

### A unique class of substance... and its complement. Acids, Bases and their reactions.





# Acids

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- Liebig Model-metal can replace hydrogen
  - Sources & History
- Properties
- ► Arrhenius Model-provides H<sup>1+</sup> & OH<sup>1-</sup> ions
  - Hydrogen Ion Electrolytes
    - Binary Acids
    - Oxyacids
      - Polyprotic Acids
  - Arrhenius Bases & Salts
  - Acid Strength
  - Acid-Base Reactions
    - Neutralization Reactions
- Brønsted-Lowry Model-accepts/donates H<sup>1+</sup>
  - Hydronium lons
    - Amphoteric Substances
  - Conjugate Base Pairs
  - Self Ionization of Water

- Equilibrium
  - H<sub>3</sub>O<sup>+</sup> Concentration
    - Ion Product of Water
    - pH & pOH Scale
      - Indicators
      - Calculator Use
    - Titration
  - Gas Evolution Reactions
    - Carbonate Reactions
    - Ammonium Hydroxide
    - Acid-Metal Reactions
  - Disturbing Equilibrium
    - Establishing Equilibrium
    - Le Chatelier's Principle
    - Buffer Solutions









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### Acids & Bases

- There are only five fundamental tastes:
  - Sour, Bitter, Sweet, Savory, and Salty.
- The name "acid" comes from the Latin acidus, meaning "sour taste."
- Acids were first identified by a characteristic sharp odor and sour taste.
  - Vinegar tastes sour because it is a dilute solution of acetic acid in water.
  - Lemon juice is sour because it contains citric acid.
  - Milk turns sour when it spoils because of the formation of lactic acid.
  - The sour odor of rotten meat is due to carboxylic acids such as butyric acid formed when fat spoils.



#### Acids & Bases

- Wood ash and has a bitter taste. Solutions created from wood ash react violently with acids, and destroy the distinctive taste of each.
- Bitter tasting substances that destroy acids became known collectively as alkalies or bases.
- The two complimentary classes of substance, acids and bases, have been used to explore chemical reactions in industry, food preparation, metallurgy, and every other aspect of chemistry for thousands of years.
  - Acids Taste Sour.
  - Bases Taste Bitter.
    - Acids Destroy Bases.
    - Bases Destroy Acids.

### Acids & Bases

- Acids are all around us.
  - ... and are used for diverse purposes.

Name	Occurrence/Uses
Hydrochloric acid (HCl)	Metal cleaning; food preparation; ore refining; primary component of stomach acid
Sulfuric acid (H <sub>2</sub> SO <sub>4</sub> )	Fertilizer and explosives manufacturing; dye and glue production; automobile batteries; electroplating of copper
Nitric acid (HNO <sub>3</sub> )	Fertilizer and explosives manufacturing; dye and glue production
Acetic acid ( $HC_2H_3O_2$ )	Plastic and rubber manufacturing; food preservative; active component of vinegar
Citric acid $(H_3C_6H_5O_7)$	Present in citrus fruits such as lemons and limes; used to adjust pH in foods and beverages
Carbonic acid (H <sub>2</sub> CO <sub>3</sub> )	Found in carbonated beverages due to the reaction of carbon dioxide with water
Hydrofluoric acid (HF)	Metal cleaning; glass frosting and etching
Phosphoric acid (H <sub>3</sub> PO <sub>4</sub> )	Fertilizer manufacture; biological buffering; preservative in beverages



### Litmus

- Litmus is a water soluble mixture of dyes extracted from lichens.
  - It is often absorbed onto filter paper to produce one of the oldest forms of indicator, used to test for acids and bases.
- Litmus was used for the first time about 1300 AD by Spanish alchemist Arnaldus de Villa Nova.
- The practice was refined and expanded by Robert Boyle, who used a syrup from violets.
- Litmus enabled chemists to work out which proportion of acids and bases would neutralize each other, allowing them to determine the relative strengths of different acidic and basic substances.
  - Acids cause litmus to turn red.
  - Bases cause litmus to turn blue.











### Conductivity & Reactivity

- Acids and bases have one property in common, both increase the electrical conductivity of solutions.
- Acids produce hydrogen gas from reaction with active metals.



 Bases produces slippery (soapy) substances from animal tissue.







### Liebig Model of Acids

- Justus von Liebig defined acids as hydrogen containing substances in which the hydrogen can be replaced by a metal.
  - Nitric Acid, Hydrochloric Acid, Acetic Acid are consistent with Liebig's definition of an acid.
  - A substance can be demonstrated to be an acid by mixing it with an active metal.



Justus von Liebig 1803–1873



 $Sn_{(s)} + HNO_{3(aq)} \rightarrow H_{2(g)} \uparrow + Sn(NO_3)_{4(aq)}$ 

 $Mg_{(s)} + HCl_{(aq)} \rightarrow H_{2(g)} \uparrow + MgCl_{2(aq)}$ 

 $Na_{(s)} + HCH_3CO_{2(aq)} \rightarrow H_{2(g)} \uparrow + NaCH_3CO_{2(aq)}$ 



• The Liebig definition of acids lasted for over 50 years.

### Five Properties of Acids & Bases





#### Acids:

- 1. Taste sour.
- 2. Turn litmus red.
- 3. Destroy the properties of bases.
- 4. Conduct electricity.
- 5. Produce  $H_2$  gas from active metals.
  - Bases:
    - 1. Taste bitter (or chalky).
    - 2. Turn litmus blue.
    - 3. Destroy the properties of acids.
    - 4. Conduct electricity.
    - 5. Feel slippery ("soapy").



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    - Polyprotic Acids
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### **Svante Arrhenius**

- Svante August Arrhenius was a prodigy with education in both physics and chemistry.
  - He is sometimes described as the father of modern physical chemistry.
- Svante Arrhenius 1883 doctoral dissertation included 56 thesis, most of which would be considered a sufficient Ph.D. thesis today.
- One of these was the idea of electrolytic dissociation that electrolytes dissociate into ions when placed in water. (At the time, it was thought you needed electric current to form ions.)
  - Arrhenius got a "D" for his dissertation.
  - Upon review the University raised the grade to a C and he was begrudgingly awarded his doctorate degree.



Svante August Arrhenius 1859–1927

 $\begin{array}{ll} AB \rightarrow A^{+} + B^{-} \\ NaCl \rightarrow & Na^{+} + Cl^{-} \\ K_{2}SO_{4} \rightarrow & 2K^{+} + SO_{4}2^{-} \\ \textbf{Electrolyte} & \textbf{Ions} \end{array}$ 



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- The theory of electrolytic dissociation became one of the most widely utilized theories of the next 1,000 years.
  - Arrhenius is recognized as one of the greatest chemists of that millennia.
  - He was awarded the nobel price in 1903 for his theory of electrolytes...
    ... the theory he got a "D" on.



#### **Svante Arrhenius**

 Arrhenius' theory of electrolytic dissociation was that in forming a solution, electrolytes dissociates into charged particles (to which Michael Faraday had given the name ions many years earlier).

$$AB \rightarrow A^{+}(aq) + B^{-}(aq)$$

- Faraday suggested ions were produced by running electric current through the solution.
- Arrhenius proposed that, even in the absence of an electric current, solutions of salts contained ions. He proposed that chemical reactions in solution were reactions between ions.



Svante August Arrhenius 1859–1927



### **Svante Arrhenius**

- In an extension of his theory of electrolytic dissociation Arrhenius proposed definitions for acids and bases.
- Arrhenius offered that...
  - Acids are substances which produce hydrogen ions in solution.

 $\overset{\text{acid}}{\text{HCl}}_{(\text{aq})} \overset{\text{H}_2\text{O}(l)}{\longrightarrow} \qquad \text{H}^+_{(\text{aq})} + \text{Cl}^-_{(\text{aq})}$ 

• Bases are substances that produce hydroxide ions in solution.





Svante August Arrhenius 1859–1927







# Hydrogen Ions

- Hydrogen atoms are made of:
  - 1 proton
  - 0 neutrons
  - 1 electron
    - Hydrogen ions are made by removing that electron.
- A hydrogen ion is a bare proton.
  (there is more to this story when we add water)
- Acids release bare protons, a very small sub-atomic particle with concentrated positive charge.



DANGER Acid

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# **Binary Acids**

- An Arrhenius Acid is a substance that releases H<sup>1+</sup> into solution.
- Binary compounds where one of the two non-metals is hydrogen are not acids. But some become acids when combined with water.
  - Binary compounds only release protons into water solutions.
- The binary compounds that become acids are: HF, HCl, HBr, and HI
- We indicate something is dissolved in water by putting "(aq)" after its formula. Aqueous is Latin for "with water."
- To name a binary acid add "-ic acid" to the anion and prefix the name with "hydro".
  - HBr is hydrogen bromide. It's not an acid.
  - HBr<sub>(aq)</sub> is hydrobromic acid.
  - HCl<sub>(aq)</sub> is hydrochloric acid, a very powerful acid.



Hydrogen Bromide Hydrobromic Acid Hydroiodic Acid

HF HCI (aq) Answer:

Hydrochloria Acid, GR

HBr HBr (aq) HI (aq)

Hydrogen Fluoride Hydrochloric Acid





### **Arrhenius Acids**

- Arrhenius acids are electrolytes where the cation is hydrogen ion.
- Binary acids are acids where the anion is a monatomic ion.
- Acids can also be formed with polyatomic anions, most commonly oxyions.



• Some acid molecules can release more than a single hydrogen ion.





### **Polyprotic Acids**

- Acids can dissociate to release more than one proton.
  - Acids that dissociate to release a single proton are monoprotic.

HCl (aq) 
$$\rightleftharpoons$$
 H<sup>1+</sup> (aq) + Cl<sup>1-</sup> (aq)

• Acids that dissociate to release two protons are diprotic.

$$H_2SO_4 (aq) \rightleftharpoons H^{1+} (aq) + HSO_4^{1-} (aq)$$

$$HSO_4^{1-}$$
 (aq)  $\rightleftharpoons H^{1+}$  (aq) +  $SO_4^{2-}$  (aq)

• Acids that dissociate to release more than two protons are polyprotic.

$$H_{3}PO_{4} (aq) \rightleftharpoons H^{1+} (aq) + H_{2}PO_{4}^{1-} (aq)$$
$$H_{2}PO_{4}^{1-} (aq) \rightleftharpoons H^{1+} (aq) + HPO_{4}^{2-} (aq)$$

 $HPO_4^{2-}$  (aq)  $\rightleftharpoons H^{1+}$  (aq)  $+ PO_4^{3-}$  (aq)



HCI





# Naming Oxyacids

- Polyatomic ions with enough hydrogens on them to neutralize their charge <u>become acids</u>.
- Oxyacids do not need to be in water to be acids, they are acids with or without (aq).

- To name acids of oxy-ions, replace
  - the "-ate ion" with "-ic acid"
  - the "-ite ion" with "-ous acid"

		Dhosphate ion		Dhosphorie acid	ר ר
P	PO <sub>4</sub> <sup>3-</sup>	Phosphate ion	H <sub>3</sub> PO <sub>4</sub>	Phosphoric acid	
	PO <sub>3</sub> <sup>3-</sup>	Phosphite ion	H <sub>3</sub> PO <sub>3</sub>	Phosphorous acid	
S	SO <sub>4</sub> <sup>2-</sup>	Sulf <mark>ate</mark> ion	H <sub>2</sub> SO <sub>4</sub>	Sulfur <mark>ic acid</mark>	ç
	<b>SO</b> <sub>3</sub> <sup>2-</sup>	Sulf <mark>ite</mark> ion	H <sub>2</sub> SO <sub>3</sub>	Sulfur <mark>ous acid</mark>	
С	CO <sub>3</sub> <sup>2-</sup>	Carbon <mark>ate</mark> Ion	H <sub>2</sub> CO <sub>3</sub>	Carbon <mark>ic acid</mark>	
	CO <sub>2</sub> <sup>2-</sup>	Carbon <mark>ite</mark> Ion	H <sub>2</sub> CO <sub>2</sub>	Carbonous acid	
Ν	NO <sub>3</sub> <sup>1-</sup>	Nitr <mark>ate</mark> lon	HNO <sub>3</sub>	Nitric acid	- 500
	NO <sub>2</sub> <sup>1-</sup>	Nitr <mark>ite</mark> lon	HNO <sub>2</sub>	Nitrous acid	3
Cl,Br, I	ClO <sub>4</sub> 1-	Perchlorate Ion	HClO <sub>4</sub>	Perchloric acid	
	ClO <sub>3</sub> 1-	Chlor <mark>ate</mark> Ion	HClO <sub>3</sub>	Chloric acid	- 400
	ClO <sub>2</sub> <sup>1-</sup>	Chlor <mark>ite</mark> Ion	HClO <sub>2</sub>	Chlorous acid	
	ClO <sup>1-</sup>	Hypochlorite Ion	НСІО	Hypochlorous acid	

### **Arrhenius Acids**

- When writing the formula of an ionic solid, the cation always goes first.
- With Acids that cation is hydrogen ion.
  - The formula of an acid always starts with the acidic hydrogen.
  - For larger molecules, not all the hydrogens may be acidic.
    - Look at where in the formula the hydrogens are shown for structural hints.





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#### **Arrhenius Bases**

- Arrhenius bases
  - produce hydroxide ions (OH-) in water
  - taste bitter or chalky
  - are also electrolytes because they produce hydroxide ions (OH<sup>1-</sup>) in water
  - feel soapy and slippery
  - turn litmus paper blue
- Typical Arrhenius bases are named as hydroxides.
  - NaOH sodium hydroxide
  - KOH potassium hydroxide
  - Ba(OH)<sub>2</sub> barium hydroxide
  - Al(OH)<sub>3</sub> aluminum hydroxide







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- Acids are electrolytes.
- Electrolyte strength describes the dissociation process of electrolytes.
  - The stronger they are, the more they break apart.
  - The stronger they are, the more ions they put into solution.
- Acid and base strength is the same process.
  - Strong acids, put more "H<sup>+</sup>" into solution.
  - Stronger bases, put more "OH<sup>1-</sup>" (or other proton accepters) into solution.





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# Acid Strength



mostly an associated pair of ions.

They produce low concentrations of H<sup>+</sup>



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HI

HBr

HClO<sub>4</sub> HC1

 $H_2SO_4$ 

HNO<sub>3</sub>

 $H_3O^+$ 

 $HSO_4^-$ 

H<sub>3</sub>PO<sub>4</sub>

 $HC_2H_3O_2$ 

 $H_2CO_3$ 

 $H_2PO_4^{2-}$ 

 $NH_4^+$ 

HCN

 $HCO_3^{-}$  $HS^{-}$ 

 $H_2O$ 

 $H_2S$ 

HF  $HNO_2$ 

### Acid Strength

- The stronger an acid is at donating H, the weaker the conjugate base is at accepting H.
- Higher oxidation number = stronger oxyacid
- More oxygens in oxyacid = stronger acid
  - $\bullet \quad H_2SO_4 > H_2SO_3$
  - ► HNO<sub>3</sub> > HNO<sub>2</sub>
- Cations are stronger acids than neutral molecules.
- Neutral molecules are stronger acids than anions.
  - $H_3O^+ > H_2O > OH^-;$
  - $NH_4^+ > NH_3 > NH_2^-$
  - Trend in base strength opposite



Strong Acids	
Hydroiodic acid	HI
Hydrobromic acid	HBr
Perchloric acid	HClO <sub>4</sub>
Hydrochloric acid	HC1
Sulfuric acid	$H_2SO_4$
Nitric acid	HNO <sub>3</sub>
Weak Acids	
Hydronium ion	$H_3O^+$
Hydrogen sulfate ion	$\mathrm{HSO_4}^-$
Phosphoric acid	$H_3PO_4$
Hydrofluoric acid	HF
Nitrous acid	HNO <sub>2</sub>
Acetic acid	HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>
Carbonic acid	$H_2CO_3$
Hydrosulfuric acid	$H_2S$
Dihydrogen phosphate	${\rm H_2PO_4^{\ 2-}}$
Ammonium ion	$\mathrm{NH_4}^+$
Hydrocyanic acid	HCN
Bicarbonate ion	$HCO_3^-$
Hydrogen sulfide ion	$HS^{-}$
Water	$H_2O$

. . .



### **Important Dates**



- This week: Lab Checkout & Practical Counts for 30 pts
- Wed, June 22nd: Final Exam <u>7:00-9:00 am</u> you must take the final exam to pass the course.
   In Lecture Room
- After, June 29th: Final Grades will provided to De Anza.



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#### **Acid-Base Reactions**

Neutralization reactions produce heat.

- An acid-base reaction is any reaction between an acid and base.
- A neutralization reaction is an acid-base reaction where the product is water.
  - Arrhenius bases react with acids to form water.
  - Neutralization reactions occur by a double displacement mechanism.

HCl (aq) + NaOH  $(aq) \rightarrow ?$ 

#### $AB + CD \rightarrow AD + CB$

- Water is not an electrolyte.
- Formation of water will drive a double displacement reaction, the same way formation of a precipitate will cause the reaction to occur.

HCl 
$$_{(aq)}$$
 + NaOH  $_{(aq)} \rightleftharpoons$  H<sup>+</sup>  $_{(aq)}$  + Cl<sup>-</sup>  $_{(aq)}$  + Na<sup>+</sup>  $_{(aq)}$  + OH<sup>-</sup>  $_{(aq)}$ 

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### Problems with the Arrhenius Model

- > The Arrhenius model of acids has some problems...
  - It says solvents play no role in acidity.
    - HCl is an acid only in water.
    - HCl dissolved in benzene is not an acid.
      - It will not turn litmus red.
  - It says salts should produce only neutral solutions.
    - AlCl<sub>3</sub> aqueous solution is acidic.
      - Even though there is no hydrogen in aluminum chloride.
  - ▶ It caused Arrhenius to suggest the base of ammonia in water is NH<sub>4</sub>OH.
    - The base is NH<sub>3</sub>.
    - NH<sub>3</sub> is just as basic dissolved in aniline, where no hydroxide is present.
  - H<sup>+</sup> (a bare proton) doesn't last long in water.
    - The lone pairs on water grab it.
    - Even in the strongest aqueous acid solutions the concentration of H<sup>+</sup> is something like 10<sup>-130</sup> M
- The Arrhenius theory of acids and bases was replaced by a theory proposed independently by Johannes Brønsted and Thomas Lowry in 1923.



 $NH_2$ 





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# Hydronium Ion Formation



- Hydrogen Ion is extremely reactive.
  - That bare proton will grab any electrons available to it.
  - It will even react with water.
  - ▶ Hydrogen ions react with water to form hydronium ion (H<sub>3</sub>O<sup>1+</sup>).

 $HCl + H_2O \rightarrow H_3O^{1+}(aq) + Cl^{1-}(aq)$ 



- Solutions of acids are not the same as solutions of salts.
  - Because of the small size and extreme concentration of positive charge, the hydrogen is not floating in solution.
  - It is bound to another molecule, often the solvent.
  - Until it finds a more stable home.

 $NaCl + H_2O \rightarrow Na^{1+}(aq) + Cl^{1-}(aq)$ 



# Hydronium Ion Formation



- That bare proton will grab any electrons available to it.
- It will even react with water.
- Hydrogen ions react with water to form hydronium ion  $(H_3O^{1+})$ .

$$HCl + H_2O \rightarrow H_3O^{1+}(aq) + Cl^{1-}(aq)$$

 Hydrogen Ion will protonate other solvents or solutes, it will attach itself to the best lone pair it can find.

> HCl + NH<sub>3</sub> → NH<sub>4</sub><sup>+</sup>(aq) + Cl<sup>1-</sup>(aq) HCl + HS<sup>1-</sup> → H<sub>2</sub>S<sub>(aq)</sub> + Cl<sup>1-</sup>(aq)

 Hydronium ions themselves are very reactive, they readily deliver that proton to any bases found in solution.

 $H_3O^{1+} + F^{1-}_{(aq)} \rightarrow HF_{(aq)} + H_2O$ 




- The Brønsted-Lowry theory was proposed independently by Johannes Nicolaus Brønsted and Thomas Martin Lowry in 1923.
  - They expanded the definition of acids and bases, focusing on what substances could do with a hydrogen ion.
    - The definition of a base was expanded: any substance which accepts a hydrogen ion – even if that something wasn't hydroxide.
    - The definition of an acid was expanded: anything which could provide a hydrogen ion – even in very small concentration.



Johannes Brønsted Thomas Lowry







- Acids and Bases are the same particle, just with or without the H<sup>+</sup> that get's traded back and forth in solution.
- When you have the ball, you're an acid.
- Before you have the ball, you're a base.



Johannes Brønsted

Thomas Lowry







#### 38

- Acids and Bases are the same particle, just with or without the H<sup>+</sup> that get's traded back and forth in solution.
- When you have the ball, you're an acid.
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Thomas Lowry Johannes Brønsted









- Acids and Bases are the same particle, just with or without the H<sup>+</sup> that get's traded back and forth in solution.
- Amphoteric substances, like water, can play either role.
- Amphoteric substances can be a base or an acid.



Johannes Brønsted Thomas Lowry



# Ch10

## **Acids**

- Liebig Model-metal can replace hydrogen
  - Sources & History
  - Properties
- Arrhenius Model-provides H<sup>1+</sup> & OH<sup>1-</sup> ions
  - Hydrogen Ion Electrolytes
    - **Binary Acids**
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      - Polyprotic Acids
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  - Acid-Base Reactions
    - Neutralization Reactions
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  - Hydronium lons
    - Amphoteric Substances
    - **Conjugate Base Pairs**
  - Self Ionization of Water

 $NH_3 +$ 

 $H_2O \iff$ 

- Equilibrium
  - ► H<sub>3</sub>O<sup>+</sup> Concentration
    - Ion Product of Water
    - pH & pOH Scale
      - Indicators
      - Calculator Use
    - Titration
  - **Gas Evolution Reactions** 
    - Carbonate Reactions
    - Ammonium Hydroxide
    - Acid-Metal Reactions
  - Disturbing Equilibrium
    - Establishing Equilibrium
    - Le Chatelier's Principle

 $OH^{-}$ 

**Buffer Solutions** 

 $NH_4^+$ 









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#### Conjugate Acid-Base Pairs

- An acid-base reaction is two bases trading a shared hydrogen ion between them.
- When one base receives the hydrogen ion, it becomes an acid.
- In any acid-base reaction, there are two conjugate acid-base pairs
  - each pair is related by the loss and gain of H<sup>+1</sup>
    - $\, \bullet \,$  each pair is tossing a ball back and forth, the ball is  $H^{1+}$
  - one pair donates in the forward direction; one pair donates in the reverse direction

#### acid and conjugate base pair 1







#### Conjugate Acid-Base Pairs

HA and A<sup>1-</sup> are one acid base conjugate pair.HA is the acid. A<sup>1-</sup> is the base.





B and BH<sup>1+</sup> are the other acid base conjugate pair.B is the base. BH<sup>1+</sup> is the acid.



#### Find the Acid-Base Pairs.



When you add HCl to water, which of the two is the acid?

What is it's conjugate base?

When you mix ammonia and water, which of the two is the acid?

What is it's conjugate base?





### Identify the Acid

In each of the following equations, identify the Brønsted-Lowry acid and base in the reactants:





## Write the conjugate base for...

	ACID	BASE
ery acid has an nbedded base.	HBr →	$H^{1+} + Br^{1-}$
	HNO <sub>3</sub> →	H <sup>1+</sup> + NO <sub>3</sub> <sup>1-</sup>
	$H_3PO_4 \rightarrow$	$H^{1+} + H_2 PO_4^{1-}$
	$NH_4^{1+} \rightarrow$	$H^{1+} + NH_3$
	$H_2CO_3 \rightarrow$	$H^{1+} + HCO_3^{1-}$
	$H_3O^{1+} \rightarrow$	$H^{1+}$ + $H_2O$ (water

Ev er

## Write the conjugate acid for...

BASE ACID The conjugate acid  $Cl^{1-} + H^{1+} \rightarrow HCl$ is just the base with H<sup>+</sup> added.  $BrO_3^{1-} + H^{1+} \rightarrow HBrO_3$  $HPO_4^{2-} + H^{1+} \rightarrow H_2PO_4^{1-}$  $OH^{1-} + H^{1+} \rightarrow H_2O$  (water)  $NH_3 + H^{1+} \rightarrow NH_4^{1+}$  $H_2O + H^{1+} \rightarrow H_3O^{1+}$  (hydronium ion)

# Ch10

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- Water is amphoteric—it can act as an acid or a base.
- Water reacts with itself.
  - H<sup>1+</sup> is transferred from one H<sub>2</sub>O molecule to another
    - one acts as an acid
    - another acts as a base
  - equilibrium is reached between the conjugate acid-base pairs
- This process is called auto ionization.
- Because of auto ionization, pure water always has a small concentration of both:
  - ▶ hydronium ion ( H<sub>3</sub>O<sup>1+</sup> )
  - hydroxide ion ( OH<sup>1-</sup> )

 $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+(aq) + OH^-(aq)$ 



- Because of auto ionization, pure water always has a small concentration of both:
  - hydronium ion (H<sub>3</sub>O<sup>1+</sup>)
  - hydroxide ion ( OH<sup>1-</sup> )
- Because auto ionization produces both ions at the same time, pure water always has an equal amount of each of the two ions.

 $H_{2}O_{(l)} + H_{2}O_{(l)} \rightleftharpoons H_{3}O^{+}(aq) + OH^{-}(aq)$   $= H_{2}O_{(l)} + H_{2}O_{(l)} \rightleftharpoons H_{2}O_{(l)} \rightleftharpoons H_{2}O_{(l)} \leftrightarrow H_{3}O^{+} = OH^{-}(aq)$   $= H_{1}O_{1} + H_{2}O_{(l)} \Leftrightarrow H_{2}O_{(l)} \leftrightarrow H_{3}O^{+} = OH^{-}(aq)$   $= H_{1}O_{1} + OH^{-}(aq)$   $= H_{1}O_{1} + OH^{-}(aq)$   $= H_{1}O_{1} + OH^{-}(aq)$ 



- Because of auto ionization, pure water always has a small concentration of both:
  - hydronium ion ( $H_3O^{1+}$ )
  - hydroxide ion ( OH<sup>1-</sup> )
- Because auto ionization produces both ions at the same time, pure water always has an equal amount of each of the two ions.
- The association of water occurs much faster than the dissociation of water.
- Because of this water is a very weak acid. It is 99.99999 % associated and only 0.00001 % dissociated.

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





(in pure water)

Concentration  $H_3O^{1+} = [H_3O^{1+}] = 0.0000001 \text{ M} = 1.0 \text{ x } 10^{-7} \text{ M}$ Concentration  $OH^{1-} = [OH^{1-}] = 0.0000001 \text{ M} = 1.0 \text{ x } 10^{-7} \text{ M}$ 

(in pure water)

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- If the solution is *not pure water*, they may not be equal.
  - If they do happen to be equal we say the solution is overall neutral.
  - If the hydronium ion concentration is higher, the overall solution is acidic.
  - If the hydroxide ion concentration is higher, the overall solution is basic.



- The ionization reaction of water connects the concentration of the two ions in these solutions.
  - They consume each other.
    - An excess of  $H_3O^{1+}$  means more  $H_3O^{1+}$ ... which depletes the  $OH^{1-}$  in solution.
    - An excess of  $OH^{1-}$  means more  $OH^{1-}$ ... which depletes the  $H_3O^{1+}$  in solution.



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#### $H_2O_{(l)} + H_2O_{(l)} \rightleftharpoons H_3O^+(aq) + OH^-(aq)$

- To model and explore the relationship between these two connected concentrations we define the product of them.
  - The ion product of water (K<sub>w</sub>) is defined as the sum of the concentration of hydroxide and hydronium ion in solution.





• The ion product of water (K<sub>w</sub>) is defined as the sum of the concentration of hydroxide and hydronium ion in solution.

	on Product $K_{\rm w} = \begin{bmatrix} \mathbf{I} \end{bmatrix}$	$H_3O^+$ OH	-]	
	Type of Solution	[H <sub>3</sub> O <sup>+</sup> ]	[OH <sup>-</sup> ]	К <sub>w</sub> (25 °С)
H <sub>3</sub> O	Neutral	$1.0 \times 10^{-7} \mathrm{M}$	$1.0 \times 10^{-7} \mathrm{M}$	$1.0 \times 10^{-14}$
	Acidic	$1.0 \times 10^{-2} \mathrm{M}$	$1.0 \times 10^{-12} \mathrm{M}$	$1.0 \times 10^{-14}$
	Acidic	$2.5 \times 10^{-5} \mathrm{M}$	$4.0 \times 10^{-10} \mathrm{M}$	$1.0 \times 10^{-14}$
	Basic	$1.0 \times 10^{-8} \mathrm{M}$	$1.0  imes 10^{-6} \mathrm{M}$	$1.0 \times 10^{-14}$
	Basic	$5.0 \times 10^{-11} \mathrm{M}$	$2.0 imes10^{-4}\mathrm{M}$	$1.0 \times 10^{-14}$

- The ion product of water is constant at any given temperature, regardless of whether the solution is acidic or basic.
- ▶ At 25°C K<sub>w</sub> = 1.0 x10<sup>-14</sup>

$$K_{\rm w} = \left[ \mathrm{H}_{3}\mathrm{O}^{+} \right] \left[ \mathrm{OH}^{-} \right] = 1.0 \times 10^{-14}$$



- The ion product can be used to identify the concentration of either hydronium ion or hydroxide ion in solution, if you know the concentration of the other.
- The ion product of water ( $K_w$ ) at 25°C is a constant 1.0 x10<sup>-14</sup>.

$$K_{\rm w} = \left[ \mathrm{H}_{3}\mathrm{O}^{+} \right] \left[ \mathrm{OH}^{-} \right] = 1.0 \times 10^{-14}$$

$$\begin{bmatrix} OH^{-} \end{bmatrix} = \frac{K_{w}}{\begin{bmatrix} H_{3}O^{+} \end{bmatrix}}$$
 and  $\begin{bmatrix} H_{3}O^{+} \end{bmatrix} = \frac{K_{w}}{\begin{bmatrix} OH^{-} \end{bmatrix}}$ 



 $H_3O^{\dagger}$ 

OH<sup>-</sup>

- The ion product can be used to identify the concentration of either hydronium ion or hydroxide ion in solution, if you know the concentration of the other.
- The ion product of water ( $K_w$ ) at 25°C is a constant 1.0 x10<sup>-14</sup>.

$$K_{\rm w} = \left[ \mathrm{H}_{3}\mathrm{O}^{+} \right] \left[ \mathrm{OH}^{-} \right] = 1.0 \times 10^{-14}$$

OH

 $H_3O^{+}$ 

• If a solution has a  $[H_3O^{1+}]$  of 0.0012 M at 25°C, what is it's hydroxide ion concentration?

$$\begin{bmatrix} OH^{-} \end{bmatrix} = \frac{K_{w}}{\begin{bmatrix} H_{3}O^{+} \end{bmatrix}}$$
  

$$Eol^{+} = \frac{K_{w}}{\begin{bmatrix} H_{3}O^{+} \end{bmatrix}} = \frac{I.0 \times 10^{-14}}{0.012} = \begin{bmatrix} 8.3 \times 10^{-13} \text{ M} \end{bmatrix}$$

#### How much acid in Milk of Magnesia?

If milk of magnesia has  $[OH^-]$  of 5.0 × 10<sup>-4</sup> M, what is the  $[H_3O^+]$  of the solution?



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Calculator Use

Ion Product of Water

pH & pOH Scale

Indicators

H<sub>3</sub>O<sup>+</sup> Concentration

Titration

Equilibrium

- Gas Evolution Reactions
  - Carbonate Reactions
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#### What pH means.



#### What pH means.



- pH stands for pouvoir hydrogen (power of hydrogen).
- It's the negative power of the hydrogen ion concentration in a solution.
- pH is a measure of how acidic or basic that solution is.
  - pH is defined as the negative log of the H<sup>+</sup> concentration.

# $[H^+] = 10^{-pH}$

 $pH = -log [H^+]$ 

## The pH Scale

#### The pH of a solution

- is used to indicate the acidity of a solution
- has values that usually range from 0 to 14
- is acidic when the values are less than 7
- is neutral at a value of 7
- is **basic** when the values are greater than 7

Acidic solution	pH < 7.0
Neutral solution	pH = 7.0
Basic solution	pH > 7.0



## pH related to [OH1-]

- As pH increases
  - ▶ [H<sup>+</sup>] Concentration goes up.

$$K_{\rm w}$$
 = [H<sub>3</sub>O<sup>+</sup>] [OH<sup>-</sup>] = 1.0 x 10<sup>-14</sup>

▶ [OH<sup>1-</sup>] Concentration goes down.

#### The relationship between H<sup>+</sup>, OH<sup>-</sup> and pH

con	OH <sup>-</sup> centration(mol/l)	pН	H <sup>+</sup> concentration(m	ol/l)
$\begin{array}{c}1 \ge 10^{-14}\\1 \ge 10^{-13}\\1 \ge 10^{-12}\\1 \ge 10^{-11}\\1 \ge 10^{-10}\\1 \ge 10^{-9}\\1 \ge 10^{-9}\\1 \ge 10^{-9}\\1 \ge 10^{-9}\\1 \ge 10^{-7}\\1 \ge 10^{-7}\\1 \ge 10^{-7}\\1 \ge 10^{-7}\\1 \ge 10^{-3}\\1 \ge 10^{-2}\\1 \ge 10^{-1}\\1 \ge 10^{-7}\end{array}$	0.000000000000000000000000000000000000	0 1 2 3 4 5 6 7 9 10 11 12 13	1         Increasing acidity           0.1         acidity           0.01         Implementation           0.001         Implementation           0.0001         Implementation           0.00001         Implementation           0.000001         Implementation           0.0000001         Implementation           0.000000001         Implementation           0.00000000001         Implementation           0.0000000000001         Implementation           0.0000000000001         Implementation           0.00000000000000000000000000000000000	$\begin{array}{c} 1 \times 100 \\ 1 \times 10^{-1} \\ 1 \times 10^{-2} \\ 1 \times 10^{-2} \\ 1 \times 10^{-3} \\ 1 \times 10^{-3} \\ 1 \times 10^{-4} \\ 1 \times 10^{-5} \\ 1 \times 10^{-5} \\ 1 \times 10^{-8} \\ 1 \times 10^{-10} \\ 1 \times 10^{-10} \\ 1 \times 10^{-11} \\ 1 \times 10^{-12} \\ 1 \times 10^{-13} \\ 1 \times 10^{-14} \end{array}$
1 x 10 1 x 100	0.1 basicity 1	13	0.0000000000000000000000000000000000000	$1 \times 10$ $1 \times 10$



#### Acidic, Basic, or Neutral?



# Ch10

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#### Measuring pH

The pH of a solution is commonly measured using

- pH meter, a device that determines [H<sup>1+</sup>] of a solution by electrical conductivity
- pH paper, an indicator that turns specific colors at specific pH values

The pH of a solution is found by comparing the colors of indicator paper to a chart.



- Some substances change color, depending on the pH of the solution they are in.
- A dye extracted from lichens, called litmus, is one example.
  - Litmus turns red in acid.
    - ▶ pH < 4.5
  - Litmus turns blue in base.
    - ▶ pH > 8.3
- Litmus is one of the oldest and most generally useful acid-base indicators.
- It was discovered in 1300 AD by the Spanish alchemist Arnaldus de Villa Nova.





- Some substances change color, depending on the pH of the solution they are in.
- Phenolphthalein is another example:
  - Is colorless in acid.
    - ▶ pH < 7.0
  - Is light pink in base.
    - ▶ pH > 7.0
  - Is bright pink in stronger base.
    - ▶ pH > 8.3







- Indicators are found in many common plants.
- The most useful indicators have more than one color change.
- Red cabbage for example produces an indicator with many different transitions.





- Universal indicator is a mixture of four indicators that produces a complete spectrum.
- It's components include:
  - methyl red
  - bromothymol blue
  - thymol blue
  - phenolphthalein





pH 1.0 = Cherry Red pH 2.0 = Rose pH 3.0 = Reddish-Orange pH 4.0 = Orange-Red pH 5.0 = Orange pH 6.0 = Yellow pH 7.0 = Yellowish-Green pH 8.0 = Green pH 9.0 = Bluish-Green pH 10.0 = Blue


## pH Paper

• Litmus, Universal Indicator, and other indicators are often soaked into paper strips to make them more convenient to use.



## pH Meters

- A pH meter is a device that directly measures the pH of a solution.
- It usually has a glass electrode plus a calomel reference electrode, or a combination electrode.
- Basic potentiometric pH meters simply measure the voltage between two electrodes and display the result converted into the corresponding pH value.
- Probes need to be kept clean of contamination as far as possible, and not touched by hand.
- Probes are best kept moist with a medium appropriate for the particular probe (distilled water, which can encourage diffusion out of the electrode, is undesirable) when not in use.
- For very precise work the pH meter should be calibrated before each measurement. For normal use calibration should be performed at the beginning of each day.
- Calibration should be performed with at least two standard buffer solutions that span the range of pH values to be measured. For general purposes buffers at pH 4.01 and pH 10.00 are acceptable.



# Ch10

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## pH Calculations

- The pH scale
  - is a logarithmic scale that corresponds to the [H<sub>3</sub>O<sup>+</sup>] of aqueous solutions.
  - ▶ is the negative logarithm (base 10) of the [H<sub>3</sub>O<sup>+</sup>]

### $\mathbf{pH} = -\log \left[\mathbf{H}_{3}\mathbf{O}^{+}\right]$

 To calculate the pH, the negative powers of 10 in the molar concentrations are converted to positive numbers.

If  $[H_3O^+]$  is  $1.0 \times 10^{-2}$  M

 $pH = -log [1.0 \times 10^{-2}] = -(-2.00) = 2.00$ 











 $[H_3O^+] = 10 \ ^{\text{pH}}$  $= 10 \ ^{\text{(8.31)}}$  $= 10 \ ^{\text{-8.31}}$  $= 4.9 \ \times \ 10^{-9}$ 

When you know the pH and wish to calculate the hydronium ion concentration  $([H_3O^+] = 10^{-pH})$ , you must use the *antilog*  $(10^x)$  function. The antilog function on the TI-30Xa is the second function of the **LOG** key. To calculate the  $[H_3O^+]$  of a solution with pH 8.31, you must push the following sequence of keys:







	$[H_3O^+] = 10^{-pH}$
TEXAS INSTRUMENTS THOSE TAX	= 10 - (8.31)
SLUENG MS	= 10 -8.31
LOG DRG DEL LO <sup>4</sup> F==D R==P LOG PRB • • P e <sup>2</sup> Able == de Able == de LN Able DATA STATE CELL STAT LN Able DATA STATE	$= 4.9 \times 10^{-9}$
TT SIN COS TAN $\div$ T SIN COS TAN $\div$ T $\chi^{-1}$ ( ) $\chi$ ( ) $\chi^{-1}$ (	When you know the pH and wish to calculate the hydronium ion concentration ( $[H_3O^+] = 10^{-pH}$ ), you must use the <i>antilog</i> ( $10^x$ ) function. The antilog function on the TI-30XIIS is the second function of the <b>LOG</b> key. To calculate the $[H_3O^+]$ of a solution with pH 8.31, you must push the following sequence of keys:
OF 1 2 3 ALSSLT TIX ANS 0 • (-)	sequence of keys:



 $[H_3O^+] = 10 \ ^{\text{pH}}$  $= 10 \ ^{\text{(8.31)}}$  $= 10 \ ^{\text{-8.31}}$  $= 4.9 \ \text{x} \ 10^{-9}$ 

When you know the pH and wish to calculate the hydronium ion concentration ( $[H_3O^+] = 10^{-pH}$ ), you must use the *antilog* ( $10^x$ ) function. The antilog function on the Sharp model EL-531X is the second function of the **log** key. To calculate the  $[H_3O^+]$  of a solution with pH 8.31, you must push the following sequence of keys:

#### 2nd F) log +/\_ 8 • 3 1 ) =





### Significant Figures

- To determine the number of significant figures in the pH value, consider the following:
  - The number of decimal places in the pH value is the same as the number of significant figures in the coefficient of [H<sub>3</sub>O<sup>+</sup>].
  - The number to the left of the decimal point in the pH value is the power of 10

(this is determined by counting how how far we move the decimal point, it's not part of the significant figures of the pH measaurement)

$$[H_{3}O^{+}] = 1.0 \times 10^{-2} M \qquad pH = 2.00$$
  
Two SFs Two SFs

## What's the pH of Aspirin?

- Aspirin, which is acetylsalicylic acid, was the first nonsteroidal anti-inflammatory drug used to alleviate pain and fever.
- If a solution of aspirin has a [H<sub>3</sub>O<sup>+</sup>] = 1.7 × 10<sup>-3</sup> M
- What is the pH of the solution?

 $pH = -\log [H_3O^+]$  $pH = -\log 1.7 \times 10^{-3}$ 



1.7 EE or Exp +/- 3 (enter) 
$$\log$$
 +/-  
pH = 2.76 9551079 = 2.77  
 $\uparrow \uparrow \uparrow$   
Exact Two SFs

### Find the pH

• Find the pH of a solution with a  $[H_3O^+]$  of 5.6 × 10<sup>-5</sup>.

 $pH = -\log [H_{3}O^{+}]$   $pH = -\log 5.6 \times 10^{-5}$ 5.6 EE or Exp +/- 5 log +/ pH = 4.251811 = 4.25pH = 4.251811 = 4.25



## What's the [H<sup>+</sup>]?

 Determine the [H<sub>3</sub>O<sup>+</sup>] for solutions having each of the following pH values:

pH = 3.0

For pH values that are whole numbers, the  $[H_3O^+]$  can be written  $1 \times 10^{-pH}$ .

$$[H_3O^+] = 1 \times 10^{-3} M$$

$$3.42 +/- 2nd 10^{x}$$
$$= 3.801893963E-04$$
$$= 3.8 \times 10^{-4}$$





#### **Red Wine**

A red wine has a pH of 3.2. What is the: (a) Concentration of  $[H_3O^+]$ ?

(b) Concentration of [OH<sup>1-</sup>]?





# Ch10

## Acids

¦Η

- Liebig Model-metal can replace hydrogen
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  - Properties
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  - Hydrogen Ion Electrolytes
    - Binary Acids
    - Oxyacids
      - Polyprotic Acids
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  - Acid Strength
  - Acid-Base Reactions
    - Neutralization Reactions
- Brønsted-Lowry Model-accepts/donates H<sup>1+</sup>
  - Hydronium lons
    - Amphoteric Substances
  - Conjugate Base Pairs
  - Self Ionization of Water

- ► Equilibrium
  - ► H<sub>3</sub>O<sup>+</sup> Concentration
    - Ion Product of Water
    - ▶ pH & pOH Scale
      - Indicators
      - Calculator Use
      - Titration
  - Gas Evolution Reactions
    - Carbonate Reactions
    - Ammonium Hydroxide
    - Acid-Metal Reactions
  - Disturbing Equilibrium
    - Establishing Equilibrium
    - Le Chatelier's Principle
    - Buffer Solutions













### Titration

### $AB + CD \rightarrow AD + CB$

HOAC (aq) + NaOH (aq)  $\rightarrow$  H<sub>2</sub>O (I) + NaOAC (aq)







### Titration

- Titration is an analytic technique for determining the concentration in one solution by carefully adding a measured quantity of a known solution and observing a clear end point.
- The unknown is called an analyte.
- The standard solution is called a titrant or titrator.
- The end point is the point in the experiment where an indicator suggests the quantities of analyte and titrant are equal.
- The equivalence point is the point where they actually are.
  - With a good chemical indicator, the two should be close, but your equivalence point is almost always reached before you see the end point.
- An indicator is a chemical added to the mixture that changes color close to the equivalence point.
- Finding the end point with a chemical indicator requires some skill.











## Titration

#### $HCI_{(aq)} + NaOH_{(aq)} \rightarrow H_2O_{(l)} + NaCI_{(aq)}$

- Titration is an analytic technique for determining the concentration in one solution by carefully adding a measured quantity of a known solution and observing a clear end point.
- The unknown is called an analyte.
- The standard solution is called a titrant or titrator.
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Final volume

Neutralized

reading

solution

color)

(indicator

has changed

91

## Stoichiometry



## Stoichiometry





# Ch10

## Acids

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- Equilibrium
  - H<sub>3</sub>O<sup>+</sup> Concentration
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    - ▶ pH & pOH Scale
      - Indicators
      - Calculator Use
    - Titration

#### **Gas Evolution Reactions**

- Carbonate Reactions
- Ammonium Hydroxide
- Acid-Metal Reactions
- Disturbing Equilibrium
  - Establishing Equilibrium
  - Le Chatelier's Principle
  - Buffer Solutions















## Irreversible Double Displacement Rxns

• We call the class of reaction, where two electrolytes react in solution, a double displacement reaction.

$$AB + CD \rightarrow AD + CB$$

• If a non-electrolyte is produced it is irreversible ( $\rightarrow$ ), otherwise the reaction can equilibrate ( $\rightleftharpoons$ ).

 $KI_{(s)} + NaNO_{3(s)} \rightleftharpoons KNO_{3(aq)} + NaI_{(aq)} \longleftarrow$  both reactants and products form  $KI_{(s)} + Pb(NO_{3})_{2(s)} \rightarrow PbI_{2(s)\downarrow} + KNO_{3(aq)} \longleftarrow$  reaction is driven entirely to products

The reaction is driven by a net change that is irreversible:

Molecular Equation  $KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)\downarrow} + KNO_{3(aq)}$ 

Complete Ionic Equation  $K^+_{(aq)} + I_{(aq)} + Pb^{2+}_{(aq)} + NO_3^{1-}_{(aq)} \rightarrow PbI_{2(s)\downarrow} + K^+_{(aq)} + NO_3^{1-}_{(aq)}$ 

Net Ionic Equation  $I^{1-}(aq) + Pb^{2+}(aq) \rightarrow PbI_{2}(s)$ 

- If the net reaction is irreversible, the reaction will be driven entirely to products.
- The formation of a non-electrolyte irreversibly

Remove the spectator Ions

### Irreversible Double Displacement Rxns

#### $AB + CD \rightleftharpoons AD + CB$

- If a non-electrolyte does not form as a product, the reaction will equilibrate back and fourth.
- No net reaction will occur.

#### $AB + CD \rightarrow AD + CB$

- If one of the two possible products is a nonelectrolyte, that product will not dissociate back to reactants.
  - The net reaction is irreversible.
- Examples of irreversible reactions are:
  - Precipitation Reactions

 $AB + CD \rightarrow AD_{(s)\downarrow} + CB$ 

Neutralization Reactions

 $AB + CD \rightarrow H_2O + CB$ 

Gas Evolution Reactions

 $AB + CD \rightarrow AD_{(g)\uparrow} + CB$ 



color

Three indicators of a

chemical change.



heat



## **Gas Formation Reactions**

- ▶ Volatile gases like CO<sub>2 (g)</sub>, H<sub>2</sub>S (g) and NH<sub>3 (g)</sub> that form immediately bubble off.
- The gases escape, their formation is irreversible.
- Sometimes the double displacement reaction forms an unstable compound that decomposes into the gases. Example:

 $H_2CO_3_{(aq)} \rightarrow H_2O_{(l)} + CO_2_{(g)} \qquad \qquad NH_4OH_{(aq)} \rightarrow H_2O_{(l)} + NH_3_{(g)}$ 

- ▶ If a double displacement reaction forms CO<sub>2 (g)</sub>, H<sub>2</sub>S (g), or NH<sub>3 (g)</sub> gases this irreversible reaction will drive equilibrium forward.
- If a double displacement reaction forms H<sub>2</sub>CO<sub>3 (aq)</sub> or NH<sub>4</sub>Cl (aq) these decompose to gases and drive equilibrium forward.
- Examples:

HCl 
$$_{(aq)}$$
 + Na<sub>2</sub>S  $_{(aq)} \rightleftharpoons$  H<sup>+</sup>  $_{(aq)}$  + Cl<sup>-</sup>  $_{(aq)}$  + Na<sup>+</sup>  $_{(aq)}$  + S<sup>2-</sup>  $_{(aq)}$ 

$$\begin{aligned} & \text{HCl}_{(aq)} + \text{Na}_2\text{S}_{(aq)} \rightarrow \text{H}_2\text{S}_{(g)\uparrow} + \text{NaCl}_{(aq)} \\ & \text{H}^+_{(aq)} + \text{Cl}^-_{(aq)} + \text{Na}^+_{(aq)} + \text{S}^{2-}_{(aq)} \rightarrow \text{H}_2\text{S}_{(g)\uparrow} + \text{Cl}^-_{(aq)} + \text{Na}^+_{(aq)} \\ & \text{H}^+_{(aq)} + \text{S}^{2-}_{(aq)} \rightarrow \text{H}_2\text{S}_{(g)\uparrow} \end{aligned}$$



## **Gas Formation Reactions**

- Volatile gases like CO<sub>2 (g)</sub>,  $H_2S_{(g)}$  and  $NH_{3 (g)}$  that form immediately bubble off.
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- If a double displacement reaction forms H<sub>2</sub>CO<sub>3 (aq)</sub> or NH<sub>4</sub>Cl (aq) these decompose to gases and drive equilibrium forward.

 $\begin{aligned} H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightleftharpoons H^+_{(aq)} + SO_4^{2^-}_{(aq)} + Na^+_{(aq)} + HCO_3^-_{(aq)} \\ H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightarrow SO_4^{2^-}_{(aq)} + Na^+_{(aq)} + H_2CO_{3(aq)} \end{aligned}$ 

 $H_2SO_{4(aq)} + NaHCO_{3(aq)} \rightarrow Na_2CO_{3(aq)} + H_2CO_{3(aq)}$ 

 $H_{2}SO_{4(aq)} + NaHCO_{3(aq)} \rightarrow Na_{2}SO_{4(aq)} + H_{2}O_{(l)} + CO_{2(g)} \uparrow$   $H^{+}_{(aq)} + SO_{4}^{2^{-}}_{(aq)} + Na^{+}_{(aq)} + HCO_{3}^{-}_{(aq)} \rightarrow SO_{4}^{2^{-}}_{(aq)} + Na^{+}_{(aq)} + H_{2}O_{(l)} + CO_{2(g)} \uparrow$   $H^{+}_{(aq)} + HCO_{3}^{-}_{(aq)} \rightarrow H_{2}O_{(l)} + CO_{2(g)} \uparrow$  (A)



## **Gas Formation Reactions**

- Volatile gases like CO<sub>2 (g)</sub>,  $H_2S_{(g)}$  and  $NH_{3 (g)}$  that form immediately bubble off.
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- If a double displacement reaction forms H<sub>2</sub>CO<sub>3 (aq)</sub> or NH<sub>4</sub>Cl (aq) these decompose to gases and drive equilibrium forward.

 $NaOH_{(aq)} + NH_4Cl_{(aq)} \rightleftharpoons Na^+_{(aq)} + OH^-_{(aq)} + NH_4^+_{(aq)} + Cl^-_{(aq)}$  $NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + NH_4OH_{(aq)}$ 

 $NaOH_{(aq)} + NH_4Cl_{(aq)} \rightarrow NaCl + NH_4OH_{(aq)}$ 

$$\begin{split} & \mathsf{NaOH}_{(aq)} + \mathsf{NH}_4\mathsf{Cl}_{(aq)} \twoheadrightarrow \mathsf{NaCl}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} + \mathsf{NH}_{3(g)} \uparrow \\ & \mathsf{Na}^+_{(aq)} + \mathsf{OH}^-_{(aq)} + \mathsf{NH}_4^+_{(aq)} + \mathsf{Cl}^-_{(aq)} \twoheadrightarrow \mathsf{Na}^+_{(aq)} + \mathsf{Cl}^-_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} + \mathsf{NH}_{3(g)} \uparrow \\ & \mathsf{OH}^-_{(aq)} + \mathsf{NH}_4^+_{(aq)} \twoheadrightarrow \mathsf{H}_2\mathsf{O}_{(l)} + \mathsf{NH}_{3(g)} \uparrow \end{split}$$



## Double Displacement Reactions

- If I mix two electrolytes (AB & CD), I can look at the two possible double displacement products (AD & CB) to predict if a reaction will occur.
- If either of the two products forms irreversibly, a reaction will occur.
  - Irreversible reactions include precipitation formation, neutralization and gas formation.
- For each pair of **possible products** below, did a reaction occur?



# Ch10

## **Acids**

¦Η

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  - Hydrogen Ion Electrolytes
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    - Acid-Metal Reactions
  - **Disturbing Equilibrium** 
    - Establishing Equilibrium
    - Le Chatelier's Principle
    - **Buffer Solutions**















### **Acid-Metal Reactions**

- Another kind of gas evolution reaction, is the oxidation of active metals by acid.
- When the metal is oxidized, hydrogen ions are reduced and hydrogen gas is produced.
- These reactions occur by a single displacement mechanism.
- There is no equilibrium in these reactions, the reaction is irreversible.
- It is driven by difference in oxidation potential of hydrogen and the active metal.

 $A + BC \rightarrow AC + B$   $Zn_{(s)} + HBr_{(aq)} \rightarrow ZnBr_{2(aq)} + H_{2(g)\uparrow}$   $Mn_{(s)} + HNO_{3(aq)} \rightarrow Mn(NO_{3})_{2(aq)} + H_{2(g)\uparrow}$   $Cu_{(s)} + HNO_{3(aq)} \rightarrow N/R$ 



#### Acid-Metal Reactions are also:

Single Displacement Reactions Red-Ox Reactions Gas Evolution Reactions



### **Acid-Metal Reactions**

- Acids will react with any metal more active than hydrogen.
  - The reaction may be rapid and violent, like magnesium and HCl<sub>(aq)</sub>.
  - It might occur more slowly, like iron and HCl<sub>(aq)</sub>.
  - Currency metals resist corrosion, because they are less active than hydrogen.

	<u>k</u>	
	Latter 1	
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and the second second	2	Contraction of the
	Anna	in the second
Summer of the		The second
Contraction of the local distribution of the		
		ALL PROPERTY.

	Metal	Oxidation Reaction							
	Lithium	$Li(s) \longrightarrow$	Li <sup>+</sup> (aq) +	e <sup>—</sup>	$\wedge$				
	Potassium	$K(s) \longrightarrow$	K <sup>+</sup> (aq) +	e 🖌 🖌					
	Barium	$Ba(s) \longrightarrow$	$Ba^{2+}(aq) +$	2e <sup>-</sup>					
	Calcium	$Ca(s) \longrightarrow$	$Ca^{2+}(aq) +$	$2e^{-}$					
	Sodium	$Na(s) \longrightarrow$	$Na^+(aq) +$	e					
	Magnesium	$Mg(s) \longrightarrow$	$Mg^{2+}(aq) +$	2e <sup>-</sup>					
	Aluminum	$Al(s) \longrightarrow$	$Al^{3+}(aq) +$	3e <sup>-</sup>	S				
	Manganese	$Mn(s) \longrightarrow$	$Mn^{2+}(aq) +$	2e <sup>-</sup>	cas				
	Zinc	$Zn(s) \longrightarrow$	$Zn^{2+}(aq) +$	2e <sup>-</sup>	U U				
	Chromium	$Cr(s) \longrightarrow$	$Cr^{3+}(aq) +$	3e <sup>-</sup>	1 ir				
	Iron	$Fe(s) \longrightarrow$	Fe <sup>2+</sup> (aq) +	2e <sup>-</sup>	ioi				
	Cobalt	$Co(s) \longrightarrow$	$Co^{2+}(aq) +$	2e <sup>-</sup>	dat				
	Nickel	$Ni(s) \longrightarrow$	$Ni^{2+}(aq) +$	2e <sup>-</sup>	- Xi				
	Tin	$Sn(s) \longrightarrow$	$Sn^{2+}(aq) +$	2e <sup>-</sup>	of				
	Lead	$Pb(s) \longrightarrow$	$Pb^{2+}(aq) +$	2e <sup>-</sup>	se				
	Hydrogen	$H_2(g) \longrightarrow$	2H <sup>+</sup> (aq) +	2e -	Ea				
	Copper	$Cu(s) \longrightarrow$	$Cu^{2+}(aq) +$	2e <sup>-</sup>					
	Silver	$Ag(s) \longrightarrow$	$Ag^+(aq) +$	e					
	Mercury	$Hg(l) \longrightarrow$	$Hg^{2+}(aq) +$	2e <sup>-</sup>					
	Platinum	$Pt(s) \longrightarrow$	$Pt^{2+}(aq) +$	2e <sup>-</sup>					
	Gold	$Au(s) \longrightarrow$	Au <sup>3+</sup> (aq) +	3e <sup>-</sup>					



## Oxidation & Reduction

- How do we know which metal gives up it's electrons? Check "activity." The more active ion is the one more likely to turn into a cation (give up it's electrons).
- Which reactions will occur?

 $A + BC \rightarrow AC + B$ 1 **H**  $Na_{(s)} + HBr_{(aq)} \rightarrow ?$ 3 Li 11 Na Na more active than H? Yes. 19 **K**  $Na(s) + HBr(aq) \rightarrow H_{2(g)} \uparrow + NaBr(aq)$ 37 Rb 55 Cs 87 Fr  $Fe_{(s)} + HClO_{3(aq)} \rightarrow ?$ Cu more active than H? No.  $Cu_{(s)} + HClO_{3(aq)} \rightarrow N/R$  $Sn_{(s)} + HNO_{3(aq)} \rightarrow ?$ Sn more active than H? Yes.

 $Sn_{(s)} + HNO_{3(aq)} \rightarrow H_{2(g)\uparrow} + Sn(NO_3)_{4(aq)}$ 

2	2A 2												3A 13	4A 14	54 15
	4 <b>Be</b>												5 <b>B</b>	6 C	7 N
1	12 Mg		3B 3	${}^{4\mathrm{B}}_{4}$	5B 5	6B 6	7B 7	8	9 9	10	1B 11	2B 12	13 Al	14 Si	15 P
	20 Ca		21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 <b>Co</b>	28 Ni	29 Cu	30 <b>Zn</b>	31 <b>Ga</b>	32 Ge	33 A
,	38 Sr		39 Y	40 <b>Zr</b>	41 Nb	42 <b>Mo</b>	43 Tc	44 Ru	45 <b>Rh</b>	46 <b>Pd</b>	47 Ag	48 Cd	49 In	50 <b>Sn</b>	51 S1
	56 <b>Ba</b>	/	71 Lu	72 Hf	73 <b>Ta</b>	74 W	75 <b>Re</b>	76 <b>Os</b>	77 Ir	78 Pt	79 Au	80 <b>Hg</b>	81 Tl	82 Pb	83 <b>B</b> :
,	88 <b>Ra</b>		103 Lr	104 <b>Rf</b>	105 <b>Db</b>	106 <b>Sg</b>	107 <b>Bh</b>	108 <b>Hs</b>	109 <b>Mt</b>	110 <b>Ds</b>	111 <b>Rg</b>	112	113	114	11
	Meta	ls		57 La	58 Ce	59 <b>Pr</b>	60 <b>Nd</b>	61 <b>Pm</b>	62 Sm	63 Eu	64 <b>Gd</b>	65 <b>Tb</b>	66 <b>Dy</b>	67 <b>Ho</b>	68 E
	Meta	11	oids	89 Ac	90 <b>Th</b>	91 <b>Pa</b>	92 U	93 Np	94 <b>Pu</b>	95 <b>Am</b>	96 <b>Cm</b>	97 <b>Bk</b>	98 Cf	99 <b>Es</b>	10 <b>F</b> r
	Nonr	n	etals												
											1	-	-20	-	



# Ch10

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### **Reaction Rates**

- How fast a reaction occurs in solution depends on the concentration of reactants.
  - In a concentrated solution, reactants can find each other more quickly and so product get's formed faster.
  - As the reactant is consumed the mixture becomes more dilute in reactants.
  - It becomes harder for reactants to find each other and the rate slows down.
  - If the reaction is irreversible, it will slow but eventually reach completion.







### **Reaction Rates**

If the reaction is reversible, two reactions are occurring concentration of reactant in the same mixture.

**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightarrow \mathsf{F}^-_{(aq)} + \mathsf{H}_3\mathsf{O}^+_{(aq)}$$

**Reverse Reaction** 

$$\mathbf{F}^{-}(aq) + \mathbf{H}_{3}\mathbf{O}^{+}(aq) \rightarrow \mathbf{HF}(aq) + \mathbf{H}_{2}\mathbf{O}(l)$$






#### **Reaction Rates**

 If the reaction is reversible, two reactions are occurring in the same mixture.

**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$

**Reverse Reaction** 

- When you first mix HF in water, the forward reaction has the highest concentration.
- The reverse reaction is starting with zero reactants.
  - But as the forward reaction occurs the concentration of  $F^{1-}_{(aq)}$  and  $H_3O^{1+}$  increases.
  - And the rate of the reverse reaction increases.





 If the reaction is reversible, two reactions are occurring in the same mixture.

#### **Forward Reaction**

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$

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- When you first mix HF in water, the forward reaction has the highest concentration.
- The reverse reaction is starting with zero reactants.
  - But as the forward reaction occurs the concentration of  $F^{1-}_{(aq)}$  and  $H_3O^{1+}$  increases.
  - And the rate of the reverse reaction increases.
- The forward reaction slows.
- The reverse reaction speeds up.
  - Until you hit a point where product is being produced as fast it is being consumed and the concentrations stop changing.





- The forward reaction slows.
- The reverse reaction speeds up.

**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$

**Reverse Reaction** 

 Until you hit a point where product is being produced as fast it is being consumed and the concentrations stop changing.





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**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$

**Reverse Reaction** 

 Until you hit a point where product is being produced as fast it is being consumed and the concentrations stop changing.





#### Fill in the Blanks

- Complete each of the following with
  - equal or not equal
  - faster or slower
  - change or do not change:



- A. Before equilibrium is reached, the concentrations of the reactants and products <u>change</u>. (change or do not change?)
- B. Initially, reactants have a rate of reaction <u>faster</u> than the rate of reaction of the products. (faster or slower?)
- C. At equilibrium, the rate of the forward reaction is <u>equal</u> to the rate of the reverse reaction. (equal or not equal?)
- D. At equilibrium, the concentrations of the reactants and products <u>do not change</u>. (change or do not change?)



- What that equilibrium concentration becomes, depends on the substance.
- Substances that dissociate very quickly will have equilibrium favoring the dissociated state.
  - These are strong electrolytes.
    - Which includes strong acids (like H<sub>2</sub>SO<sub>4</sub>) and strong bases (like NaOH).

$$\operatorname{NaCl}_{(aq)} \rightleftharpoons \operatorname{Na}^{1+}_{(aq)} + \operatorname{Cl}^{1-}_{(aq)}$$



- What that equilibrium concentration becomes, depends on the substance.
- Substances that dissociate slowly will have equilibrium favoring the associated state.
  - These are weak electrolytes.
    - Which includes weak acids (like HOAc) and weak bases (like NaOAc).

 $HOAc_{(aq)} \rightleftharpoons OAc^{1-}_{(aq)} + H^{1+}_{(aq)}$ 



#### $AB + CD \rightleftharpoons AD + CB$

- If a double displacement reaction is reversible, you will come to some final equilibrium concentration of reactants and products.
- Where equilibrium will lie when you mix two electrolytes, depends on the relative electrolytic strength of the reactants and possible products.

$$HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$$



What happens when we add new substances to a mixture at equilibrium?

$$HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$$



## Ch10

#### Acids

- Liebig Model-metal can replace hydrogen
  - Sources & History
  - Properties
- ► Arrhenius Model—provides H<sup>1+</sup> & OH<sup>1-</sup> ions
  - Hydrogen Ion Electrolytes
    - Binary Acids
    - Oxyacids
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  - Arrhenius Bases & Salts
  - Acid Strength
  - Acid-Base Reactions
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- Brønsted-Lowry Model-accepts/donates H<sup>1+</sup>
  - Hydronium lons
    - Amphoteric Substances
  - Conjugate Base Pairs
  - Self Ionization of Water

- Equilibrium
  - H<sub>3</sub>O<sup>+</sup> Concentration
    - Ion Product of Water
    - ▶ pH & pOH Scale
      - Indicators
      - Calculator Use
    - Titration
  - Gas Evolution Reactions
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  - Disturbing Equilibrium
    - Establishing Equilibrium
    - Le Chatelier's Principle
    - Buffer Solutions













- Once you've reached equilibrium, concentrations don't change.
- Both reactions are still happening, but as fast as a substance is consumed — it's reproduced.
- Like two tanks of water, draining into each other at the same rate.



Water at equilibrium



- If you add water to one tank, the equilibrium is stressed.
- One tank has more water pressure.
- It starts draining faster.





- If you add water to one tank, the equilibrium is stressed.
- One tank has more water pressure.
- It starts draining faster.
- Eventually though, the two tanks will get back to equilibrium.
- > It will be a new equilibrium, but the system will return to a balanced state.



- Chemical reactions behave similarly.
- If you stress a system and equilibrium, the rates will change and the concentrations will change.
- But eventually the system will return to equilibrium.
- ... and we can predict how the concentrations will change.

### $HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$



Le Châtelier's principle states that when equilibrium is disturbed, the rates of the forward and reverse reactions change to relieve that stress and reestablish equilibrium.

### $HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$



Le Châtelier's principle predicts how disturbances will effect equilibrium.

 $\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$ 

**Forward Reaction** 

**Reverse Reaction** 

When the populations of Country A and Country B are in equilibrium, the emigration rates between the two countries are equal so the populations stay constant.



Le Châtelier's principle predicts how disturbances will effect equilibrium.

 $\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$ 

**Forward Reaction** 

**Reverse Reaction** 



• Le Châtelier's principle predicts how disturbances will effect equilibrium.

 $HF_{(aq)} + H_2O_{(l)} \rightleftharpoons F^{-}_{(aq)} + H_3O^{+}_{(aq)}$ 

**Forward Reaction** 

**Reverse Reaction** 

The result will be people moving from Country B into Country A faster than people moving from Country A into Country B.

This will continue until a new equilibrium between the populations is established; the new populations will have different numbers of people than the old ones.





#### Applying Le Châtelier's Principle

#### $HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$

If you add NaCI...





Applying Le Châtelier's Principle

#### $HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$

If you add NaOAc...





Applying Le Châtelier's Principle

#### $HOAc_{(aq)} + NaCl_{(aq)} \rightleftharpoons NaOAc_{(aq)} + HCl_{(aq)}$

If you remove HOAc...





• Le Châtelier's principle predicts how disturbances will effect reaction rates.

**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$

#### **Reverse Reaction**

Stress	Change	Shift in the Direction of
Add reactant	Increases forward reaction rate	Products
Remove reactant	Decreases forward reaction rate	Reactants
Add product	Increases reverse reaction rate	Reactants
Remove product	Decreases reverse reaction rate	Products



• Le Châtelier's principle predicts how disturbances will effect reaction rates.

**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$



**Reverse Reaction** 

#### Shift Equilibrium to

Adding  $F^-$  will...ReactantsAdding HF will...ProductsRemoving F- will...ProductsRemoving H\_2O will...Reactants



• Le Châtelier's principle predicts how disturbances will effect reaction rates.

**Forward Reaction** 

$$\mathsf{HF}_{(aq)} + \mathsf{H}_2\mathsf{O}_{(l)} \rightleftharpoons \mathsf{F}_{(aq)} + \mathsf{H}_3\mathsf{O}_{(aq)}^+$$



**Reverse Reaction** 

#### Effect H<sub>3</sub>O<sup>+</sup> concentration

Adding  $F^-$  will...Decrease  $H_3O^+$ Adding HF will...Increase  $H_3O^+$ Removing F- will...Increase  $H_3O^+$ Removing H\_2O will...Decrease  $H_3O^+$ 



## Ch10

#### **Acids**

¦Η

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    - **Binary Acids**
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 $OH^{-}$ 

**Buffer Solutions** 

 $NH_4^+$ 









 $NH_3 +$ 

 $H_2O \iff$ 





#### **Buffer Solutions**

- Solutions at equilibrium can have their equilibrium concentrations changed dramatically by small additions of substances.
  - For acid-base equilibria, this means big changes in pH.
- Many systems, particularly biological systems, can be dramatically effected by small changes in pH.
- Nature has found a way to protect itself from these changes.
- Solutions like blood have a strong resistance to changes in pH.
- A buffer solution is a mixture that has a resistance changes in pH.
- Here's how buffers accomplish that resistance...





#### Acidic Buffer Solutions

- An acidic buffer solution starts with a weak acid like HOAc.
  - > The weak acid dissociates poorly, so most of it stays associated in solution.





#### **Acidic Buffer Solutions**

- An acidic buffer solution starts with a weak acid like HOAc.
  - > The weak acid dissociates poorly, so most of it stays associated in solution.
- The conjugate base of the acid is then added as a salt.
  - This strong electrolyte dissociates completely.
  - The counter-ion (sodium for example) has no effect on the system.
  - This stacks the deck against changes in H<sup>1+</sup>.

$$\begin{array}{ll} \mathsf{HOAC}_{(aq)} & \rightleftharpoons \mathsf{OAc}^{1-}_{(aq)} + \mathsf{H}^{1+}_{(aq)} \\ & \mathsf{NaOAc} \rightarrow \mathsf{OAc}^{1-}_{(aq)} + & \mathsf{Na}^{1+}_{(aq)} \end{array}$$





#### Acidic Buffer Solutions

- The new system has:
  - ▶ A large concentration of HOAc that is a reservoir against removal of H<sup>1+</sup>.
  - A large concentration of OAc<sup>1-</sup> to eat up any extra H<sup>1+</sup> added.
- By Le Chateliers Principle, the solution will have a strong resistance to addition or removal of H+ making any changes to the persistent H+ concentration.
- > This means pH levels are maintained even if you add acid or base.

```
HOAc_{(aq)} + NaOAc \rightleftharpoons OAc^{1-}_{(aq)} + H^{1+}_{(aq)}
```



#### **Basic Buffer Solutions**

- An basic buffer solution starts with a weak base like NH<sub>3</sub>.
  - > The weak base grabs hydrogen poorly, so most of it stays un-associated in solution.

#### $NH_{3(aq)} + H_2O \rightleftharpoons NH_4^{1+}(aq) + OH^{1-}(aq)$



#### **Basic Buffer Solutions**

- An basic buffer solution starts with a weak base like NH<sub>3</sub>.
  - > The weak base grabs hydrogen poorly, so most of it stays un-associated in solution.
- The conjugate acid of the base is then added as a salt.
  - This strong electrolyte dissociates completely.
  - The counter-ion (chloride for example) has no effect on the system.
  - This stacks the deck against changes in OH<sup>1-</sup>.

$$\begin{array}{rcl} \mathsf{NH}_{3(aq)} & + & \mathsf{H}_2\mathsf{O} & \rightleftharpoons & \mathsf{NH}_4^{1+}(aq) + \mathsf{OH}^{1-}(aq) \\ & & \mathsf{NH}_4\mathsf{Cl} & \rightarrow & \mathsf{NH}_4^{1+}(aq) + & \mathsf{Cl}^{1-}(aq) \end{array}$$



#### Basic Buffer Solutions

- The new system has:
  - ▶ A large concentration of NH<sub>3</sub> to regenerate any lost OH<sup>1-</sup>.
  - ▶ A large concentration of NH<sub>4</sub><sup>1+</sup> to eat up any extra OH<sup>1-</sup> added.
- By Le Chateliers Principle, the solution will have a strong resistance to addition or removal of H+ making any changes to the persistent H+ concentration.
- > This means pH levels are maintained even if you add acid or base.

$$NH_{3(aq)} + NH_{4}Cl \stackrel{H_{2}O}{\rightleftharpoons} NH_{4}^{1+}(aq) + OH^{1-}(aq)$$





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# Questions?

