

# Chapter 14

# Acids and Bases

Based on slides provided with Introductory Chemistry, Fifth Edition  
Nivaldo J. Tro

By Dr. Sylvia Esjornson  
Southwestern Oklahoma State University  
Weatherford, OK

With some modifications and additions by  
Dr. Deniz Cizmeciyan-Papazyan  
and  
Dr. Arno Papazyan

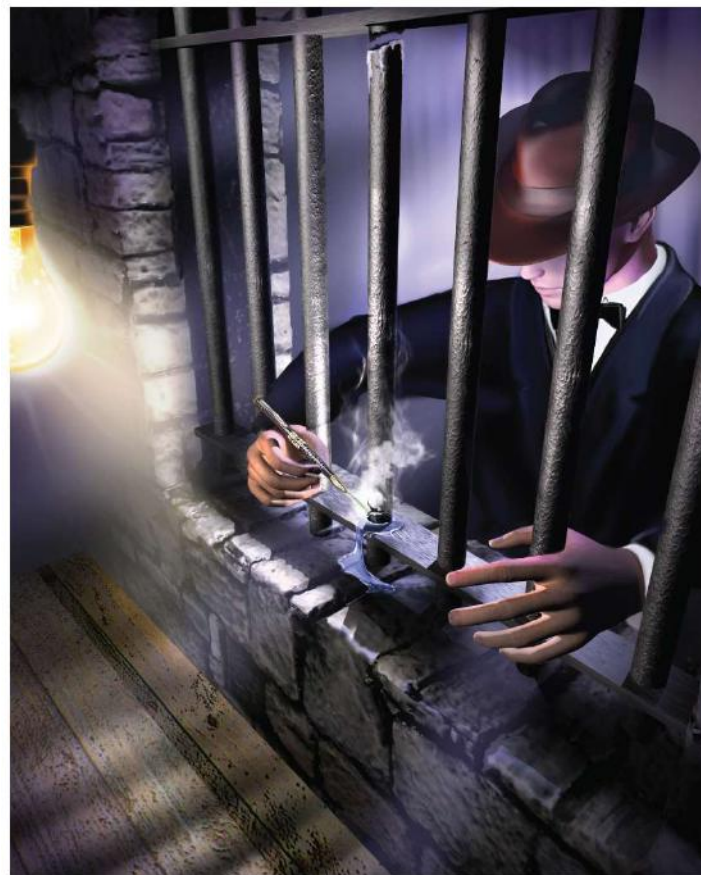
## Sour Patch Kids: Acids Taste Sour

- When you first put a Sour Patch Kid in your mouth, it tastes sour.
- The taste is caused by a white powder coating, a mixture of citric acid and tartaric acid.
- The citric and tartaric acids from a Sour Patch Kid combine with saliva in your mouth, and they produce  $H^+$  ions. Those  $H^+$  ions react with protein molecules on your tongue.
- The protein molecules change shape, sending an electrical signal (a nerve impulse) to your brain that you experience as a sour taste.



## Acids and Bases

- Acids have been used in spy movies and other thrillers to dissolve the metal bars of a prison cell.



## International Spy Movies: Acids Dissolve Many Metals

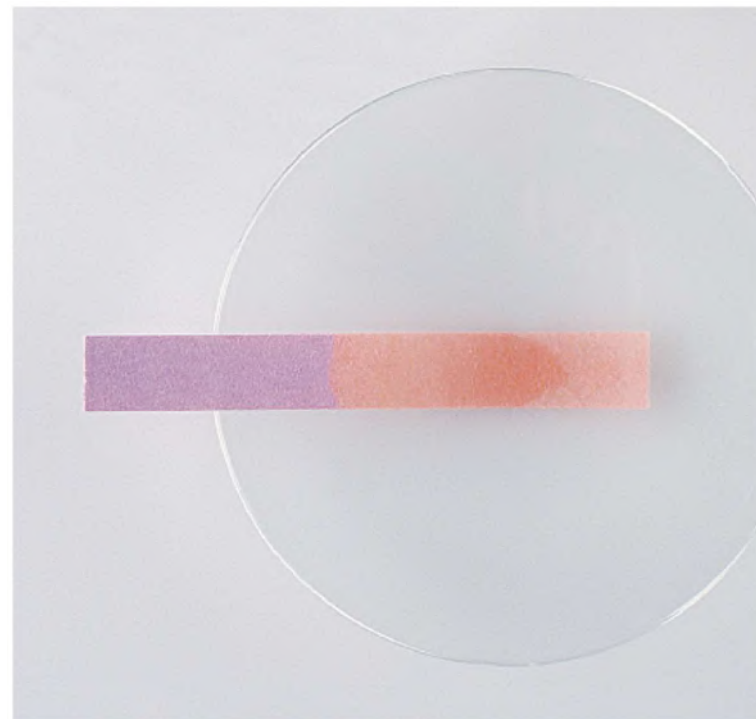
- James Bond often carries an acid-filled gold pen. Bond's pen is made of gold because gold is one of the few metals that is not dissolved by most acids.
- When Bond is imprisoned, he squirts some acid out of his pen and onto the iron bars of his cell.
- With enough acid, it would be possible to dissolve the iron bars of a prison cell, but it would take more acid than the amount that fits in a pen.
- We see here: A small piece of aluminum placed in hydrochloric acid dissolves away in about 10 minutes.



## Acids: Properties and Examples

**Acids** have the following properties:

- Acids have a sour taste.
- Acids dissolve many metals.
- Acids turn blue litmus paper red.



## Acid Examples: Hydrochloric Acid

- Hydrochloric acid is found in most chemistry laboratories.
- It is used in industry to clean metals, prepare and process foods, and refine metal ores.



Hydrochloric acid

- Hydrochloric acid is the main component of stomach acid.
- Hydrochloric acid helps break down food, and it kills harmful bacteria that might enter the body through food.
- The sour taste associated with indigestion is caused by the stomach's hydrochloric acid refluxing up into the esophagus.

## Acid Examples: Sulfuric and Nitric Acid



Sulfuric acid

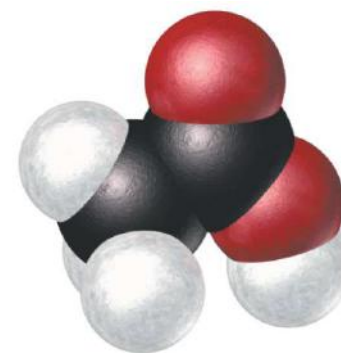
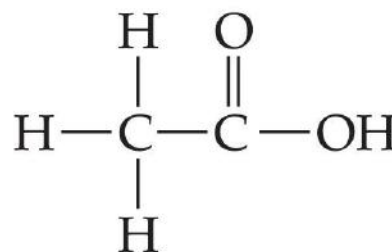


Nitric acid

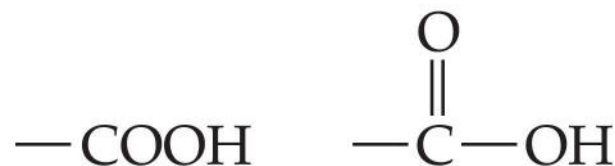
- Sulfuric acid is the most widely produced chemical in the United States; annual U.S. production of sulfuric acid exceeds 36 million tons.
- Sulfuric acid and nitric acid are commonly used in the laboratory.
- Sulfuric acid and nitric acid are used in the manufacture of fertilizers, explosives, dyes, and glue.
- Sulfuric acid is contained in most automobile batteries.

## Acid Examples: Acetic Acid

- Acetic acid forms in improperly stored wines.
- The word *vinegar* originates from the French *vin aigre*, which means “sour wine.”
- Acetic acid is an example of a **carboxylic acid**, an acid containing the COOH grouping of atoms, known as the carboxylic acid group.



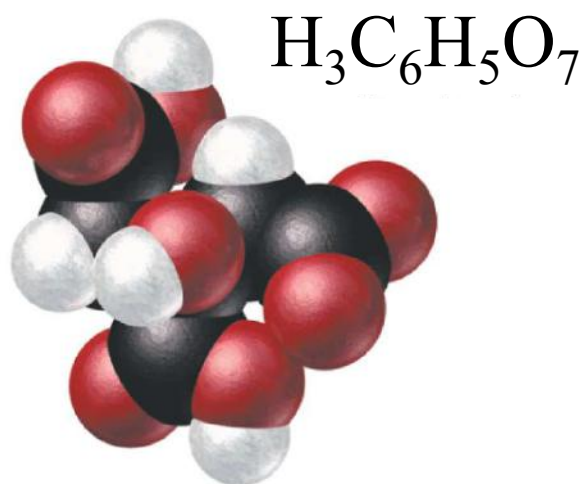
Acetic acid



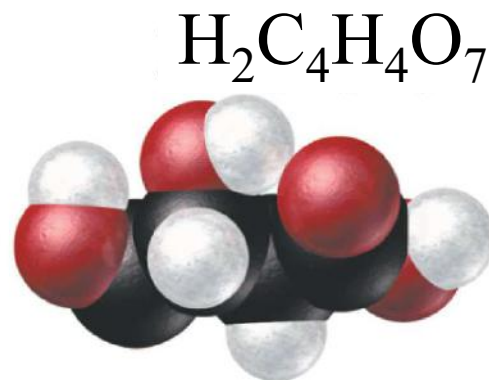
Carboxylic acid group



## Acid Examples: Carboxylic Acids



Citric acid



Malic acid

- We often find carboxylic acids in substances derived from living organisms.
- Other carboxylic acids include citric acid, the main acid in lemons and limes, and malic acid, an acid found in apples, grapes, and wine.

## Bases: Properties and Examples

**Bases** have the following properties:

- Bases have a bitter taste.  
Probably because they turn oil in living tissue into soap, which is bitter.  
For example, quinoa has a natural soapy coating to taste bitter and be unattractive to pests.
- Bases have a slippery feel.
- Bases turn red litmus paper blue.

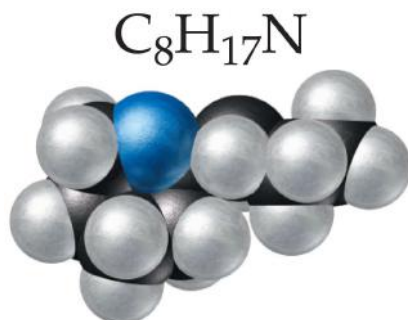


## Bases Feel Slippery

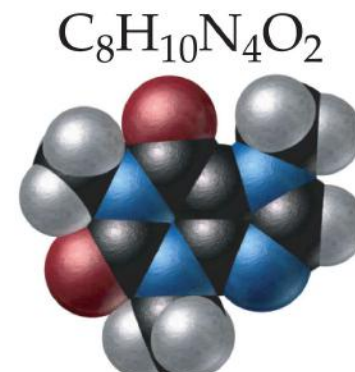
- All of these consumer products contain bases.
- Bases feel slippery because they react with oils on your skin to form soaplike substances.
- Soap feels slippery because:
  - Soap molecules have a polar “head” and an “oily” “tail”. The oily tails go to the surface of the solution because their surface tension (energy penalty for being on the surface) is lower.
  - Oily materials have lower viscosity, so offer less resistance for slipping
- Some household cleaning solutions, such as ammonia, are also basic and have the typical slippery feel of a base.



## Bases Taste Bitter



Coniine



Caffeine

- Bases are less common in foods than acids because of their bitter taste.
- Our aversion to the taste of bases is probably an adaptation to protect us against **alkaloids**, organic bases found in plants.
- Alkaloids are often poisonous—the active component of hemlock, for example, is the alkaloid coniine—and their bitter taste warns us against eating them.
- Some foods contain small amounts of base.
- Coffee is acidic overall, but bases present in coffee—such as caffeine—impart a bitter flavor.

## Base Examples: Sodium Hydroxide, Potassium Hydroxide, Sodium Bicarbonate

- Sodium hydroxide and potassium hydroxide are found in most chemistry laboratories.
- They are used in processing petroleum and cotton, and in soap and plastic manufacturing.
- Sodium hydroxide is the active ingredient in products such as Drano that work to unclog drains.
- Sodium bicarbonate can be found in most homes as baking soda and is an active ingredient in many antacids.
  - Technically is a salt with one acidic hydrogen remaining, but acts as a base for reasons that will become clear later
- When taken as an antacid, sodium bicarbonate neutralizes stomach acid, relieving heartburn and sour stomach.

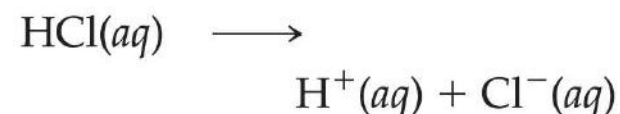
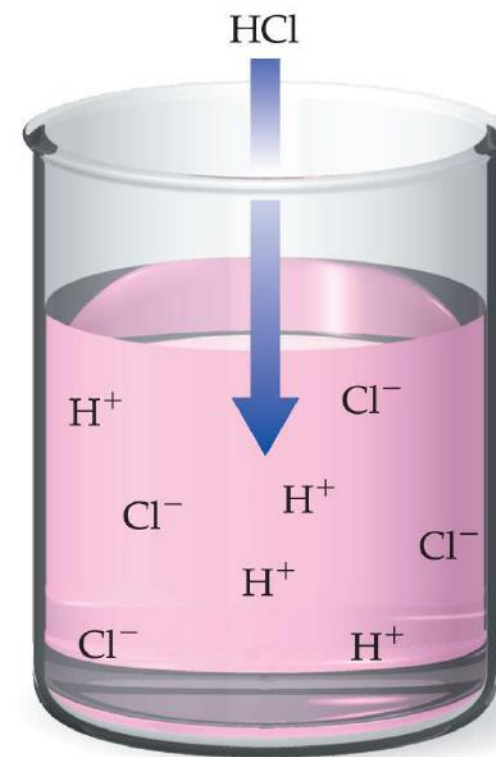
# The Arrhenius Definition of Acids

## Acid:

A compound that produces  $\text{H}^+$  ions when dissolved in water.

$\text{HCl}$  is a covalent compound and does not contain ions.

In water,  $\text{HCl}$  ionizes to form  $\text{H}^+$  ions and  $\text{Cl}^-$  ions.



## The Arrhenius Definition of Acids and Bases

- The  $\text{H}^+$  ions are highly reactive. In aqueous solution, they bond to water molecules according to the reaction.

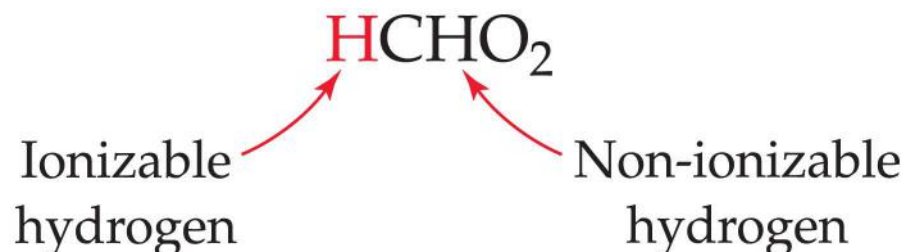


The  $\text{H}_3\text{O}^+$  ion is the **hydronium ion**.

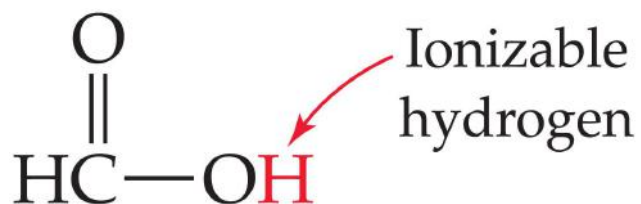
- In water,  $\text{H}^+$  ions *always* associate with  $\text{H}_2\text{O}$  molecules.
- Chemists often use  $\text{H}^+(\text{aq})$  and  $\text{H}_3\text{O}^+(\text{aq})$  interchangeably to refer to a hydronium ion.

## In the molecular formula for an acid, we write the ionizable (acidic) hydrogen First

We write the formula for formic acid as:



We represent the *structure* of formic acid with its structural formula:



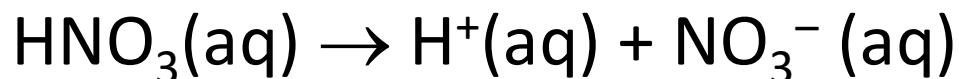
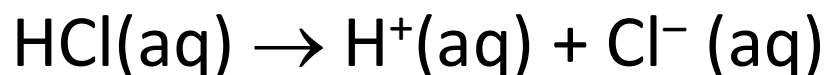
The structural formula indicates how the atoms are bonded together; the molecular formula indicates the number of each type of atom.



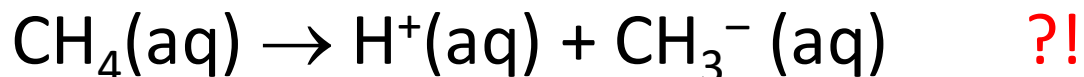
## What makes the acidic hydrogen acidic?

When a  $\text{H}^+$  leaves the acid molecule, it leaves behind an anion.

The following substances produce anions that are quite stable



But if we take methane,  $\text{CH}_4$ , for example



The reaction is extremely unfavorable because  $\text{CH}_3^-$  is a high-energy, unstable anion and being dissolved in water (favorable attractions by water molecules) won't be enough to stabilize it.

# The Arrhenius Definition of Bases

## Base

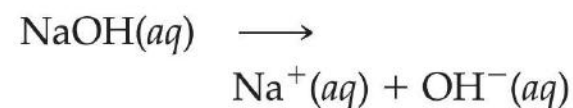
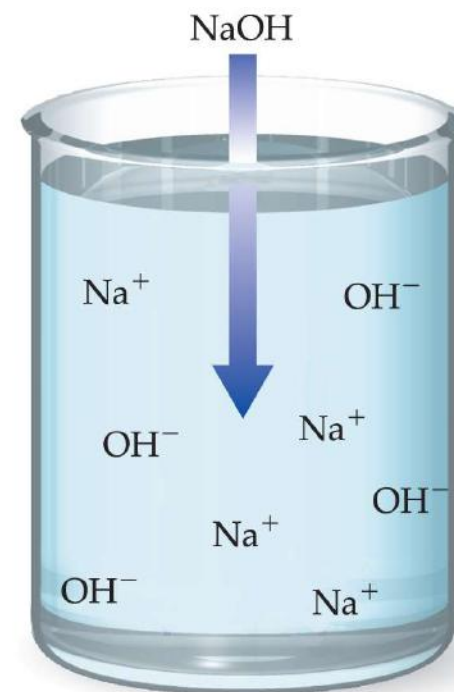
A substance that produces  $\text{OH}^-$  ions when dissolved in water

$\text{NaOH}$  is an Arrhenius base because it produces  $\text{OH}^-$  ions in solution.

$\text{NaOH}$  is an ionic compound and contains  $\text{Na}^+$  ions and  $\text{OH}^-$  ions.

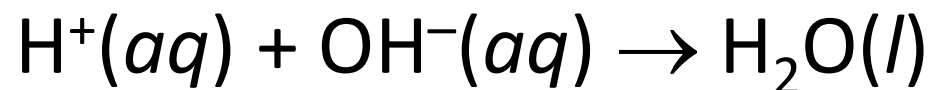
When  $\text{NaOH}$  is added to water, it dissociates into its component ions.

Molecular compounds containing an  $\text{OH}$  group, such as methanol,  $\text{CH}_3\text{OH}$ , do not dissociate in solution and therefore do not act as bases.



## The Arrhenius Definition of Acids and Bases

Under the Arrhenius definition, acids and bases naturally combine to form water, neutralizing each other in the process.



- Note that Arrhenius definition of acids and bases doesn't require that an acid (or a base) be in an aqueous solution to be called an acid or base
  - A common misreading of the original statement
  - It's the capacity to produce  $H^+$  or  $OH^-$  in aqueous solution that makes a substance an acid or a base
- For example,  $HNO_3$  in pure, non-aqueous form, is still called "nitric acid"
  - It then produces  $H^+$  (and  $NO_3^-$ ) ions when in aqueous solution
  - What makes it an acid is its ability to produce  $H^+$  ions when dissolved in water

- We don't really need (*aq*) after the acid formula for it to be called an acid
  - except for those whose anion name ends with “ide”, as a matter of ancient chemical tradition
    - HCl (*g*) “hydrogen chloride”
    - HCl(*aq*) “hydrochloric acid”
- But we normally have acids and bases in already aqueous form anyway when we use them in the lab
- **In future classes: Make sure you are aware of your instructor's expectations about the (*aq*) after acid formulas**

# The Brønsted–Lowry Definition of Acids and Bases

## Brønsted–Lowry definition

- **Acid**—An acid is a proton *donor*.
- **Base**—A base is a proton *acceptor*.
  
- In 1923, Johannes Brønsted, working in Denmark, and Thomas Lowry, working in England, developed the concept of proton transfer in acid–base behavior independently and simultaneously.

## The Brønsted–Lowry Definition of Acids

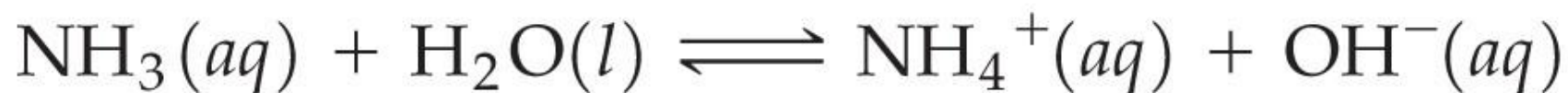
- According to this definition, HCl is a **Brønsted–Lowry acid** because, in solution, it donates a proton to water.



- This definition more clearly accounts for what happens to the  $\text{H}^+$  ion from an acid: it associates with a water molecule to form  $\text{H}_3\text{O}^+$  (a hydronium ion).

## The Brønsted–Lowry Definition of Bases

- The Brønsted–Lowry definition works well with bases such as  $\text{NH}_3$  that do not inherently contain  $\text{OH}^-$  ions but still produce  $\text{OH}^-$  ions in solution.
- $\text{NH}_3$  is a **Brønsted–Lowry base** because it accepts a proton from water.



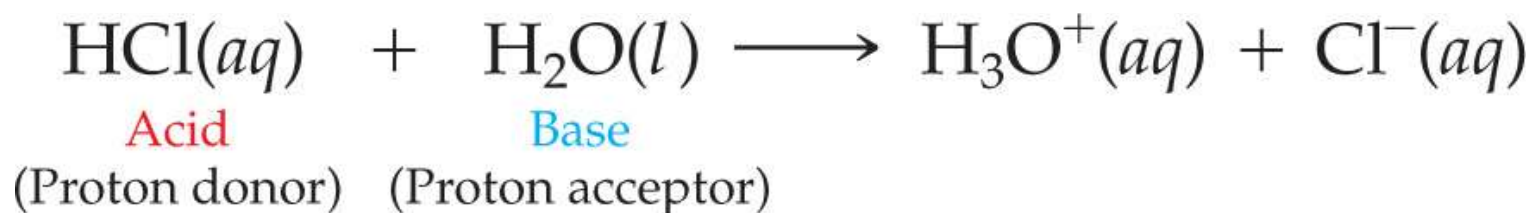


## Comparing the Arrhenius and the Brønsted–Lowry Definitions of Acids and Bases

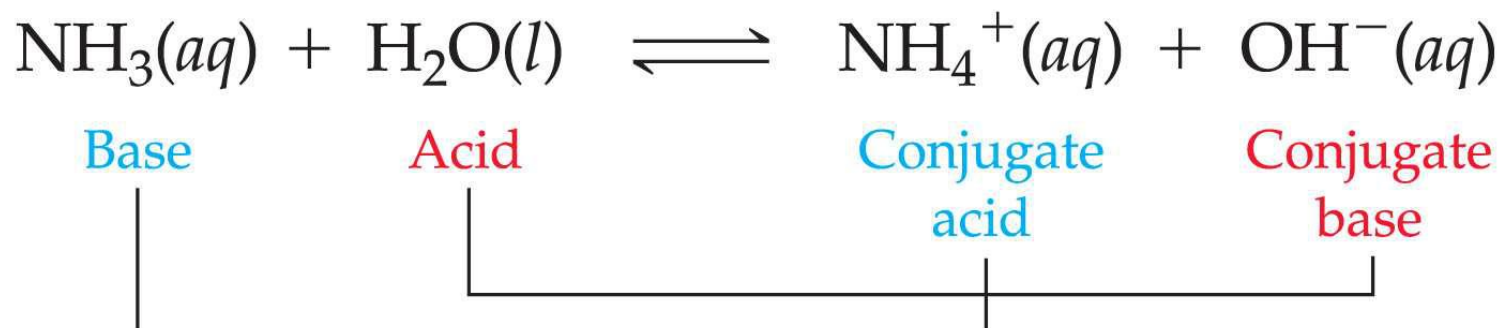
- The Arrhenius definition of acids and bases cannot explain why some substances act as bases even though they do not contain  $\text{OH}^-$ .
- The Arrhenius definition does not apply to nonaqueous solvents.
- The **Brønsted–Lowry definition of acids and bases** applies to a wider range of acid–base phenomena.
- This definition focuses on the *transfer* of  $\text{H}^+$  ions in an acid–base reaction.
- Since an  $\text{H}^+$  ion is a proton—a hydrogen atom with its electron taken away—this definition focuses on the idea of a **proton donor** and a **proton acceptor**.
- Important: The name “proton” can be misleading. This has nothing to do with any changes in the nucleus or nuclear reaction.  $\text{H}^+$  ions are acting chemically only.

## In the Brønsted–Lowry Definition, Acids (Proton Donors) and Bases (Proton Acceptors) Always Occur Together

- In the reaction between HCl and H<sub>2</sub>O, HCl is the proton donor (acid) and H<sub>2</sub>O is the proton acceptor (base).

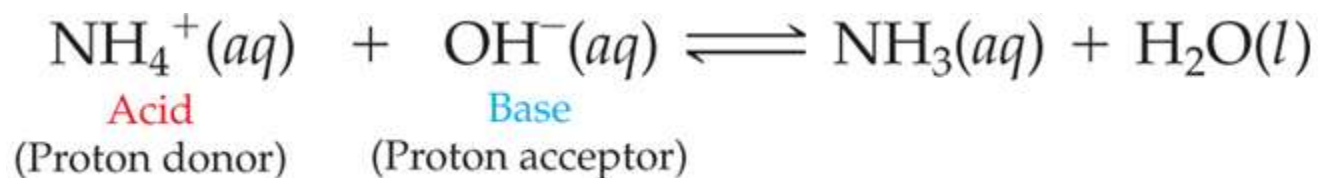


- In the reaction between NH<sub>3</sub> and H<sub>2</sub>O, H<sub>2</sub>O is the proton donor (acid) and NH<sub>3</sub> is the proton acceptor (base).



## Acid Base Defined by Behavior

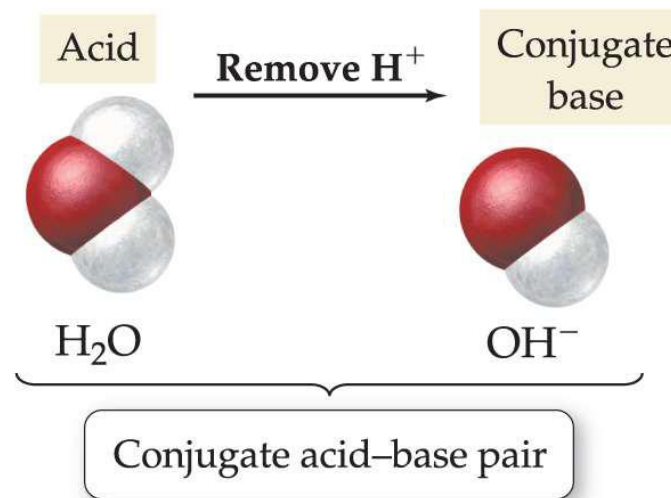
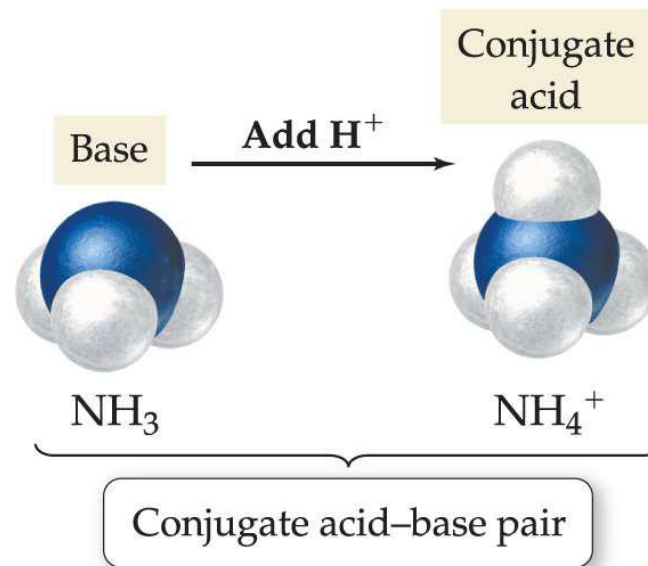
- Some substances—such as water—can act as acids *or* bases either accepting or donating a proton.
- Substances that can act as acids or bases are **amphoteric**.
- What happens when an equation representing Brønsted–Lowry acid–base behavior is reversed?



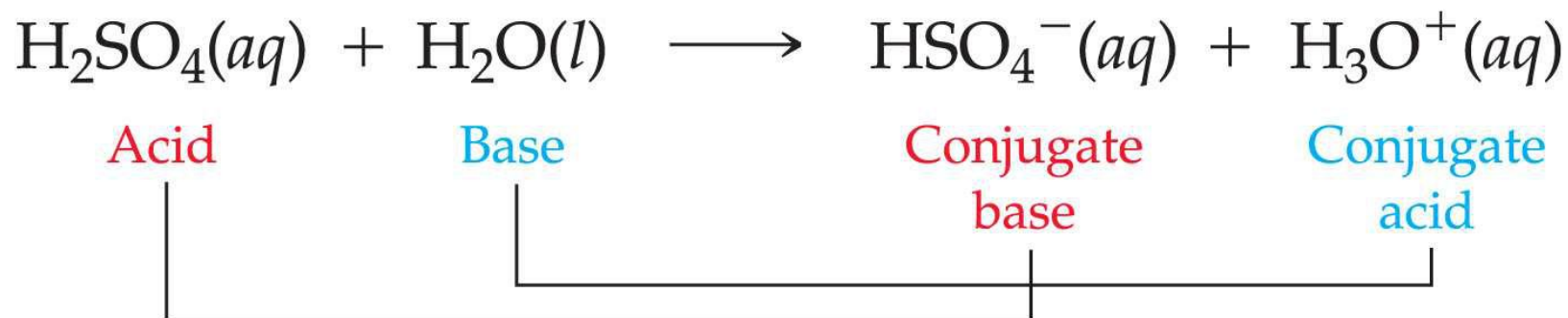
- Now  $\text{NH}_4^+$  is the proton donor (acid) and  $\text{OH}^-$  is the proton acceptor (base).
- What was the base ( $\text{NH}_3$ ) has become the acid ( $\text{NH}_4^+$ ), and what was the acid ( $\text{H}_2\text{O}$ ) has become the base ( $\text{OH}^-$ ).
- $\text{NH}_3$  and  $\text{NH}_4^+$  are often referred to as a **conjugate acid–base pair**, two substances related to each other by the transfer of a proton.

## A Conjugate Acid–Base Pair

Any two substances related to each other by the transfer of a proton can be considered a conjugate acid–base pair.



## Identifying Brønsted–Lowry Acids and Bases and Their Conjugates



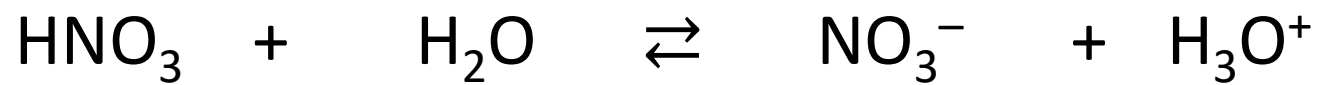
In an acid–base reaction:

- A base accepts a proton and becomes a conjugate acid.
- An acid donates a proton and becomes a conjugate base.

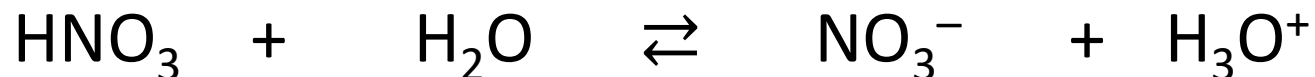
## Practice:

Check your solution on the next slide

Identify the Brønsted-Lowry acids and bases, and their conjugates, in the reaction

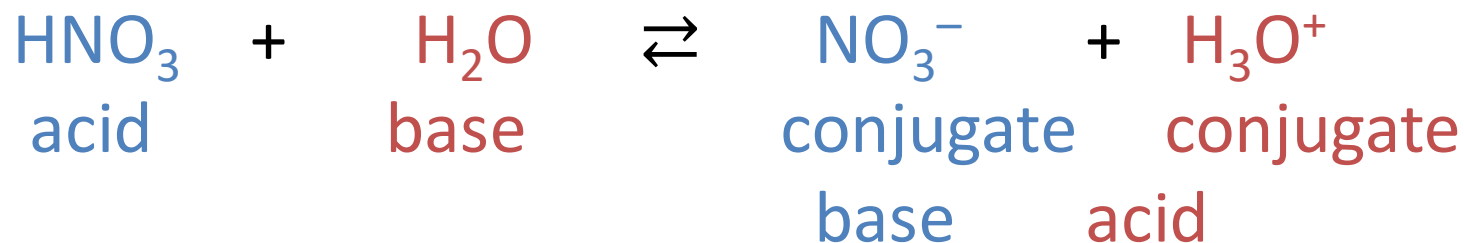


**Practice:** identify the Brønsted-Lowry acids and bases, and their conjugates, in the reaction



When the  $\text{H}_2\text{SO}_4$  becomes  $\text{HSO}_4^-$ , it loses an  $\text{H}^+$  so  $\text{H}_2\text{SO}_4$  must be the acid and  $\text{HSO}_4^-$  its conjugate base

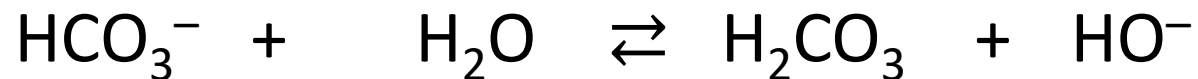
When the  $\text{H}_2\text{O}$  becomes  $\text{H}_3\text{O}^+$ , it accepts an  $\text{H}^+$  so  $\text{H}_2\text{O}$  must be the base and  $\text{H}_3\text{O}^+$  its conjugate acid



## Practice

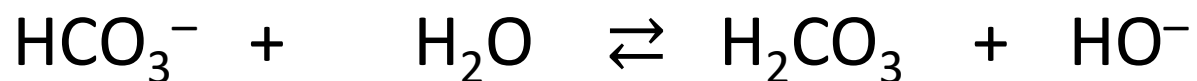
Check your answer on the next slide

Identify the Brønsted-Lowry acids and bases and their conjugates in the reaction





**Practice:** Identify the Brønsted-Lowry acids and bases and their conjugates in the reaction



When the  $\text{HCO}_3^-$  becomes  $\text{H}_2\text{CO}_3$ , it accepts an  $\text{H}^+$  so  $\text{HCO}_3^-$  must be the base and  $\text{H}_2\text{CO}_3$  its conjugate acid

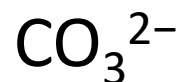
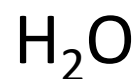
When the  $\text{H}_2\text{O}$  becomes  $\text{OH}^-$ , it donates an  $\text{H}^+$  so  $\text{H}_2\text{O}$  must be the acid and  $\text{OH}^-$  its conjugate base



## Practice

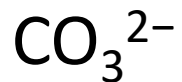
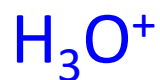
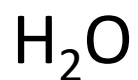
Check your answer on the next slide

Write the formula for the conjugate acid of the following



# Practice

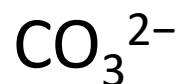
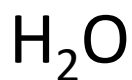
Write the formula for the conjugate acid of the following



## Practice

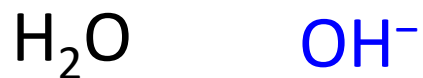
**Check your solution on the next slide**

Write the formula for the conjugate base of the following



## Practice

Write the formula for the conjugate base of the following

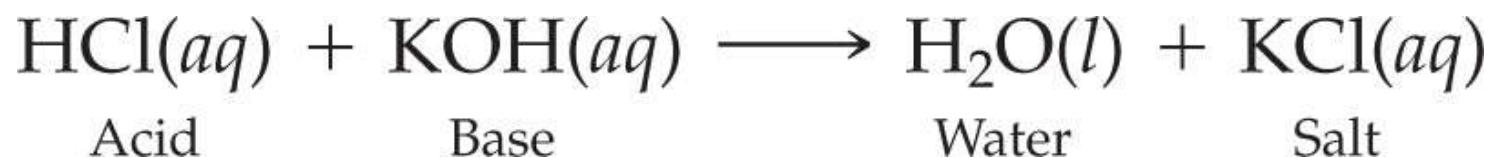


$\text{CO}_3^{2-}$       does not have an H, it cannot be an acid,  
so it cannot have a conjugate base

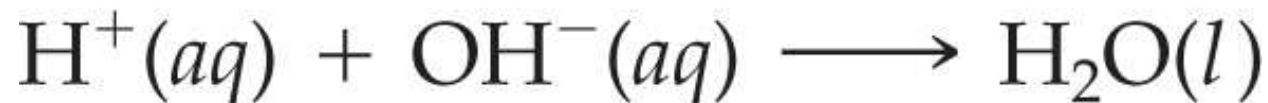


## Reactions of Acids and Bases: Neutralization Reactions

- When an acid and a base are mixed, the  $\text{H}^+$  from the acid combines with the  $\text{OH}^-$  from the base to form  $\text{H}_2\text{O}$ .

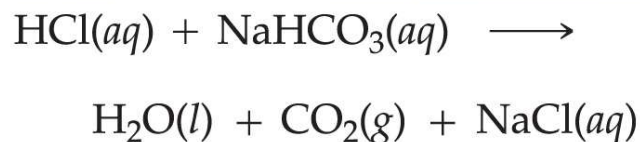


- Acid–base reactions generally form water and a **salt**—an ionic compound—that usually remains dissolved in the solution. The salt contains the cation from the base and the anion from the acid.
- The net ionic equation for many neutralization reactions is:



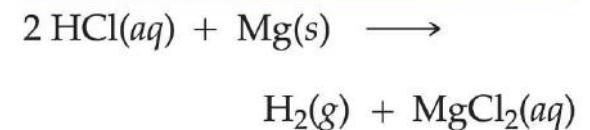
## Reactions of Acids and Bases: Neutralization Reactions

- The reaction of carbonates or bicarbonates (compounds containing  $\text{CO}_3^{2-}$  or  $\text{HCO}_3^-$ ) with acids produces water, gaseous carbon dioxide, and a salt.
  - because the produced conjugate acid  $\text{H}_2\text{CO}_3$  decomposes into  $\text{H}_2\text{O}$  and  $\text{CO}_2$
- This kind of reaction is also called a *gas evolution reaction*.



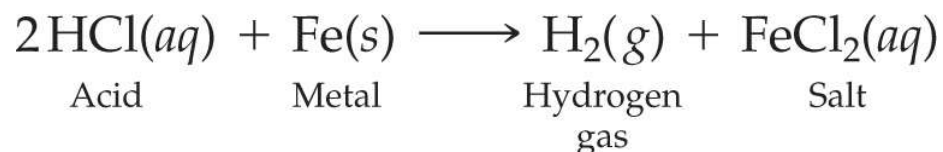
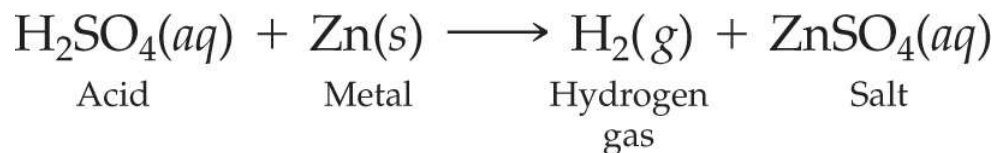
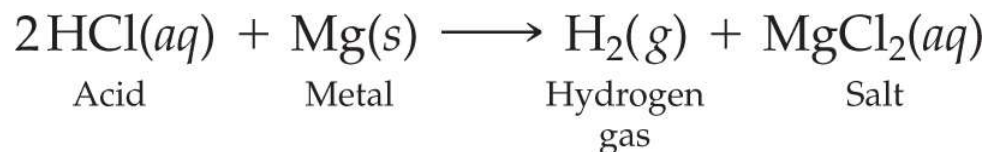
## Acids React with Metals in a Way That Causes Metals to Go Into Solution

The reaction between an acid and a metal usually produces hydrogen gas and a dissolved salt containing the metal ion.





## Acids Dissolve Many Metals



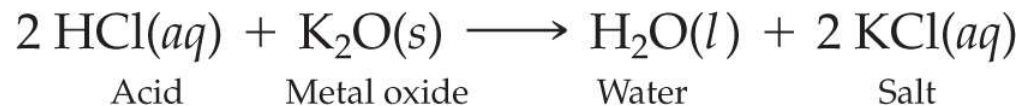
Some metals do not readily react with acids.

- If the bars that imprisoned James Bond were made of gold (or just copper), a pen filled with hydrochloric acid would not dissolve those bars.
- We will discuss the way to determine whether a particular metal dissolves in an acid in Chapter 16.

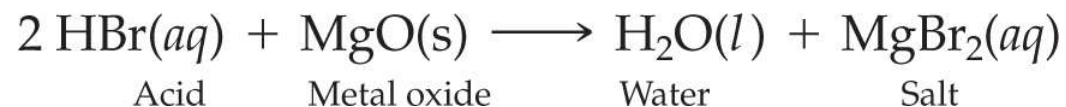
## Reactions of Acids with Metal Oxides

Acids react with metal oxides to produce water and a dissolved salt.

- Hydrochloric acid reacts with potassium oxide to form water and potassium chloride.

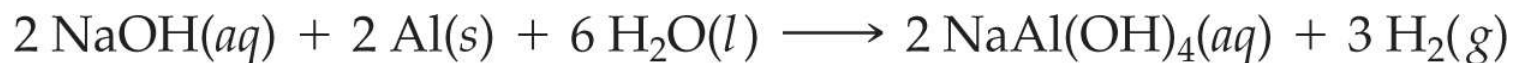


- Hydrobromic acid reacts with magnesium oxide to form water and magnesium bromide.



## Reactions of Bases: Neutralization Reactions, Reactions with Aluminum

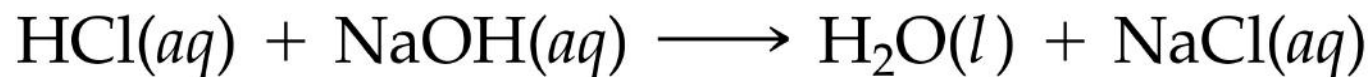
- The most important base reactions are those in which a base neutralizes an acid.
- One other kind of base reaction that we cover here is the reaction of sodium hydroxide with aluminum and water.



- Aluminum is one of the few metals that dissolves in a base.
- If your household plumbing pipes were made of aluminum, it would not be safe to use NaOH (the main ingredient in many drain-opening products) to unclog your drain.
- Fortunately, the use of aluminum pipe is forbidden in most building codes.

## Acid–Base Titration: A Way to Quantify the Amount of Acid or Base in a Solution

- Solution stoichiometry can be applied to a common laboratory procedure called a titration.
- In a **titration**, a substance in a solution of ***known*** concentration is reacted with another substance in a solution of ***unknown*** concentration.
- In the acid–base reaction between hydrochloric acid and sodium hydroxide, the net ionic equation is as follows:



## Representation of an Acid–Base Titration

- An HCl solution is represented by this molecular diagram.
- In the H<sup>+</sup> solution, the Cl<sup>-</sup> ions and the H<sub>2</sub>O molecules will not be involved in the reaction and have been omitted from the representation for clarity.
- In the OH<sup>-</sup> solution, the Na<sup>+</sup> ions and the H<sub>2</sub>O molecules will not be involved in the reaction and have been omitted from the representation for clarity.



# Representation of an Acid–Base Titration

- In titrating this sample, we slowly add a solution of known concentration to the solution of unknown concentration in the flask.
- At the equivalence point, neither reactant is present in excess, and both are limiting.
- The number of moles of the reactants are related by the reaction stoichiometry.

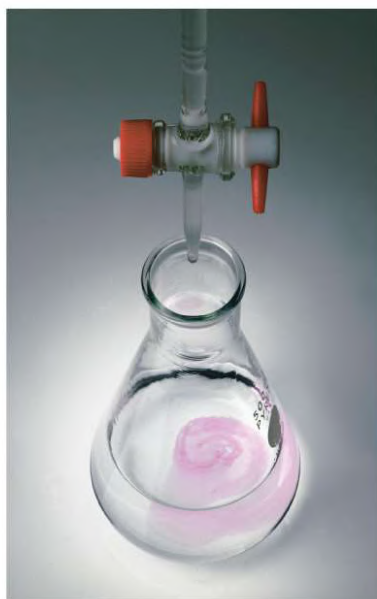


## Acid–Base Titration: A Way to Quantify the Amount of Acid or Base in a Solution

- As the  $\text{OH}^-$  is added, it reacts with and neutralizes the  $\text{H}^+$ , forming water.
- At the **equivalence point**—*the point in the titration when the number of moles of  $\text{OH}^-$  added equals the number of moles of  $\text{H}^+$  originally in solution*—the titration is complete.
- The equivalence point is usually signaled by an **indicator**, a dye whose color depends on the acidity of the solution.
- In most laboratory titrations, the concentration of one of the reactant solutions is unknown and the concentration of the other is precisely known.
- By carefully measuring the volume of each solution required to reach the equivalence point, the concentration of the unknown solution can be determined.

## In This Titration, NaOH Is Added to an HCl Solution

When the NaOH and HCl reach stoichiometric proportions (1 mol of  $\text{OH}^-$  for every 1 mol of  $\text{H}^+$ ), the indicator (phenolphthalein) changes to pink, signaling the equivalence point of the titration. Phenolphthalein is an indicator that is colorless in acidic solution and pink in basic solution.





## Acid–Base Titration Calculation

- The titration of 10.00 mL of an HCl solution of unknown concentration requires 12.54 mL of a 0.100 M NaOH solution to reach the equivalence point.
- What is the concentration of the unknown HCl solution?

### **GIVEN:**

- 10.00 mL HCl solution
- 12.54 mL of a 0.100 M NaOH solution

**FIND:** concentration of HCl solution

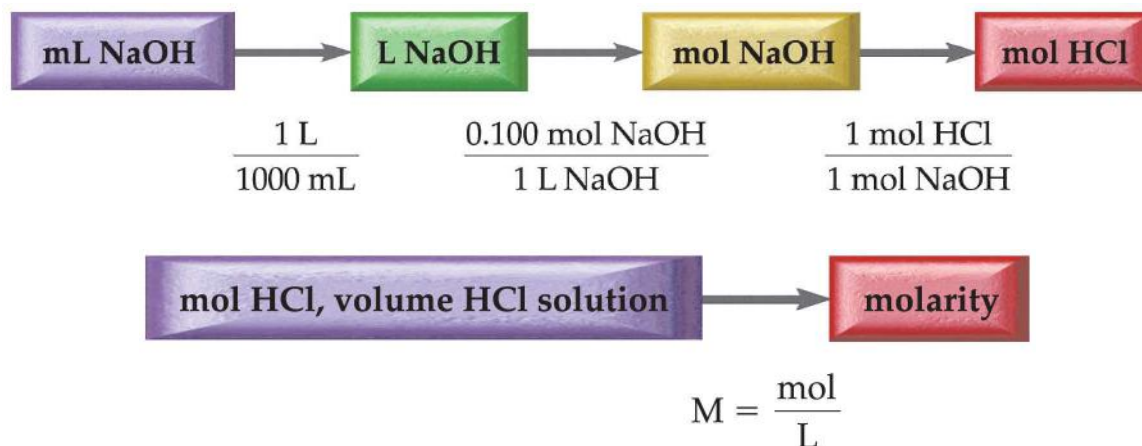
## Acid–Base Titration Calculation

- First, you must write the balanced chemical equation for the reaction between the acid and the base.



- The solution map has two parts.

SOLUTION MAP:



- The solution has two parts.

SOLUTION:

1 mol HCl:1 mol NaOH (from balanced chemical equation)

Molarity (M) =  $\frac{\text{mol solute}}{\text{L solution}}$  (definition of molarity, from Section 13.6)

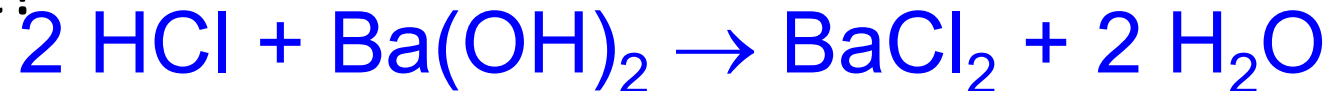
$$12.54 \text{ mL NaOH} \times \frac{1 \text{ L}}{1000 \text{ mL}} \times \frac{0.100 \text{ mol NaOH}}{1 \text{ L NaOH}} \times \frac{1 \text{ mol HCl}}{1 \text{ mol NaOH}} = 1.25 \times 10^{-3} \text{ mol HCl}$$

$$\text{Molarity} = \frac{1.25 \times 10^{-3} \text{ mol HCl}}{0.01000 \text{ L}} = 0.125 \text{ M}$$

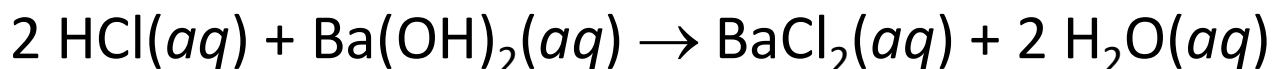
## Practice — Titration

Check your answer on the next slide

- 43.8 mL of 0.107 M HCl is needed to neutralize 37.6 mL of Ba(OH)<sub>2</sub> solution. What is the molarity of the base?



**Practice** – 43.8 mL of 0.107 M HCl is needed to neutralize 37.6 mL of Ba(OH)<sub>2</sub> solution. What is the molarity of the base?



<b>Given:</b>	43.8 mL of 0.107 M HCl, 37.6 mL Ba(OH) <sub>2</sub>
<b>Find:</b>	M Ba(OH) <sub>2</sub>
<b>Conceptual Plan:</b>	<pre> graph LR     A[L HCl] --&gt; B[mol HCl]     B --&gt; C[mol Ba(OH)2]     D[L Ba(OH)2] --&gt; E[M Ba(OH)2]     C --&gt; E             </pre>
<b>Relationships:</b>	1 mL = 0.001 L, 1 L HCl = 0.107 mol, 1 mol Ba(OH) <sub>2</sub> : 2 mol HCl $M = \frac{\text{mol}}{\text{L}}$

**Solution:**

$$43.8 \text{ mL} \times \frac{0.001 \text{ L}}{1 \text{ mL}} = 0.0438 \text{ L}$$

$$37.6 \text{ mL} \times \frac{0.001 \text{ L}}{1 \text{ mL}} = 0.0376 \text{ L}$$

$$0.0438 \text{ L HCl} \times \frac{0.107 \text{ mol HCl}}{1 \text{ L}} \times \frac{1 \text{ mol Ba}(\text{OH})_2}{2 \text{ mol HCl}} = 0.00234 \text{ mol Ba}(\text{OH})_2$$

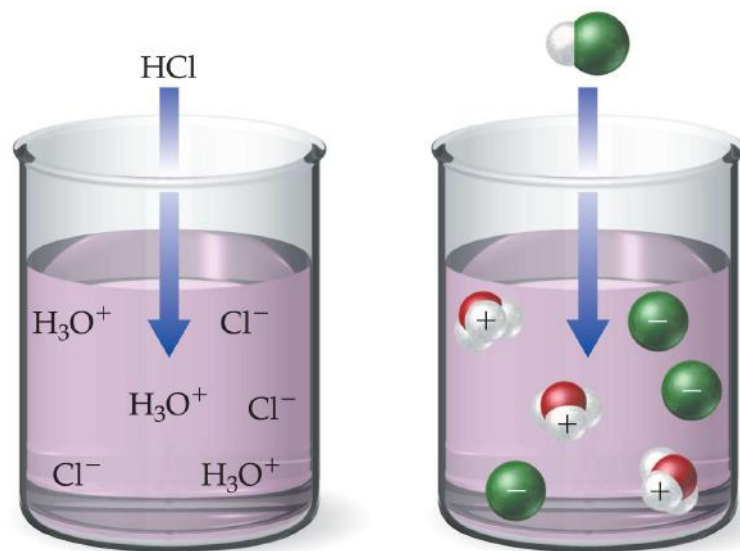
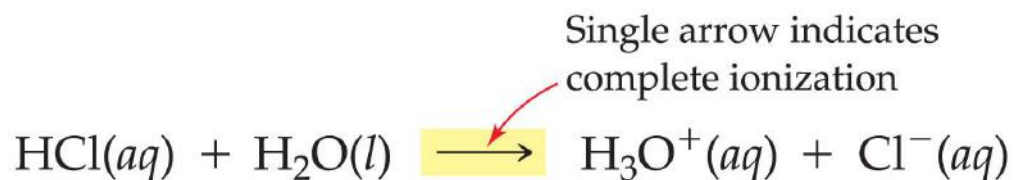
$$\frac{0.00234 \text{ mol Ba}(\text{OH})_2}{0.0376 \text{ L Ba}(\text{OH})_2} = 0.0623 \text{ M Ba}(\text{OH})_2$$

## Strong and Weak Acids and Bases: Strong Acids

- Hydrochloric acid (HCl) and hydrofluoric acid (HF) appear to be similar, but there is an important difference between these two acids.
- HCl is an example of a **strong acid**, one that completely ionizes in solution.
- HF is a **weak acid**, one that does not completely ionize in solution.
  - Actually the HF case is complicated. It only appears to be weak in dilute aqueous solution due to the super-strong hydrogen bonding its anion makes to the hydronium ion it just produced, but it is extremely strong when pure or nearly pure

## A Strong Acid

When HCl dissolves in water, it completely ionizes into  $\text{H}_3\text{O}^+$  and  $\text{Cl}^-$  ions. The solution contains no intact HCl.



## Conductivity of Pure Water

- Pure water is not a good conductor of electricity because it has relatively few charged particles.
- The danger of using electrical devices—such as a hair dryer—while sitting in the bathtub is that water is seldom pure and often contains dissolved ions.
- If the device were to come in contact with the water, dangerously high levels of electricity could flow through the water and through your body.

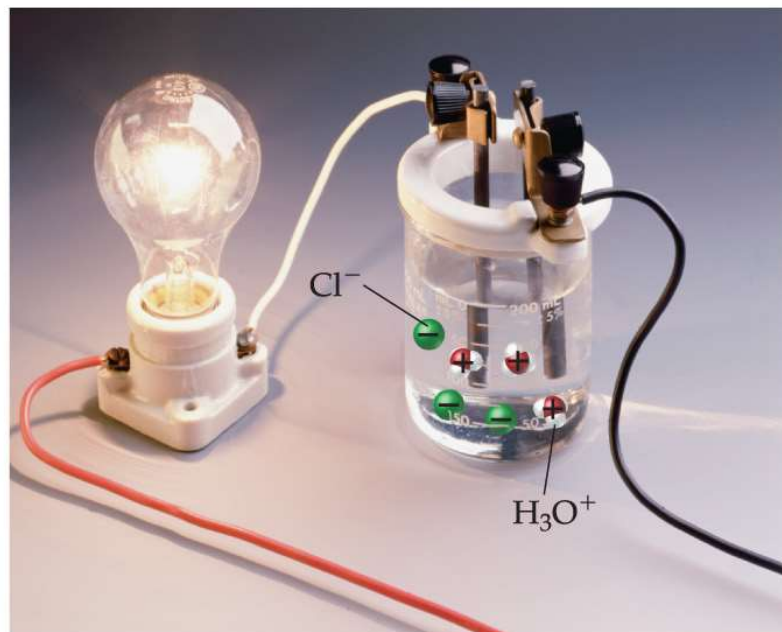


## Conductivity of a Strong Electrolyte Solution

- (a) Pure water will not conduct electricity.
- (b) The presence of ions in an HCl solution results in the conduction of electricity, causing the lightbulb to light.



(a) Pure water



(b) HCl solution

## Table of Strong Acids

- Six substances are considered strong acids in aqueous solution.
  - At introductory level chemistry
  - There are many more
- Five acids in the table are **monoprotic acids**, acids containing only one ionizable hydrogen.
- Sulfuric acid is an example of a **diprotic acid**, an acid that contains two ionizable hydrogen.
- An ionizable hydrogen is one that becomes an  $\text{H}^+$  ion (i.e. “proton”) in solution.

**TABLE 14.3 Strong Acids**

hydrochloric acid (HCl)	nitric acid (HNO <sub>3</sub> )
hydrobromic acid (HBr)	perchloric acid (HClO <sub>4</sub> )
hydroiodic acid (HI)	sulfuric acid (H <sub>2</sub> SO <sub>4</sub> ) ( <i>diprotic</i> )

almost

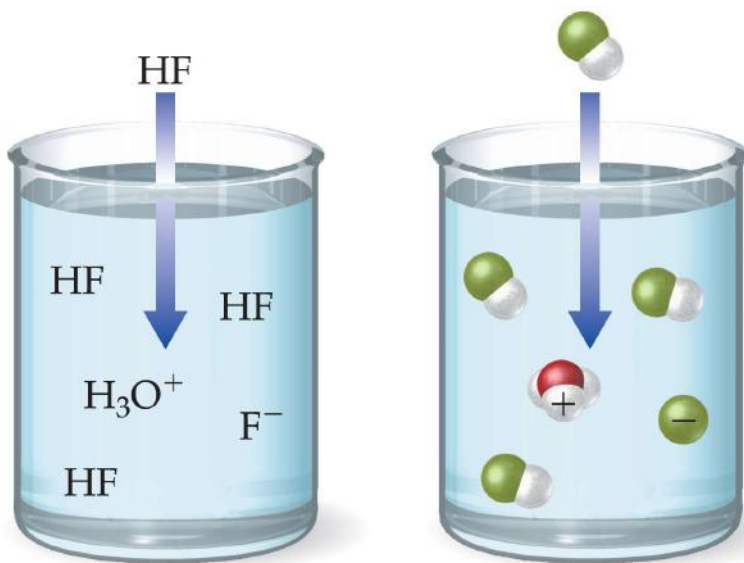
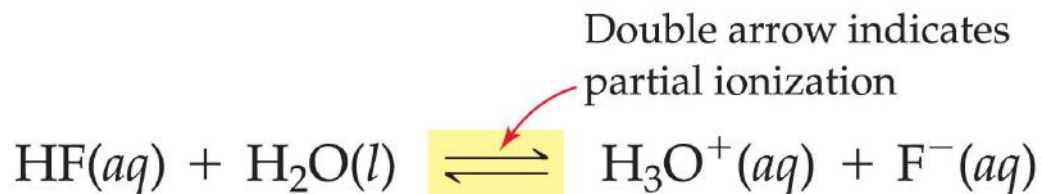


## Monoprotic, diprotic, triprotic, ....

- Monoprotic => can give one  $H^+$  cation
- Diprotic => can give two  $H^+$  cations
- Triprotic => can give three  $H^+$  cations
- The “prot” part refers to the  $H^+$  ion, which is a hydrogen atom that lost its electron
- A hydrogen atom without its electron is just a proton
- Don't think nuclear reactions (!)

## A Weak Acid

When HF dissolves in water, only a fraction of the dissolved molecules ionize into  $\text{H}_3\text{O}^+$  and  $\text{F}^-$  ions. The solution contains many intact HF molecules.\*



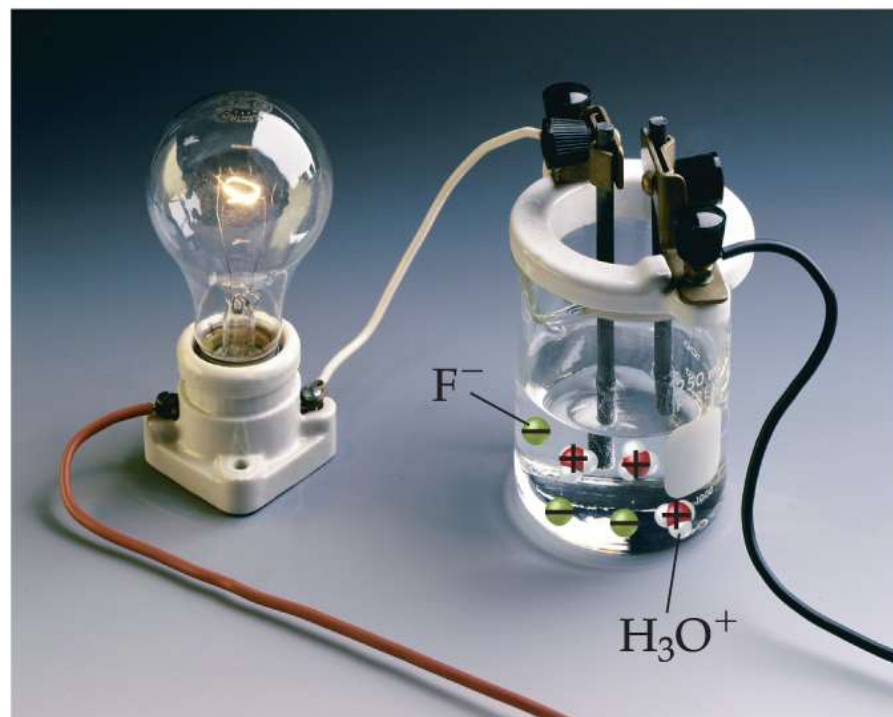
\* We will ignore the previously mentioned complexities of  $\text{HF}(aq)$  and regard any HF molecule that failed to produce truly free  $\text{H}^+$  and  $\text{F}^-$  ions as intact, which is OK.

## Conductivity of a Weak Electrolyte Solution

- (a) Pure water will not conduct electricity.
- (b) An HF solution contains some ions, but most of the HF is intact. The light glows only dimly.



(a) Pure water



(b) HF solution

## Table of Some Common Weak Acids

There are many weak acids—some organic and some inorganic. Most carboxylic (organic) acids are weak acids.

- Two of the weak acids shown here are diprotic, and one is triprotic.
- Sulfurous acid and carbonic acid are weak in both of their ionizable hydrogens, and phosphoric acid is weak in all three of its ionizable hydrogens.

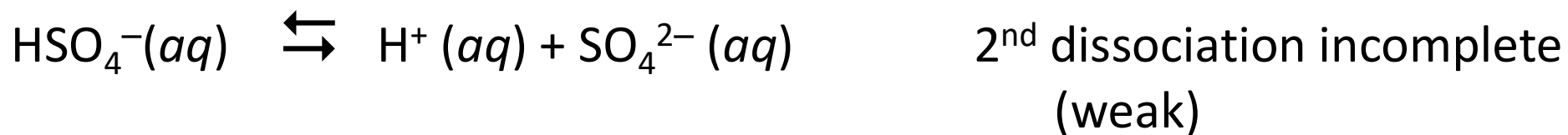
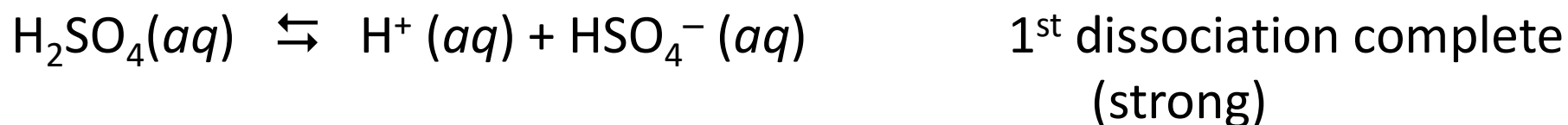
**TABLE 14.4 Weak Acids**

hydrofluoric acid (HF)	sulfurous acid ( $\text{H}_2\text{SO}_3$ ) ( <i>diprotic</i> )
acetic acid ( $\text{HC}_2\text{H}_3\text{O}_2$ )	carbonic acid ( $\text{H}_2\text{CO}_3$ ) ( <i>diprotic</i> )
formic acid ( $\text{HCHO}_2$ )	phosphoric acid ( $\text{H}_3\text{PO}_4$ ) ( <i>triprotic</i> )

## Speaking of weak acids, ...

### They can lurk in a strong acid too

Sulfuric acid (listed with the strong acids) is a diprotic acid that is strong in its first ionizable hydrogen but weak in its second ionizable hydrogen.



- No acid's second and later dissociations are complete
- The anion that still has acidic hydrogens is always a weak acid
  - Even sulfuric acid's, as we see above

## What Is the $\text{H}_3\text{O}^+$ Concentration in Each Solution?

### 1.5 M HCl

Since HCl is a strong acid, it completely ionizes. The concentration of  $\text{H}_3\text{O}^+$  will be 1.5 M.

### 3.0 M $\text{HC}_2\text{H}_3\text{O}_2$

Since  $\text{HC}_2\text{H}_3\text{O}_2$  is a weak acid, it partially ionizes. The calculation of the exact concentration of  $\text{H}_3\text{O}^+$  is beyond the scope of this text, but we know that it will be less than 3.0 M.



## Strong and Weak Acids and Bases: Strong Bases

A **strong base** is one that completely dissociates in solution.

**TABLE 14.5 Strong Bases**

---

lithium hydroxide (LiOH)	strontium hydroxide (Sr(OH) <sub>2</sub> )
sodium hydroxide (NaOH)	calcium hydroxide (Ca(OH) <sub>2</sub> )
potassium hydroxide (KOH)	barium hydroxide (Ba(OH) <sub>2</sub> )

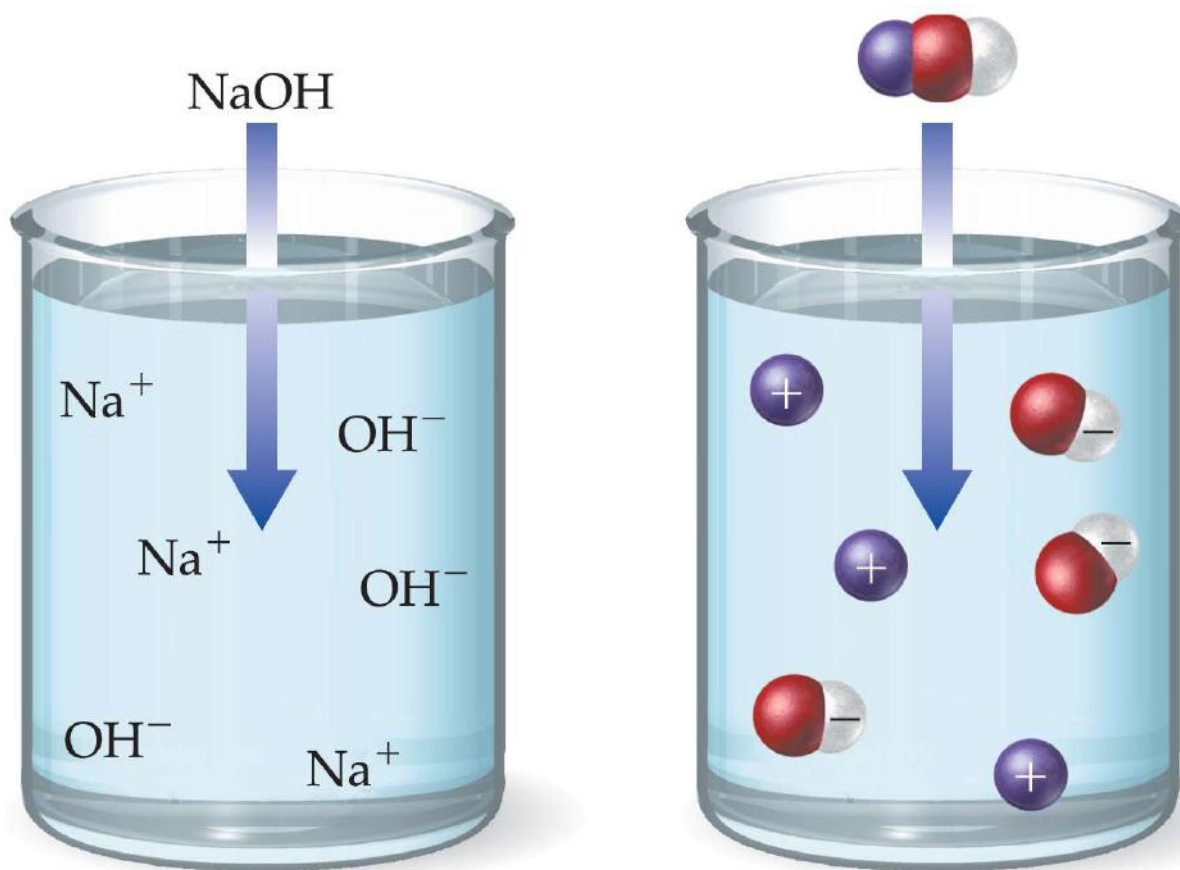
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Unlike diprotic acids, which ionize in two steps, bases containing 2 OH<sup>-</sup> ions dissociate in one step.

Because they are ionic compounds, and ionic compounds dissociate into ions completely

## A Strong Base: NaOH

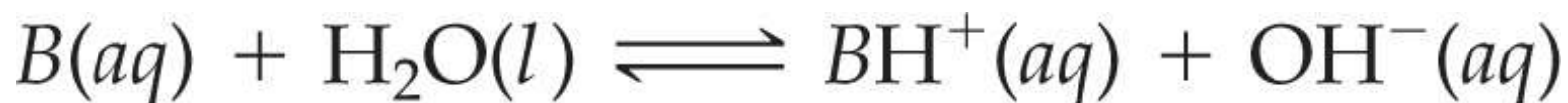
An NaOH solution contains no intact NaOH—it has all dissociated to form  $\text{Na}^+$  and  $\text{OH}^-$ .



## Strong and Weak Acids and Bases: Weak Bases

A **weak base** is analogous to a weak acid. Unlike strong bases that contain  $\text{OH}^-$  and dissociate in water, the most common weak bases produce  $\text{OH}^-$  by accepting a proton from water, causing the water to form  $\text{OH}^-$  ions.

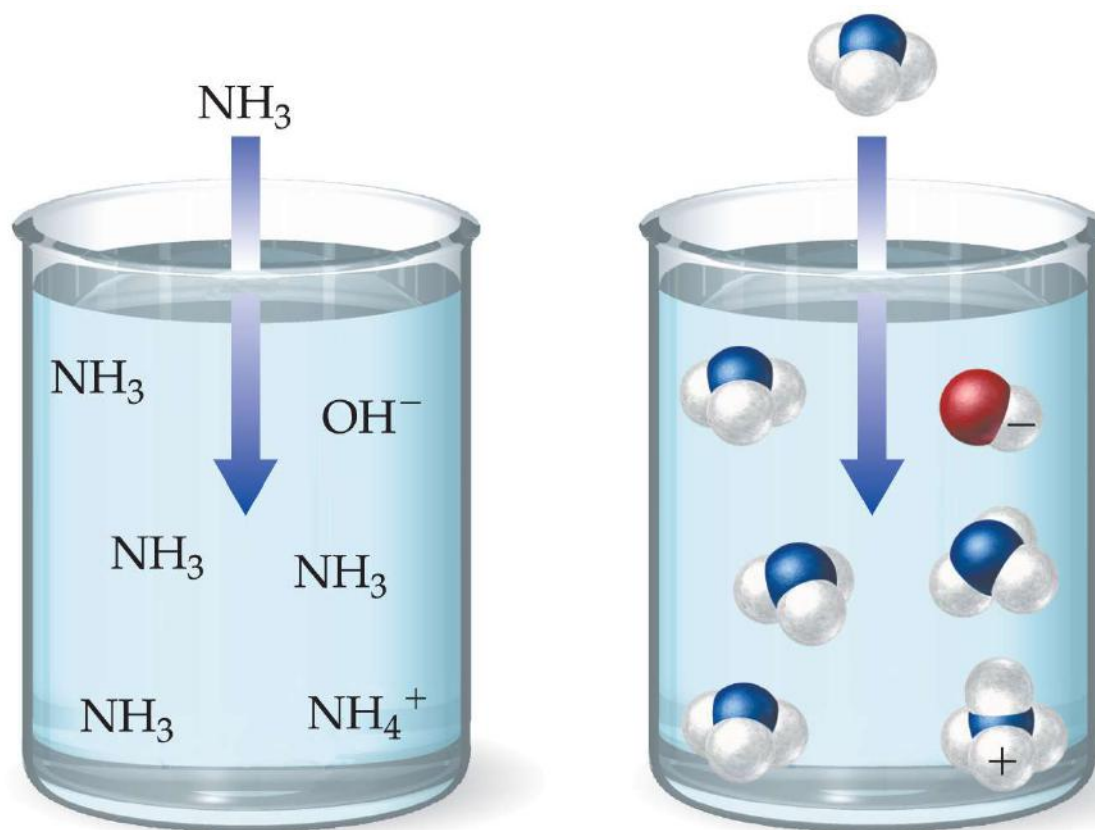
In this equation,  $B$  is generic for a weak base.



In practice only alkali metal hydroxides like NaOH and KOH are soluble enough to make concentrated base solutions.

## A Weak Base: $\text{NH}_3$

When  $\text{NH}_3$  dissolves in water, it partially ionizes to form  $\text{NH}_4^+$  and  $\text{OH}^-$ . However, only a fraction of the molecules ionize. Most molecules remain as  $\text{NH}_3$ .



## Strong and Weak Acids and Bases: Weak Bases

- There are many weak bases.
- Organic amines are weak bases.
- They are essentially ammonia molecules where one of the hydrogens is replaced with an organic (carbon-based) group.

**TABLE 14.6** Some Weak Bases

Base	Ionization Reaction
ammonia (NH <sub>3</sub> )	$\text{NH}_3(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{NH}_4^+(aq) + \text{OH}^-(aq)$
pyridine (C <sub>5</sub> H <sub>5</sub> N)	$\text{C}_5\text{H}_5\text{N}(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_5\text{H}_5\text{NH}^+(aq) + \text{OH}^-(aq)$
methylamine (CH <sub>3</sub> NH <sub>2</sub> )	$\text{CH}_3\text{NH}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{CH}_3\text{NH}_3^+(aq) + \text{OH}^-(aq)$
ethylamine (C <sub>2</sub> H <sub>5</sub> NH <sub>2</sub> )	$\text{C}_2\text{H}_5\text{NH}_2(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{C}_2\text{H}_5\text{NH}_3^+(aq) + \text{OH}^-(aq)$
bicarbonate ion (HCO <sub>3</sub> <sup>-</sup> )*	$\text{HCO}_3^-(aq) + \text{H}_2\text{O}(l) \rightleftharpoons \text{H}_2\text{CO}_3(aq) + \text{OH}^-(aq)$

\*The bicarbonate ion must occur with a positively charged ion such as Na<sup>+</sup> that serves to balance the charge but does not have any part in the ionization reaction. It is the bicarbonate ion that makes sodium bicarbonate (NaHCO<sub>3</sub>) basic.

## What Is the $\text{OH}^-$ Concentration in Each Solution?

### **2.25 M KOH**

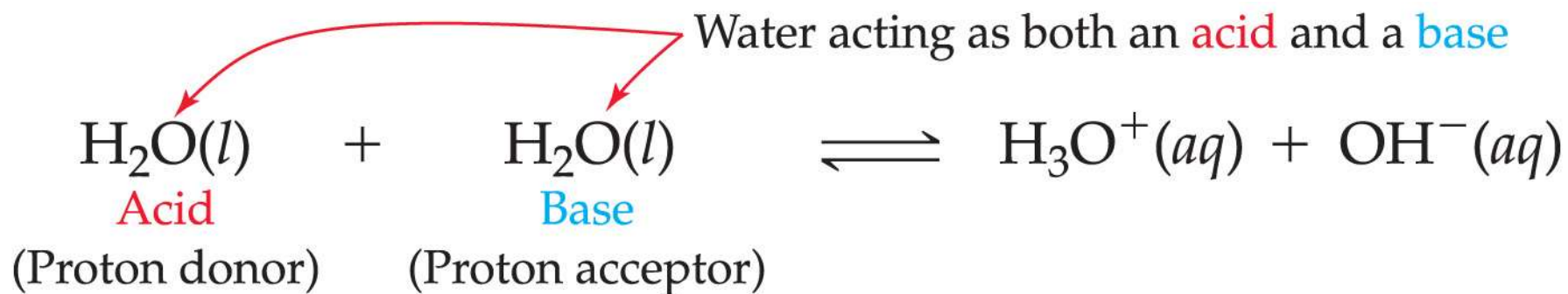
Since KOH is a strong base, it completely dissociates into  $\text{K}^+$  and  $\text{OH}^-$  in solution. The concentration of  $\text{OH}^-$  is 2.25 M.

### **0.35 M $\text{CH}_3\text{NH}_2$ (methylamine)**

Since  $\text{CH}_3\text{NH}_2$  is a weak base, it only partially ionizes water. We cannot calculate the exact concentration of  $\text{OH}^-$  here, but we know it is less than 0.35 M.

## Water: Acid and Base in One

- Water is *amphoteric*; it can act as either an acid or a base.
- Even in pure water, water acts as an acid and a base with itself, a process called self-ionization.





## The Self-Ionization Reaction of Water

Occurs only to a very small extent, resulting in equal and small concentrations of  $\text{H}_3\text{O}^+$  and  $\text{OH}^-$

- For pure water at 25 °C:

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

The ion product constant for water is:

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

$$K_w = 1.0 \times 10^{-14}$$

In a neutral solution,  $[\text{H}_3\text{O}^+] = [\text{OH}^-]$ .

## Using $K_w$ in Calculations

Calculate  $[\text{OH}^-]$  in the solution and determine whether the solution is acidic, basic, or neutral.

GIVEN:

$$[\text{H}_3\text{O}^+] = 7.5 \times 10^{-5} \text{ M}$$

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

SOLUTION:

$$[7.5 \times 10^{-5}][\text{OH}^-] = 1.0 \times 10^{-14}$$

$$[\text{OH}^-] = 1.3 \times 10^{-10} \text{ M}$$

$$[\text{H}_3\text{O}^+] > [\text{OH}^-] \Rightarrow \text{acidic solution}$$

## To Summarize Water Behavior, at 25 °C

- In a neutral solution,

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

- In an acidic solution,

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

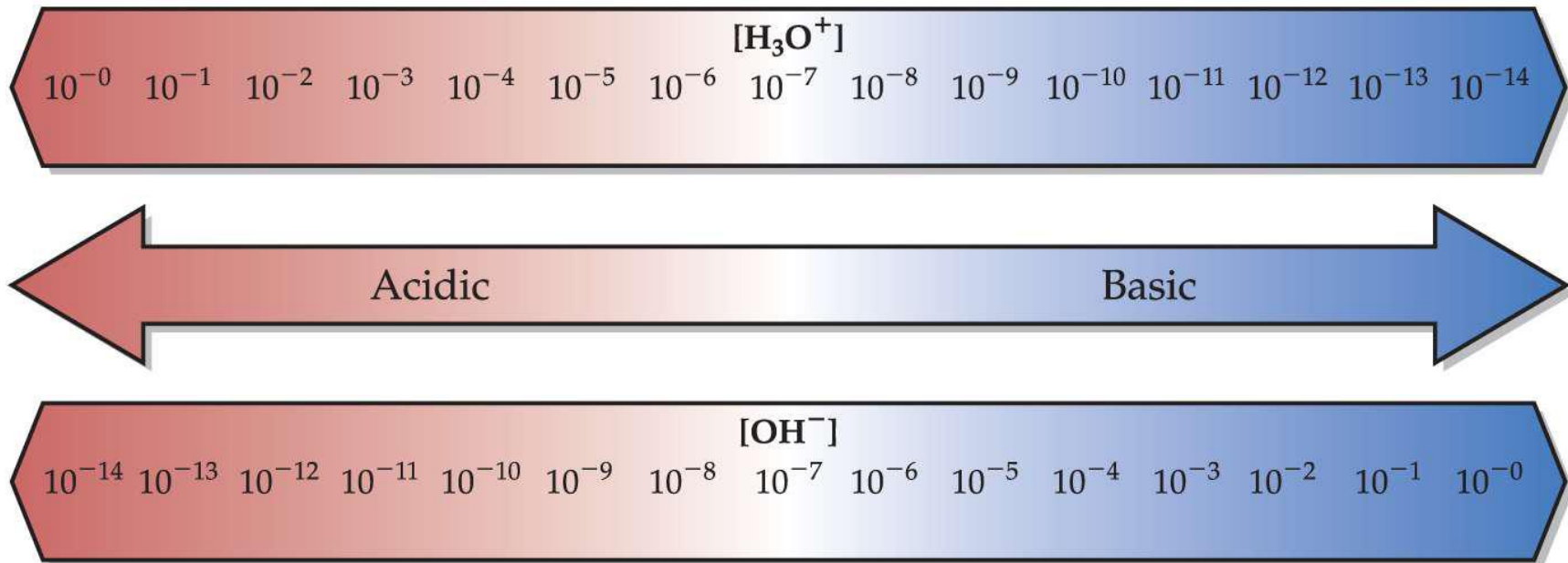
- In a basic solution,

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

- In all aqueous solutions,

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

# Interdependence of $[H_3O^+]$ and $[OH^-]$ in Acidic and Basic Solution



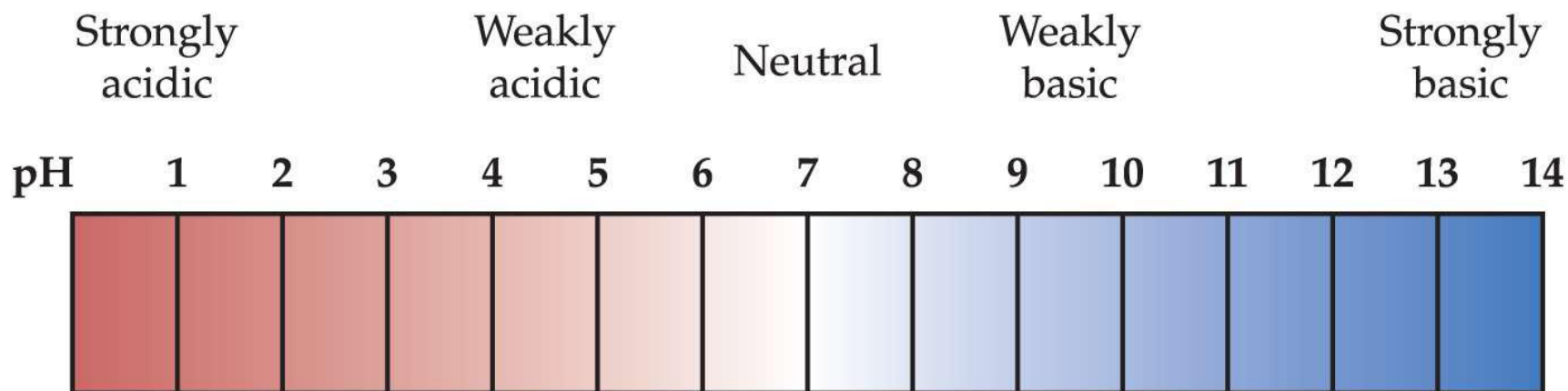
# The pH and pOH Scales: Ways to Express Acidity and Basicity

**At 25 °C:**

pH < 7 ***acidic*** solution



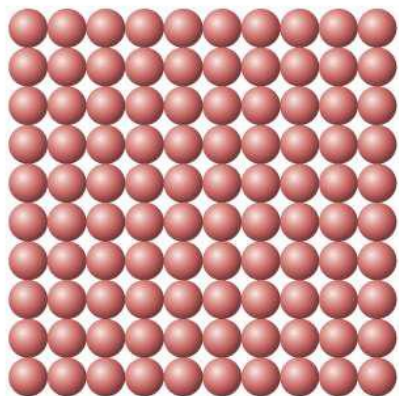
pH > 7 ***basic*** solution

pH = 7 ***neutral*** solution



## The pH Scale Is a Logarithmic Scale

A *decrease* of 1 unit on the pH scale corresponds to an *increase* in  $\text{H}_3\text{O}^+$  concentration by a factor of 10.

pH	$[\text{H}_3\text{O}^+]$	$[\text{H}_3\text{O}^+]$ Representation
4	$10^{-4}$	 ( Each circle represents $\frac{10^{-4} \text{ mol H}^+}{\text{L}}$ )
3	$10^{-3}$	
2	$10^{-2}$	

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

Calculate the pH of the solution and indicate whether the solution is acidic or basic.

GIVEN:

$$[\text{H}_3\text{O}^+] = 1.8 \times 10^{-4} \text{ M}$$

SOLUTION:

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

$$\text{pH} = -\log(1.8 \times 10^{-4})$$

$$\text{pH} = -(-3.74)$$

$$\text{pH} = 3.74$$

$\text{pH} < 7$ , ***acidic*** solution

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

$10^x$  and  $\log x$  are inverse functions that cancel each other

Calculate the  $[\text{H}_3\text{O}^+]$  for a solution with a pH of 4.80.

GIVEN:

$$\text{pH} = 4.80$$

SOLUTION:

$$4.80 = -\log[\text{H}_3\text{O}^+]$$

$$-4.80 = \log[\text{H}_3\text{O}^+]$$

$$10^{-4.80} = 10^{\log[\text{H}_3\text{O}^+]}$$

$$1.6 \times 10^{-5} \text{ M} = [\text{H}_3\text{O}^+]$$



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

Calculate the pOH of the solution and indicate whether the solution is acidic or basic.

GIVEN:

$$[\text{OH}^-] = 1.8 \times 10^{-4} \text{ M}$$

SOLUTION:

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

$$\text{pOH} = -\log(1.8 \times 10^{-4})$$

$$\text{pOH} = -(-3.74)$$

$$\text{pOH} = 3.74$$

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

$$14.00 - 3.74 = \text{pH} = 10.26, \text{ *basic* solution}$$

## Relationship between pH and pOH

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

– at 25 °C

– you can use pOH to find pH of a solution

$$10^{-\text{pOH}} = [\text{OH}^-]$$

Calculate the  $[\text{OH}^-]$  for solution with a pOH of 4.80.

GIVEN:

$$\text{pOH} = 4.80$$

SOLUTION:

$$4.80 = -\log[\text{OH}^-]$$

$$-4.80 = \log[\text{OH}^-]$$

$$10^{-4.80} = 10^{\log[\text{OH}^-]}$$

$$1.6 \times 10^{-5} \text{ M} = [\text{OH}^-]$$

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

$$[\text{H}_3\text{O}^+] = 6.25 \times 10^{-10} < [\text{OH}^-] = 1.6 \times 10^{-5} \Rightarrow \textit{basic}$$

## Practice

Check your answer on the next slide

Calculate the  $[\text{OH}^-]$  at 25 °C when  $[\text{H}_3\text{O}^+] = 1.5 \times 10^{-9} \text{ M}$ , and determine if the solution is acidic, basic, or neutral

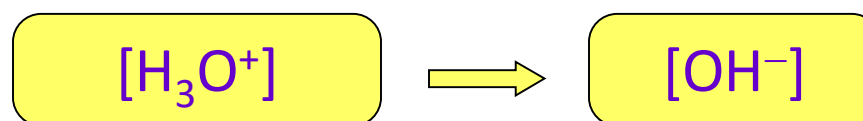
## Practice :

Calculate the  $[\text{OH}^-]$  at  $25^\circ\text{C}$  when  $[\text{H}_3\text{O}^+] = 1.5 \times 10^{-9}\text{ M}$ , and determine if the solution is acidic, basic, or neutral

**Given:**  $[\text{H}_3\text{O}^+] = 1.5 \times 10^{-9}\text{ M}$

**Find:**  $[\text{OH}^-]$

**Conceptual Plan:**



**Relationships:**

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

**Solution:**

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

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

$$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{1.5 \times 10^{-9}} = 6.7 \times 10^{-6}\text{ M}$$

## Practice

Check your answer on the next slide

Determine the  $[\text{H}_3\text{O}^+]$  when  $[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$

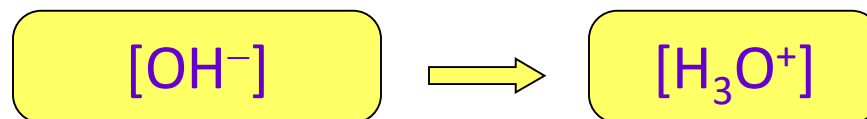
## Practice

Determine the  $[\text{H}_3\text{O}^+]$  when  $[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$

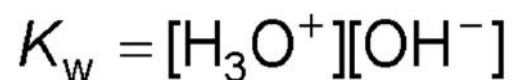
**Given:**  $[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$

**Find:**  $[\text{H}_3\text{O}^+]$

**Conceptual Plan:**



**Relationships:**



**Solution:**

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

$$[\text{H}_3\text{O}^+] = \frac{K_w}{[\text{OH}^-]} \quad [\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-9}} = 4.0 \times 10^{-6} \text{ M}$$

## Practice

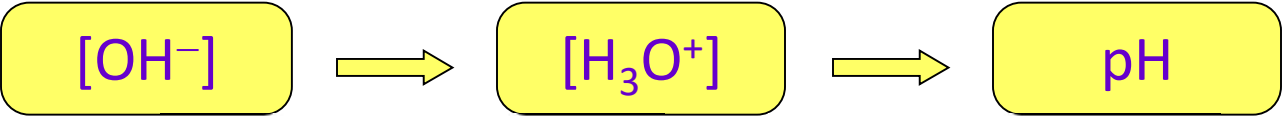
Check your answer on the next slide

Calculate the pH at 25 °C when  $[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$ , and determine if the solution is acidic, basic, or neutral



## Practice:

Calculate the pH at 25 °C when  $[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$ , and determine if the solution is acidic, basic, or neutral

<b>Given:</b>	$[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$
<b>Find:</b>	pH
<b>Conceptual Plan:</b>	
<b>Relationships:</b>	$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$ $\text{pH} = -\log[\text{H}_3\text{O}^+]$

### Solution:

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

$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{1.3 \times 10^{-2}}$$

$$[\text{H}_3\text{O}^+] = 7.7 \times 10^{-13} \text{ M}$$

$$\text{pH} = -\log(7.7 \times 10^{-13})$$

$$\text{pH} = 12.11$$

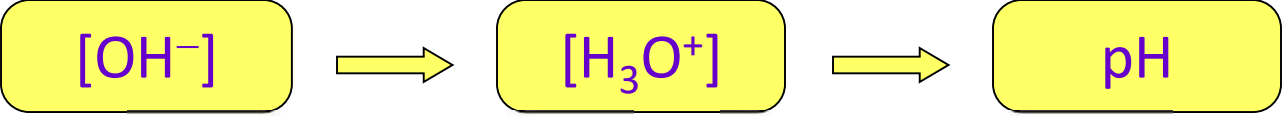
## Practice

Check your answer on the next slide

Determine the pH @ 25 °C of a solution that has  
 $[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$

## Practice:

Determine the pH @ 25 °C of a solution that has  $[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$

<b>Given:</b>	$[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$
<b>Find:</b>	pH
<b>Conceptual Plan:</b>	
<b>Relationships:</b>	$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$ $\text{pH} = -\log[\text{H}_3\text{O}^+]$

### Solution:

$$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$$
$$[\text{H}_3\text{O}^+] = \frac{1.0 \times 10^{-14}}{2.5 \times 10^{-9}}$$

$$[\text{H}_3\text{O}^+] = 4.0 \times 10^{-6} \text{ M}$$
$$\text{pH} = -\log(4.0 \times 10^{-6})$$
$$\text{pH} = 5.40$$

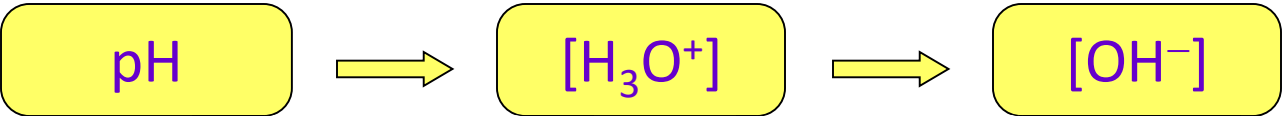
## Practice

Check your answer on the next slide

Determine the  $[\text{OH}^-]$  of a solution with a pH of 5.40

## Practice:

Determine the  $[\text{OH}^-]$  of a solution with a pH of 5.40

<b>Given:</b>	pH = 5.40
<b>Find:</b>	$[\text{OH}^-]$ , M
<b>Conceptual Plan:</b>	
<b>Relationships:</b>	$[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$ $K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$
<b>Solution:</b>	$K_w = [\text{H}_3\text{O}^+][\text{OH}^-]$
	$[\text{H}_3\text{O}^+] = 10^{-5.40}$
	$[\text{H}_3\text{O}^+] = 3.98 \times 10^{-6} \text{ M}$
	$[\text{OH}^-] = \frac{1.0 \times 10^{-14}}{3.98 \times 10^{-6}}$
	$[\text{OH}^-] = 2.5 \times 10^{-9} \text{ M}$

## Practice

Check your answer on the next slide

Calculate the pH at 25 °C when  $[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$ , and determine if the solution is acidic, basic, or neutral

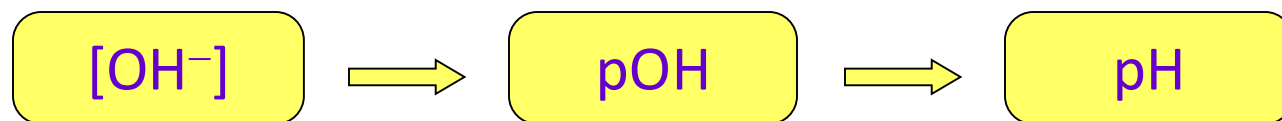
## Practice:

Calculate the pH at 25 °C when  $[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$ , and determine if the solution is acidic, basic, or neutral

**Given:**  $[\text{OH}^-] = 1.3 \times 10^{-2} \text{ M}$

**Find:** pH

**Conceptual Plan:**



**Relationships:**

$$\text{pOH} = -\log[\text{OH}^-] \quad \text{pH} + \text{pOH} = 14.00$$

**Solution:**

$$\text{pOH} = -\log(1.3 \times 10^{-2})$$

$$\text{pOH} = 1.886$$

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

$$\text{pH} = 14.00 - 1.886$$

$$\text{pH} = 12.11$$

## Practice

Check your answer on the next slide

Determine the pOH @ 25 °C of a solution that has  $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-9} \text{ M}$

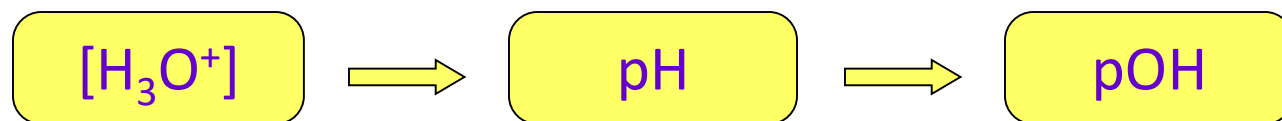


**Practice** – Determine the pOH @ 25 °C of a solution that has  $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-9} \text{ M}$

**Given:**  $[\text{H}_3\text{O}^+] = 2.5 \times 10^{-9} \text{ M}$

**Find:** pOH

**Conceptual Plan:**



**Relationships:**

$$\text{pH} = -\log[\text{H}_3\text{O}^+] \quad \text{pH} + \text{pOH} = 14.00$$

**Solution:**

$$\text{pH} = -\log(2.5 \times 10^{-9})$$

$$\text{pH} = 8.60$$

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

$$\text{pOH} = 14.00 - 8.60$$

$$\text{pOH} = 5.40$$

## Buffers: Solutions That Resist pH Change

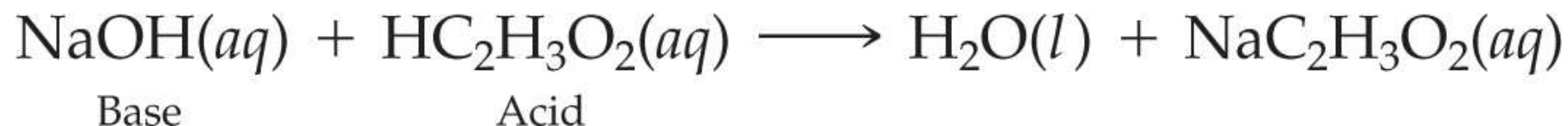
- Buffers contain significant amounts of both a weak acid and its conjugate base.
- The weak acid neutralizes added base.
- The conjugate base neutralizes added acid.

Human blood is a buffer.

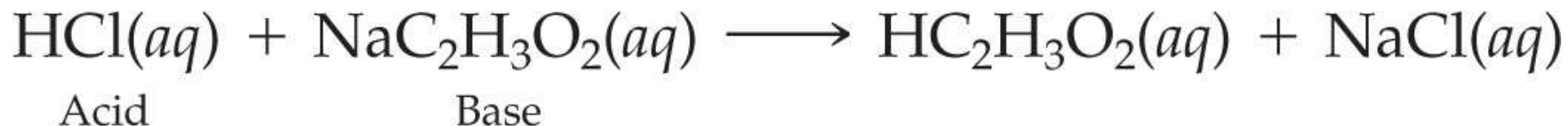
- In healthy individuals, blood pH is between 7.36 and 7.40.
- If blood pH were to drop below 7.0 or rise above 7.8, death would result.

## Buffers: Solutions That Resist pH Change

- In the acetic acid sodium acetate buffer system, with added base, acetic acid reacts:

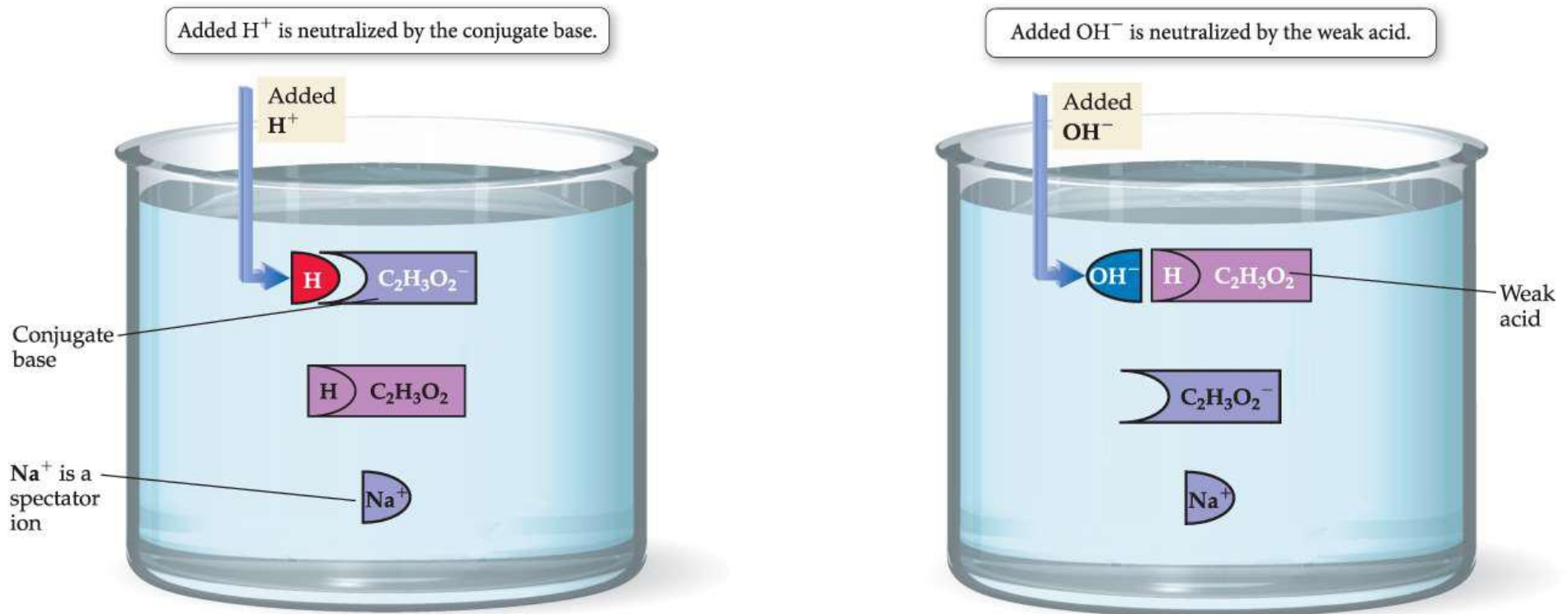


- In the acetic acid sodium acetate buffer system, with added acid, acetate ion (mostly coming from sodium acetate) reacts:



# Buffers: A Buffer Contains Significant Amounts of a Weak Acid and Its Conjugate Base

The acid consumes any added base, and the base consumes any added acid. In this way, a buffer resists pH change.



## Chemistry and Health: The Danger of Antifreeze

- Ethylene glycol is an alcohol, and the first stage of ethylene glycol poisoning is a drunken state.
- Once ethylene glycol begins to metabolize, the second and more deadly stage begins.
- Ethylene glycol is metabolized in the liver into glycolic acid, which enters the bloodstream.
- The glycolic acid overwhelms **the blood's natural buffering system**, causing blood pH to drop to dangerously low levels.
- At this point, the poisoned animal may begin hyperventilating in an effort to overcome the acidic blood's reduced ability to carry oxygen.
- If no treatment is administered, the animal will eventually go into a coma and die.

## Acid Rain

Acid rain is the result of sulfur oxides and nitrogen oxides emitted by fossil fuel combustion.

These oxides react with water to form sulfuric acid and nitric acid, which fall as acid rain.

Acids dissolve metals and metal oxides, so building materials composed of these substances are susceptible to acid rain.

## Chapter 14 in Review

**Acids** have a sour taste; acids dissolve metals.

**Bases** have a bitter taste, a slippery feel.

### Arrhenius Definition:

**Acids** produce  $\text{H}^+$  in solution.

**Bases** produce  $\text{OH}^-$  in solution.

### Brønsted–Lowry Definition:

**Acids** donate a proton.

**Bases** accept a proton.

## Chapter 14 in Review

### Reactions of Acids and Bases:

- Neutralization reactions
- Acid–metal reactions
- Acid–metal oxide reactions

Acid–base titration

Strong and weak acids and bases

Self-ionization of water

pH and pOH scales

Buffers



# Chemical Skills Learning Objectives

1. LO: Identify common acids and bases and describe their key characteristics.
2. LO: Identify Arrhenius acids and bases.
3. LO: Identify Brønsted–Lowry acids and bases and their conjugates.
4. LO: Write equations for neutralization reactions.
5. LO: Write equations for the reactions of acids with metals and with metal oxides.
6. LO: Use acid–base titration to determine the concentration of an unknown solution.
7. LO: Determine  $[\text{H}_3\text{O}^+]$  and  $[\text{OH}^-]$  in solutions.
8. LO: Identify strong and weak acids and strong and weak bases.
9. LO: Find the concentration of  $\text{H}_3\text{O}^+$  or  $\text{OH}^-$  from  $K_w$ .
10. LO: Calculate pH or pOH from  $[\text{H}_3\text{O}^+]$  or  $[\text{OH}^-]$ .
11. LO: Calculate  $[\text{H}_3\text{O}^+]$  or  $[\text{OH}^-]$  from pH or pOH.