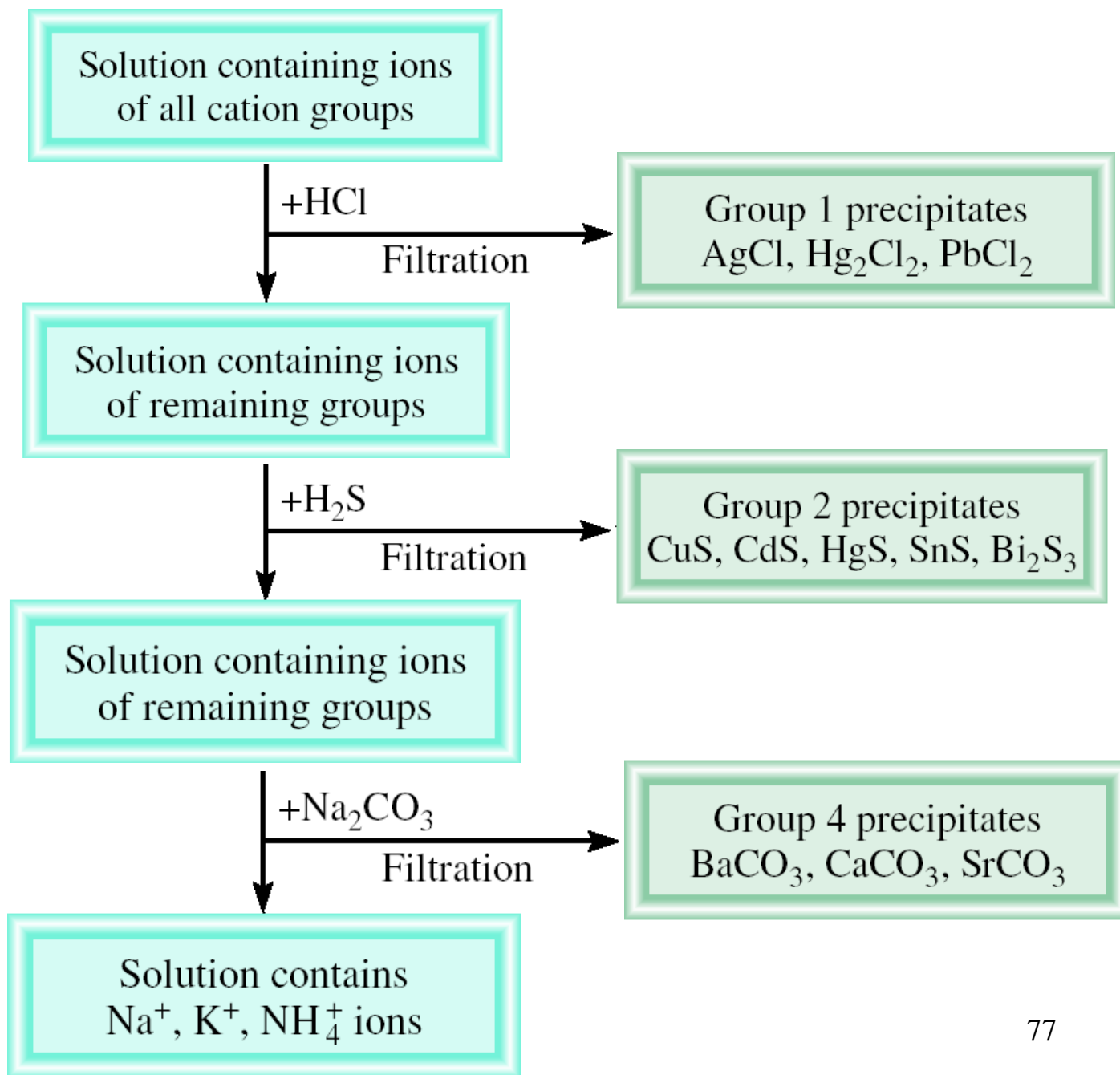
Complex Ion	Equilibrium Expression	Formation Constant (K_f)
$\text{Ag}(\text{NH}_3)_2^+$	$\text{Ag}^+ + 2\text{NH}_3 \rightleftharpoons \text{Ag}(\text{NH}_3)_2^+$	1.5×10^7
$\text{Ag}(\text{CN})_2^-$	$\text{Ag}^+ + 2\text{CN}^- \rightleftharpoons \text{Ag}(\text{CN})_2^-$	1.0×10^{21}
$\text{Cu}(\text{CN})_4^{2-}$	$\text{Cu}^{2+} + 4\text{CN}^- \rightleftharpoons \text{Cu}(\text{CN})_4^{2-}$	1.0×10^{25}
$\text{Cu}(\text{NH}_3)_4^{2+}$	$\text{Cu}^{2+} + 4\text{NH}_3 \rightleftharpoons \text{Cu}(\text{NH}_3)_4^{2+}$	5.0×10^{13}
$\text{Cd}(\text{CN})_4^{2-}$	$\text{Cd}^{2+} + 4\text{CN}^- \rightleftharpoons \text{Cd}(\text{CN})_4^{2-}$	7.1×10^{16}
CdI_4^{2-}	$\text{Cd}^{2+} + 4\text{I}^- \rightleftharpoons \text{CdI}_4^{2-}$	2.0×10^6
HgCl_4^{2-}	$\text{Hg}^{2+} + 4\text{Cl}^- \rightleftharpoons \text{HgCl}_4^{2-}$	1.7×10^{16}
HgI_4^{2-}	$\text{Hg}^{2+} + 4\text{I}^- \rightleftharpoons \text{HgI}_4^{2-}$	2.0×10^{30}
$\text{Hg}(\text{CN})_4^{2-}$	$\text{Hg}^{2+} + 4\text{CN}^- \rightleftharpoons \text{Hg}(\text{CN})_4^{2-}$	2.5×10^{41}
$\text{Co}(\text{NH}_3)_6^{3+}$	$\text{Co}^{3+} + 6\text{NH}_3 \rightleftharpoons \text{Co}(\text{NH}_3)_6^{3+}$	5.0×10^{31}
$\text{Zn}(\text{NH}_3)_4^{2+}$	$\text{Zn}^{2+} + 4\text{NH}_3 \rightleftharpoons \text{Zn}(\text{NH}_3)_4^{2+}$	2.9×10^9

TABLE 16.5

Separation of Cations into Groups According to Their Precipitation Reactions with Various Reagents

Group	Cation	Precipitating Reagents	Insoluble Compound	K_{sp}
1	Ag^+	HCl ↓	AgCl	1.6×10^{-10}
	Hg_2^{2+}		Hg_2Cl_2	3.5×10^{-18}
	Pb^{2+}		PbCl_2	2.4×10^{-4}
2	Bi^{3+}	H_2S in acidic solutions ↓	Bi_2S_3	1.6×10^{-72}
	Cd^{2+}		CdS	8.0×10^{-28}
	Cu^{2+}		CuS	6.0×10^{-37}
	Hg^{2+}		HgS	4.0×10^{-54}
	Sn^{2+}		SnS	1.0×10^{-26}
	Al^{3+}		$\text{Al}(\text{OH})_3$	1.8×10^{-33}
3	Co^{2+}	H_2S in basic solutions ↓	CoS	4.0×10^{-21}
	Cr^{3+}		$\text{Cr}(\text{OH})_3$	3.0×10^{-29}
	Fe^{2+}		FeS	6.0×10^{-19}
	Mn^{2+}		MnS	3.0×10^{-14}
	Ni^{2+}		NiS	1.4×10^{-24}
	Zn^{2+}		ZnS	3.0×10^{-23}
	Ba^{2+}		BaCO_3	8.1×10^{-9}
	Ca^{2+}		CaCO_3	8.7×10^{-9}
4	Sr^{2+}	Na_2CO_3 ↓	SrCO_3	1.6×10^{-9}
	K^+		None	
	Na^+		None	
5	NH_4^+	No precipitating reagent	None	

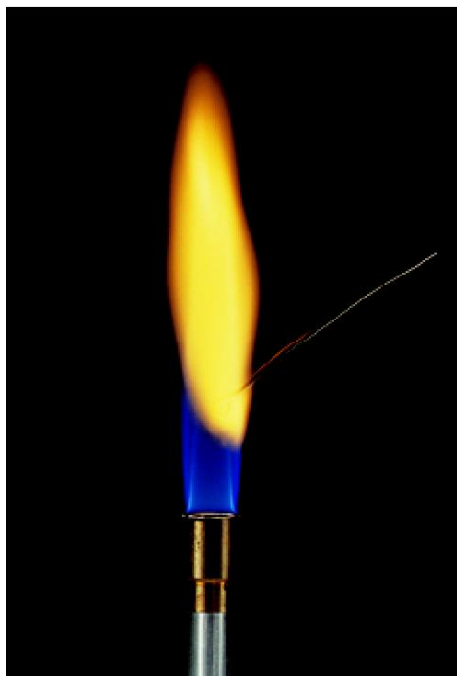
Qualitative Analysis of Cations



Flame Test for Cations



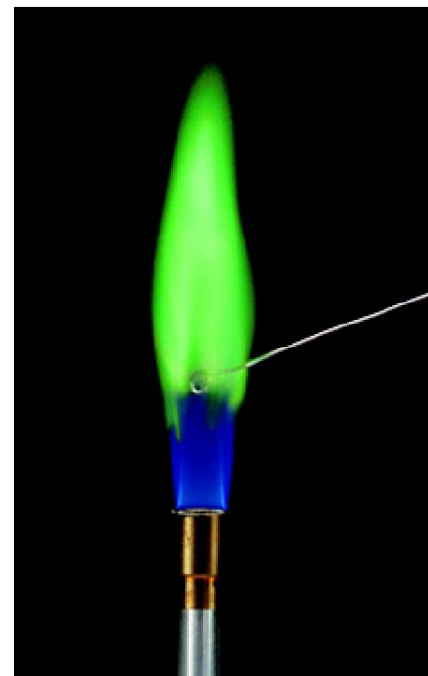
lithium



sodium

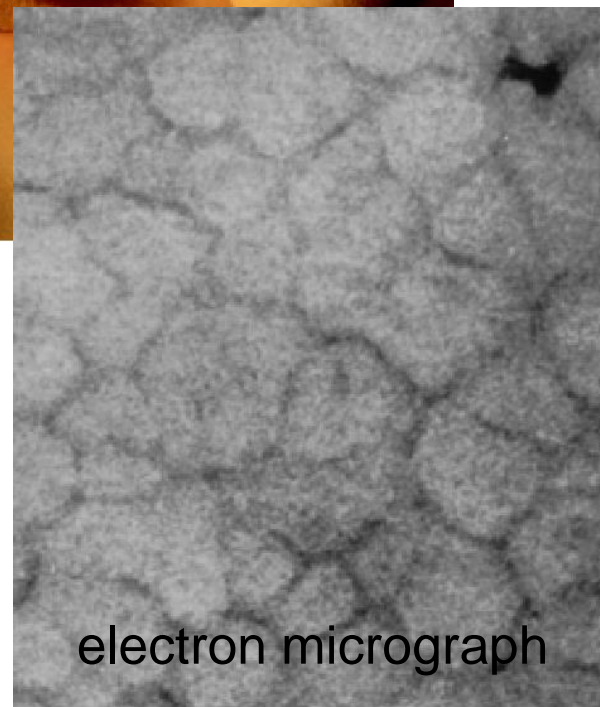
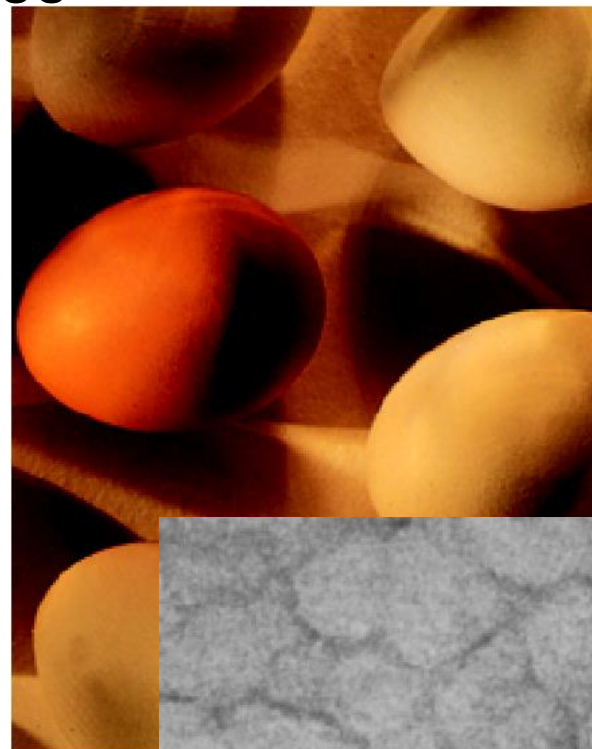
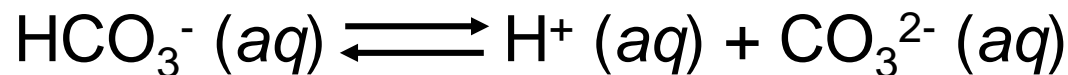
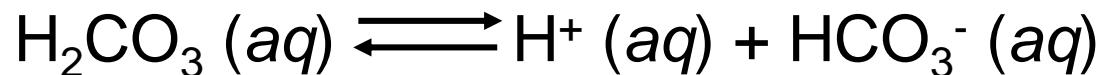
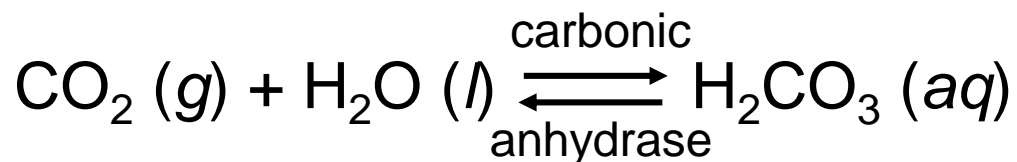
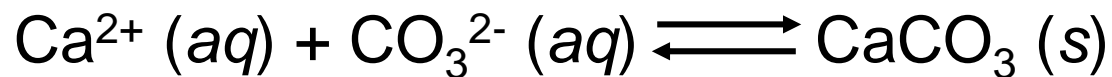


potassium



copper

Chemistry In Action: How an Eggshell is Formed



electron micrograph