

Oxidation-Reduction Reaction

Oxidation –loss of electrons

Reduction – gain of electrons

Redox reaction

oxidizing agent – substance that causes
oxidation

reducing agent – substance that cause
reduction

Common Oxidizing and Reducing Agents

TABLE 5.3 Common Oxidizing and Reducing Agents

Oxidizing agent	Reaction product	Reducing agent	Reaction product
O_2 (oxygen)	O^{2-} (oxide ion)	H_2 (hydrogen)	H^+ (hydrogen ion) or H combined in H_2O
H_2O_2 (hydrogen peroxide)	$H_2O(l)$	C (carbon) used to reduce metal oxides	CO and CO_2
F_2, Cl_2, Br_2 , or I_2 (halogens)	F^- , Cl^- , Br^- , or I^- (halide ions)	M, metals such as Na, K, Fe, or Al	M^{n+} , metal ions such as Na^+ , K^+ , Fe^{3+} , or Al^{3+}
HNO_3 (nitric acid)	Nitrogen oxides such as NO and NO_2		
$Cr_2O_7^{2-}$ (dichromate ion)	Cr^{3+} (chromium(III) ion), in acid solution		
MnO_4^- (permanganate ion)	Mn^{2+} (manganese(II) ion), in acid solution		

Recognizing Redox Reactions

TABLE 5.4 Recognizing Oxidation-Reduction Reactions

	Oxidation	Reduction
<i>In terms of oxygen</i>	Gain of oxygen	Loss of oxygen
<i>In terms of halogen</i>	Gain of halogen	Loss of halogen
<i>In terms of hydrogen</i>	Loss of hydrogen	Gain of hydrogen
<i>In terms of electrons</i>	Loss of electrons	Gain of electrons
<i>In terms of oxidation numbers</i>	Increase of oxidation number	Decrease of oxidation number

Oxidation States

Rules for Assigning Oxidation States

1. zero for uncombined element
2. charge on monatomic ion
3. F is always -1 ; other halogens -1 except when combined with more electronegative halogen or oxygen

Oxidation States

Rules for Assigning Oxidation States

4. H is +1 except in metal hydrides, where H is -1
5. O is -2 except when combined with F (then +1 or +2) or in peroxides, -1.
6. sum of oxidation states equals charge on ion or molecule

Oxidation State

What is the oxidation state of S in H_2SO_4 ?

- $\text{H} \rightarrow +1$
- $\text{O} \rightarrow -2$
- neutral compound, thus sum equals zero
- $4\text{O} \rightarrow 4 \times -2 = -8$
- $2\text{H} \rightarrow 2 \times +1 = +2$
- $0 = +2 + (\text{x}) + (-8)$

$$\text{x} = +6$$

Acid & Base

acid

- substance that donates H^+ ions to solution
- sour-tasting substances
- substances whose aqueous solutions are capable of turning blue litmus indicators red
- react with bases or alkalis to form salts
- dissolves certain metals to form salts

base

- substance that donates a OH^{-1} ion to solution
- hydroxides and oxides of metals
- bitter tasting, slippery solutions
- turn litmus blue
- react with acids to form salts

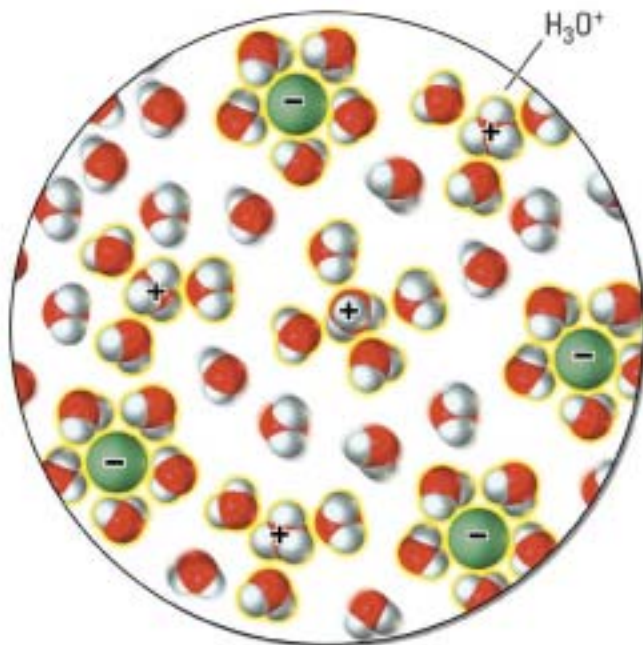
salt

- substances produced by the reaction of an acid with a base
- characterized by ionic bonds, relatively high melting points, electrical conductivity when melted or when in solution, and a crystalline structure when in the solid state

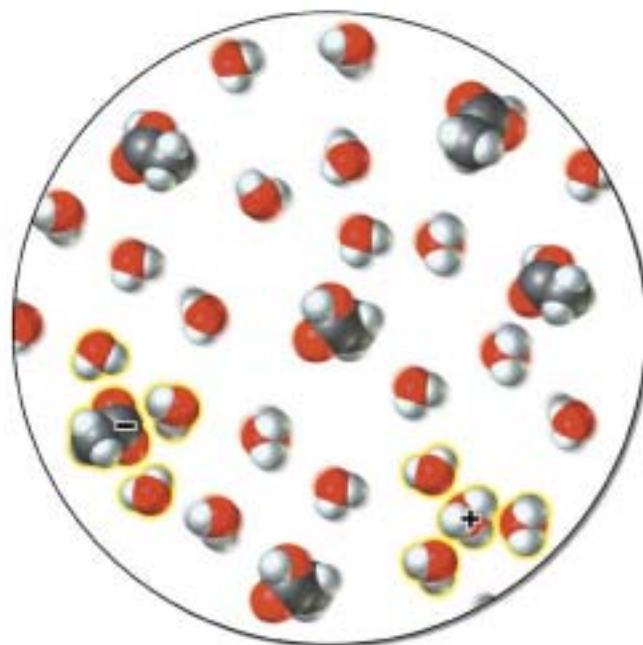
Strong vs. Weak Acids and Bases

- strong – completely ionized
- weak – partially ionized

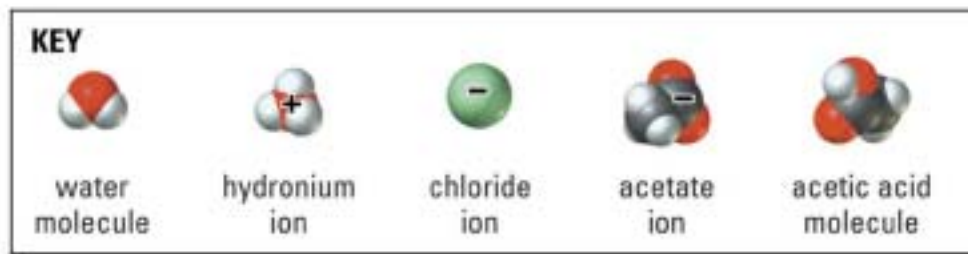
Ionization of Acids in Water



(a) Strong acid (HCl)



(b) Weak acid (CH_3COOH)



Common Acids and Bases

TABLE 5.2 Common Acids and Bases

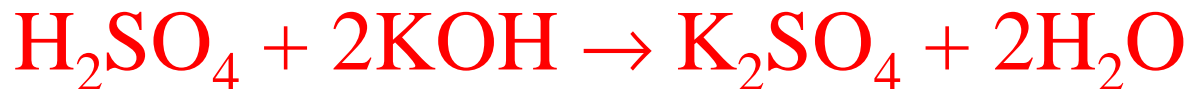
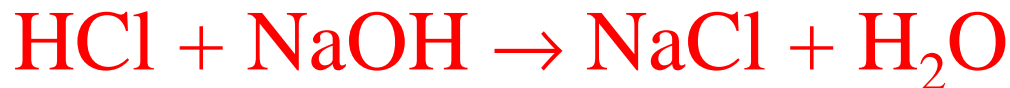
Strong acids (strong electrolytes)		Strong bases (strong electrolytes)	
HCl	Hydrochloric acid	LiOH	Lithium hydroxide
HNO ₃	Nitric acid	NaOH	Sodium hydroxide
H ₂ SO ₄	Sulfuric acid	KOH	Potassium hydroxide
HClO ₄	Perchloric acid	Ca(OH) ₂	Calcium hydroxide
HBr	Hydrobromic acid	Ba(OH) ₂	Barium hydroxide
HI	Hydroiodic acid	Sr(OH) ₂	Strontium hydroxide
Weak acids* (weak electrolytes)		Weak bases† (weak electrolytes)	
H ₃ PO ₄	Phosphoric acid	NH ₃	Ammonia
CH ₃ COOH	Acetic acid	CH ₃ NH ₂	Methylamine
H ₂ CO ₃	Carbonic acid		
HCN	Hydrocyanic acid		
HCOOH	Formic acid		
C ₆ H ₅ COOH	Benzoic acid		

* Many organic carboxylic acids are weak acids.

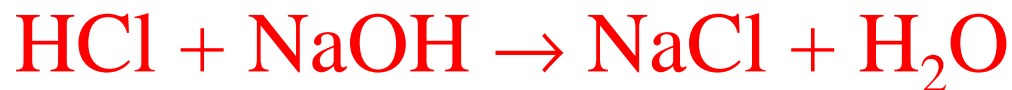
† Many organic amines (related to ammonia) are weak bases.

Neutralization Reactions

acid + base \rightarrow “salt” + water



Ionic Equations



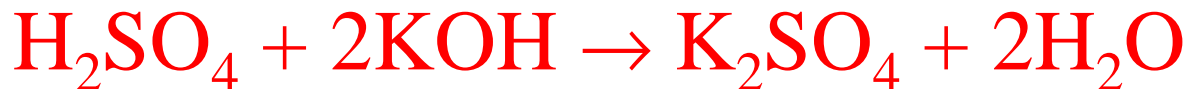
Total Ionic Equation:



Net Ionic Equation:



Ionic Equations



Total Ionic Equation:



Net Ionic Equation:



Solution

homogeneous mixtures of two or more substances.

Solute

- substance that is present in smallest quantity
- dissolved substance(s)
- can be either a gas, a liquid, or a solid
- one or more present in a solution

Solvent

- substance present in largest quantity
- water in aqueous solutions

Molarity

The number of moles of solute per liter of solution.

molarity \rightarrow M

$$M = \frac{\text{moles of solute}}{\text{liter of solution}}$$

units \rightarrow molar = moles/liter = M

Solution Preparation from Solid

- 1 Combine ~950 mL of distilled H_2O with 1.58 g (0.0100 mol) of KMnO_4 in a volumetric flask.



- 2 Shake the flask to dissolve the KMnO_4 .



- 3 After the solid dissolves, add sufficient water to fill the flask to the mark etched in the neck, indicating a volume of 1.00 L.



- 4 Shake the flask again to thoroughly mix its contents. The flask now contains 1.00 L of 0.0100 M KMnO_4 solution.

Titration

1 A buret, a volumetric measuring device calibrated in divisions of 0.1 mL, holds an aqueous solution of a base of known concentration.

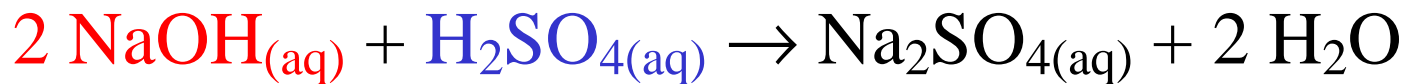
2 Add base slowly from the buret to the solution being titrated.



3 A change in the color of an indicator signals the equivalence point. (The indicator used here is phenolphthalein.)



EXAMPLE: A sample of lye, sodium hydroxide, is neutralized by sulfuric acid. How many milliliters of $0.200\text{ M H}_2\text{SO}_4$ are needed to react completely with 25.0 mL of 0.400 M NaOH ?



$$\text{\#mL H}_2\text{SO}_4 = \frac{(25.0\text{ mL NaOH})}{(1\text{ L NaOH})} \frac{(0.400\text{ mol NaOH})}{(1\text{ L NaOH})} \frac{(1\text{ L})}{(1000\text{ mL})}$$

$$\frac{(1\text{ mol H}_2\text{SO}_4)}{(2\text{ mol NaOH})} \frac{(1000\text{ mL H}_2\text{SO}_4)}{(0.200\text{ mol H}_2\text{SO}_4)}$$

$$= 25.0\text{ mL H}_2\text{SO}_4$$