Limiting reactants

How Many Cookies Can I Make?

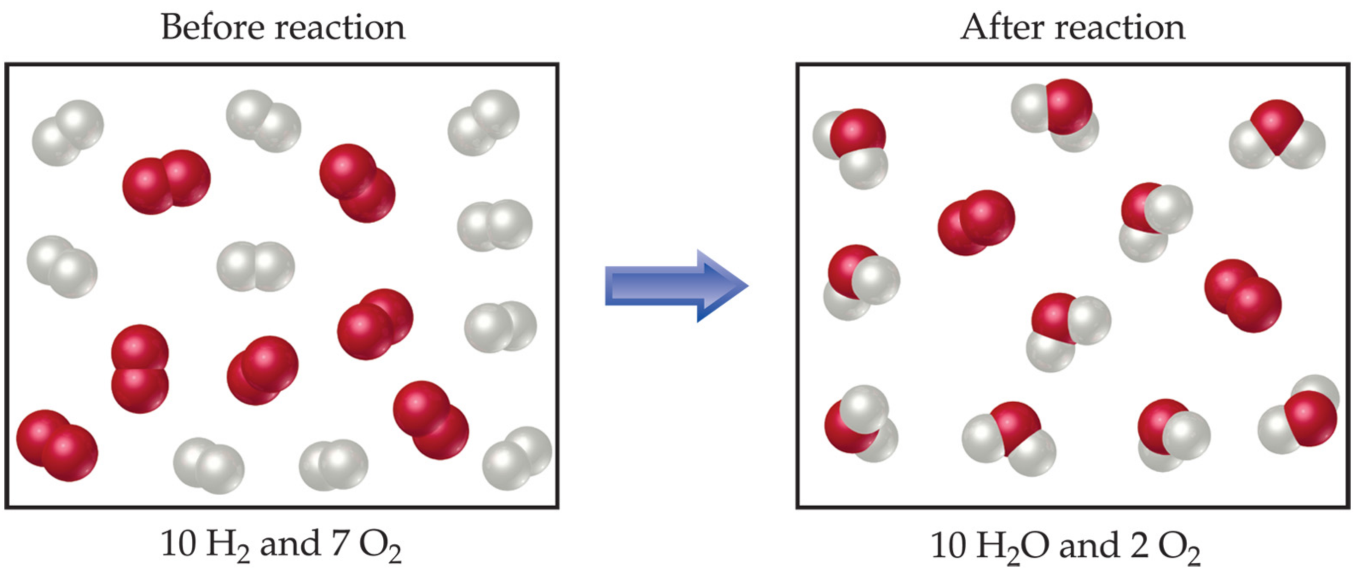
- You can make cookies until you run out of one of the ingredients
- Once you run out of sugar, you will stop making cookies

How Many Cookies Can I Make?

- In this example the sugar would be the limiting reactant, because it will limit the amount of cookies you can make

Limiting Reactants

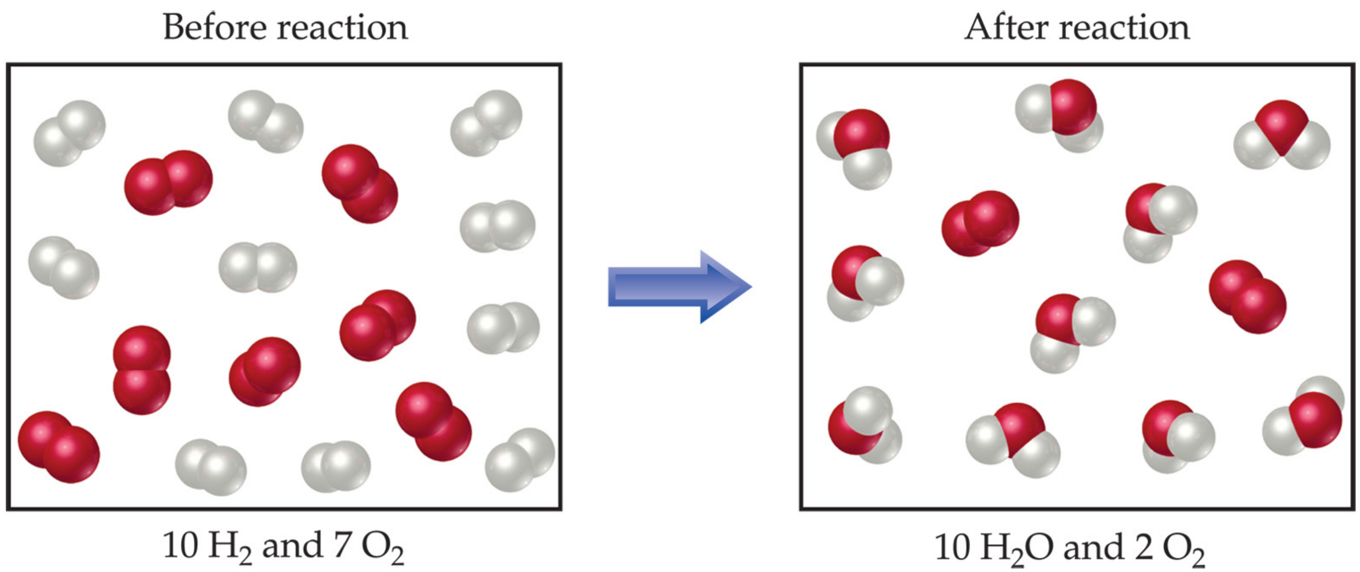
- The limiting reactant is the reactant present in the smallest **stoichiometric** amount



	$2H_2$	+	O_2	-------->	$2H_2O$
#moles	~~14~~		7		
	10		5		10
Left:	0		2		10

Limiting Reactants

In the example below, the O_2 would be the excess reagent



Limiting reagents

- Reaction of nitrous oxide with oxygen to produce nitrogen dioxide
- $2N_2O + 3O_2 \rightarrow 4NO_2$
- The mole ratio is 2 3 4
- OR 1 3/2 4/2
- If three moles of N_2O are mixed with four moles of O_2, what is the maximum amount of NO_2 that can be produced?

- $2N_2O + 3O_2 \rightarrow 4NO_2$
- 3 moles 4 moles ?? Moles
- If all N_2O used: 3 moles 3/2(3 moles)
- But, don't have 4.5 moles O_2 = 4.5 moles
- So: O_2 limiting: 2/3(4) 4 4/3(4)=16/3
- = 8/3

Limiting reagents
A quick way to tell:

Divide the number of moles you have of each reactant by the reaction coefficient for that reactant:

- $2N_2O$ + $3O_2$ → $4NO_2$
- 3 moles 4 moles ???
- 3/2=1.5 4/3=1.33
- So O_2 limiting because 1.33<1.5

Limiting reagent examples

- Suppose 6.54 g of zinc is treated with 5.47 g of hydrochloric acid (in solution). What is the maximum amount of H_2 gas that can be produced and what quantity of the non-limiting reactant remains at the end?

• Equation:	$Zn(s)$ +	$2HCl(aq)$	$\longrightarrow$	$ZnCl_2(aq)$ +	H_2
•	6.54 g	5.47 g			
• Moles:	6.54/65.4	5.47/36.5			
•	=0.1	=0.15			
If Zn:	0.1	2(0.1=.2)			
So HCl:	0.15/2=.075	0.15		0.15/2=.75	0.15/2=0.075

Hydrazine N_2H_4 reacts with dinitrogen tetroxide N_2O_4 by this equation:

$2\ N_2H_4 + N_2O_4 \rightarrow 3\ N_2 + 4\ H_2O$

When 3 mol N_2H_4 reacts with 2 mol N_2O_4, how many moles of N_2 are produced?

$2N_2H_4$	+	N_2O_4	$\rightarrow$	$3N_2$	+	$4H_2O$
3 mole		2 mole		??		
If N_2H_4: 3		3/2=1.5		3(3/2)=9/2		

Percent yield:

- The world is not perfect. When a reaction happens not all of the reactants get turned perfectly into products. You always lose some
- %yield = actual amount/theorectical amount
- Actual: what you actually got
- Theorectical: What you calculated you were going to get.

Percent yield example:

$6UO_3$ +	$8BrF_3$ →	$6UF_4$ +	$4Br_2$ +	$9O_2$
286 g mol^{-1}	137 g mol^{-1}	314 g mol^{-1}		
357 g	357 g			
1.25 mol	2.61 mol			
1.25	1.25(8/6)=1.67	1.25(6/6)=1.25=392.5g		
		380g/392.5gx100=96.8%		

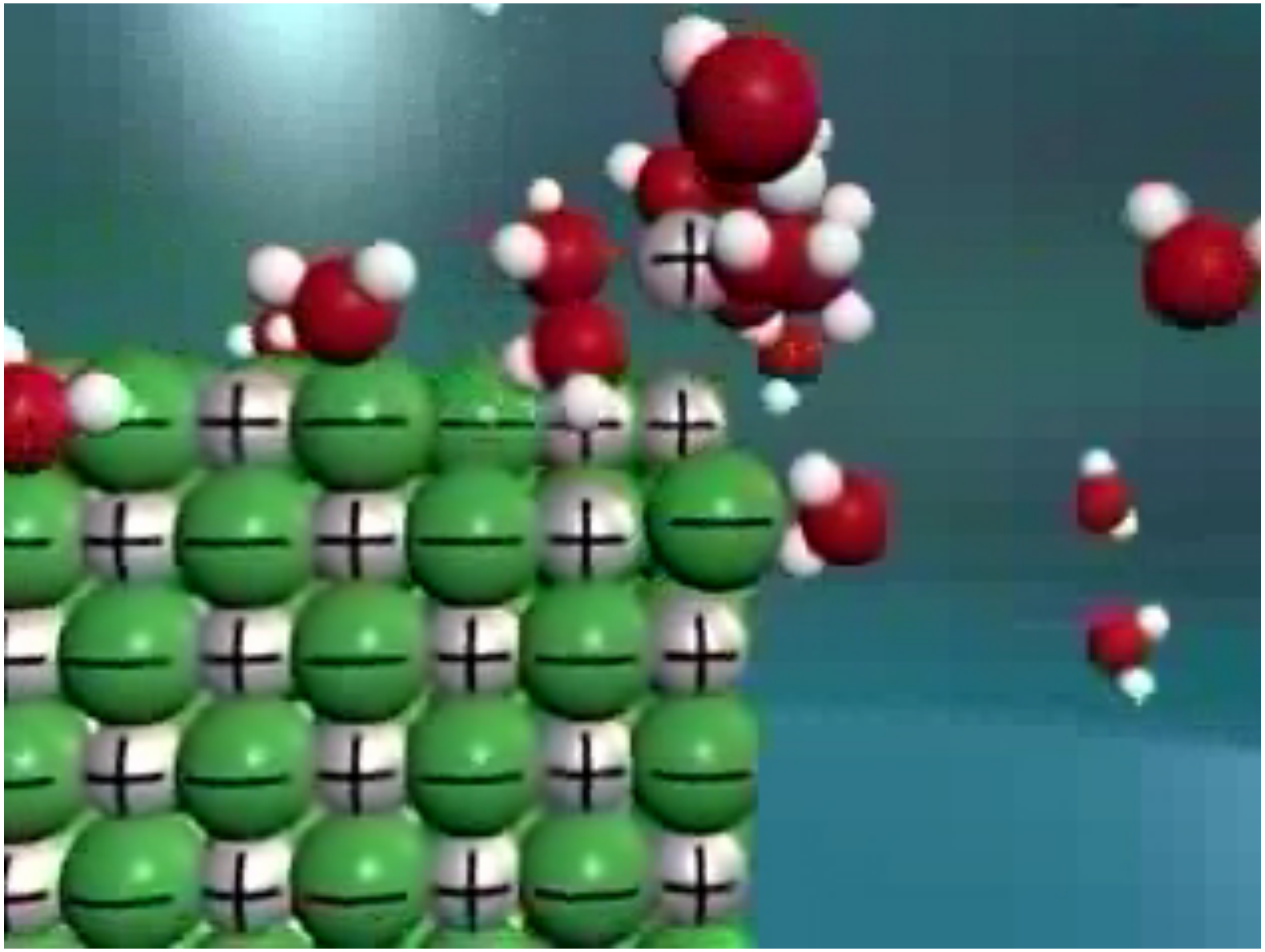
If 380 g of UF_4 was produced, what's the %yield?

Chapter 4
Aqueous Reactions and Solution Stoichiometry

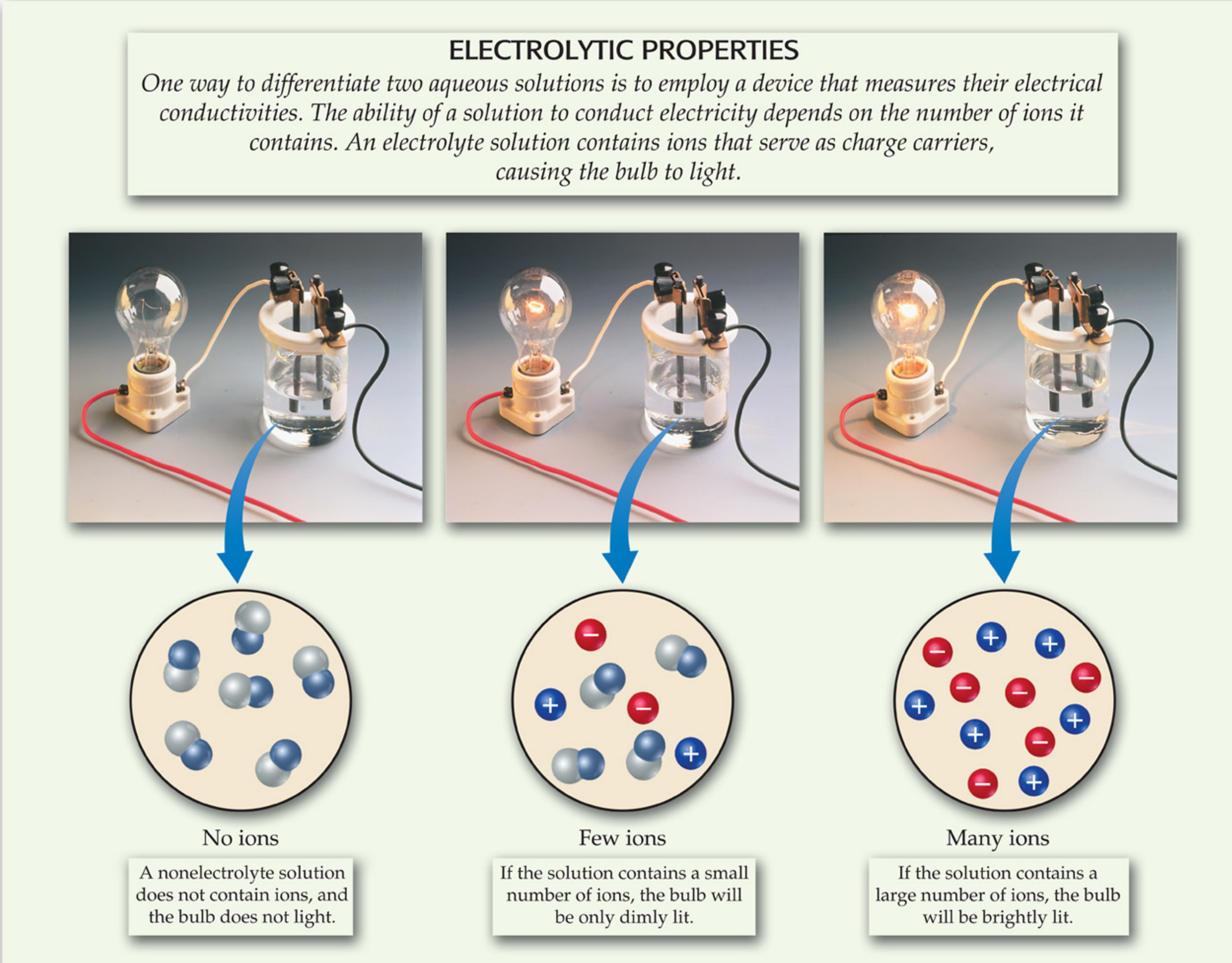
Solutions:

- Homogeneous mixtures of two or more pure substances.
- The solvent is usually present in greatest abundance.
- Or, the solvent is the liquid when a solid is dissolved
- All other substances are solutes.

Dissociation

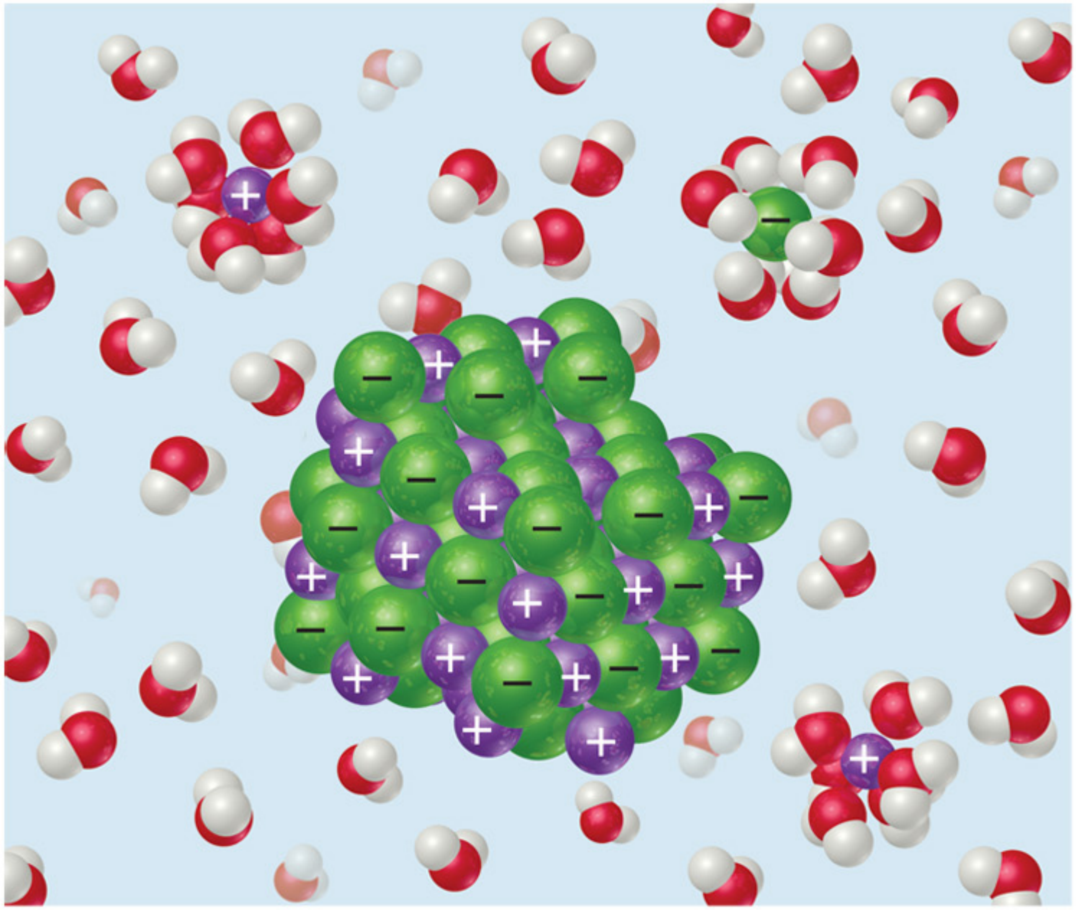


- ionic compound dissolves in water, the individual ions from the crystal are separated. This process is called dissociation.



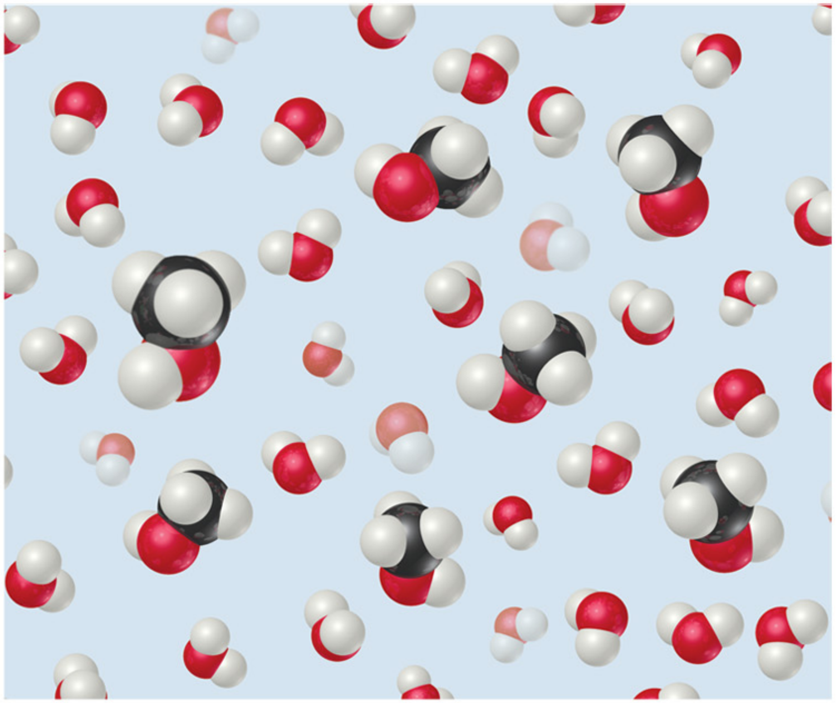
- Substances that dissociate into ions when dissolved in water are electrolytes.
- A nonelectrolyte may dissolve in water, but it does not dissociate into ions when it does so.

Electrolytes and Nonelectrolytes



Soluble ionic compounds tend to be electrolytes.

Electrolytes and Nonelectrolytes



Molecular compounds tend to be nonelectrolytes,
except for acids and bases.

Electrolytes

- A strong electrolyte dissociates completely when dissolved in water.
- A weak electrolyte only dissociates partially when dissolved in water.
- A nonelectrolyte does not dissociate in water

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids (H . . .)	
		Weak bases (NH_3)	All other compounds

Acids, definition

- Acid: Increases H^+ concentration in solution
- HCl $\rightarrow$ H^+ + Cl^-
- Base: Increases OH^- concentration in solution
- NaOH $\rightarrow$ Na^+ + OH^-

Strong Electrolytes Are…

- **Strong acids,** dissociate completely in solution

Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides (LiOH, NaOH, KOH, RbOH, CsOH)
Hydrobromic, HBr	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Hydroiodic, HI Chloric, $HClO_3$ Perchloric, $HClO_4$ Nitric, HNO_3 Sulfuric, H_2SO_4	

The 7 common strong acids

KNOW THEM

Strong Electrolytes Are…

- Strong acids
- Strong bases

NOTE THIS IS MORE STUFF YOU NEED TO KNOW

Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides ($LiOH$, $NaOH$, KOH, $RbOH$, $CsOH$)
Hydrobromic, HBr	Heavy group 2A metal hydroxides [$Ca(OH)_2$, $Sr(OH)_2$, $Ba(OH)_2$]
Hydroiodic, HI	
Chloric, $HClO_3$	
Perchloric, $HClO_4$	
Nitric, HNO_3	
Sulfuric, H_2SO_4	

The strong bases

KNOW THEM!!!!

Weak acids and bases

- Acids or bases that do not dissociate completely.
- $HCH_3CO_2 \rightarrow H^+ + CH_3CO_2^-$
- Mostly stays acetic acid.

Weak base:

NH_3 ammonia.

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

The only one you know.

Strong Electrolytes Are…

- Strong acids
- Strong bases
- Soluble ionic salts
- If the salt doesn't dissolve, it can't conduct.

- For example:
- NaCl
- KNO_3
- $Mg(NO_3)_2$
- $LiClO_4$
- Etc. Any ionic compound

Exam 1 rooms

- Exam time: Monday sept 24, 7:15-8:15 pm
- Sect 57-63: 1281 Anthony hall
- 64-67 402 computer center
- 68-70 1279 Anthony Hall

- Alternate exam: Monday 9/24 6:45 am-7:45 am 138 chemistry

Naming acids and their anions

- HCl is a gas, but in water (aqueous solution), it is hydrochloric acid, HCl(aq)

- HNO_3 is nitric acid (from the nitrate anion)
- HNO_2 is nitr<u>ous</u> acid (from the nitr<u>ite</u> anion)
- HClO is hypochlor<u>ous</u> acid (from hypochlor<u>ite</u>
- Other examples:
- H_2SO_4
- HCN
- $HBrO_2$
- CH_3CO_2H

Types of reactions and their equations

- **Acid-base**
- $HCl + NaOH \rightarrow NaCl + H_2O$
- But ions dissociate; to show that:
- $H^+ + Cl^- + Na^+ + OH^- \rightarrow Na^+ + Cl^- + H_2O$
- Called a detailed ionic equation. Now cross out everything that is the same on both sides:
- $H^+ + \cancel{Cl^-} + \cancel{Na^+} + OH^- \rightarrow \cancel{Na^+} + \cancel{Cl^-} + H_2O$
- Gives:
- $H^+ + OH^- \rightarrow H_2O$ A net ionic equation.

Types of reactions and their equations

- **<u>Precipitation</u>**, the formation of a product that is insoluble:

$AgNO_3 + NaCl \rightarrow AgCl(s) + NaNO_3$

$Ag^+ + \cancel{NO_3^-} + \cancel{Na^+} + Cl^- \rightarrow AgCl(s) + \cancel{NO_3^-} + \cancel{Na^+}$

$Ag^+ + Cl^- \rightarrow AgCl(s)$

AgCl (silver chloride) is insoluble and precipitates as a solid out of the solution. So write as formula, not ionized.

Types of reactions and their equations

- **<u>Gas forming</u>**, the formation of a product that is a gas:

$NiCO_{3(s)}$ + $2HCl$ → $NiCl_2$ + H_2O + $CO_{2(g)}$

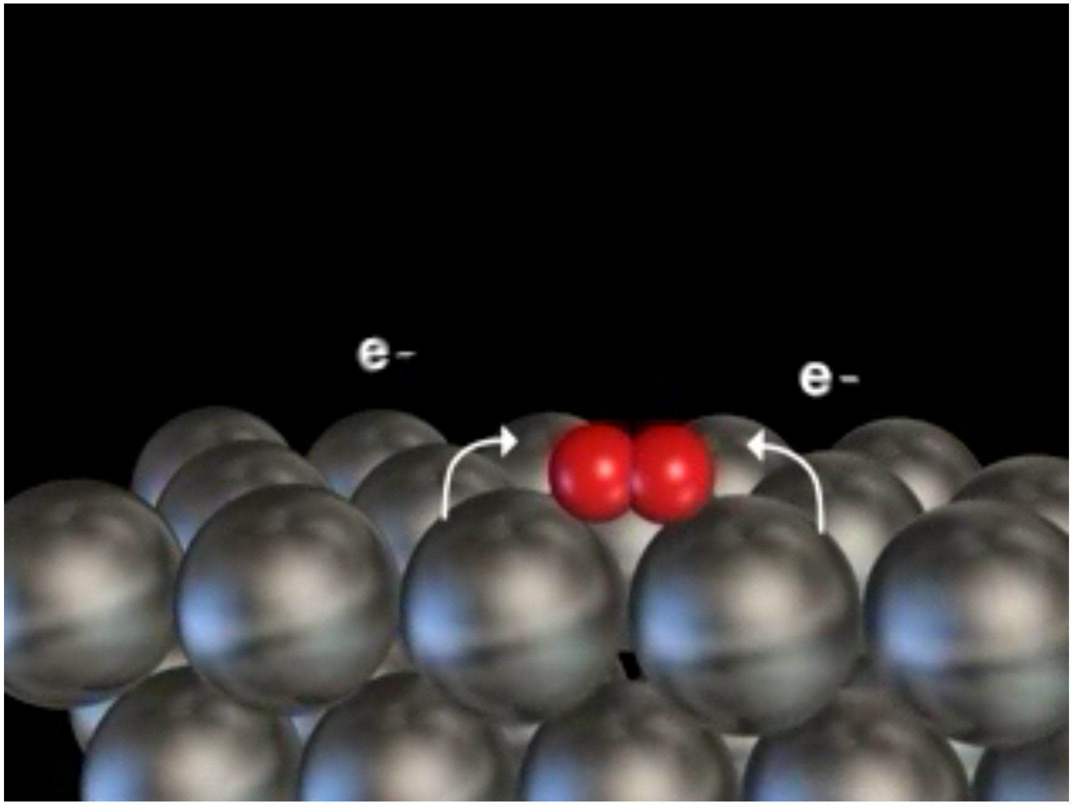
$NiCO_{3(s)}$ + $2H^+$ + $\cancel{2Cl}^{2-}$ → Ni^{2+} + $\cancel{2Cl}^{2-}$ + H_2O + $CO_{2(g)}$

$NiCO_{3(s)}$+ $2H^+$ → Ni^{2+} H_2O + $CO_{2(g)}$

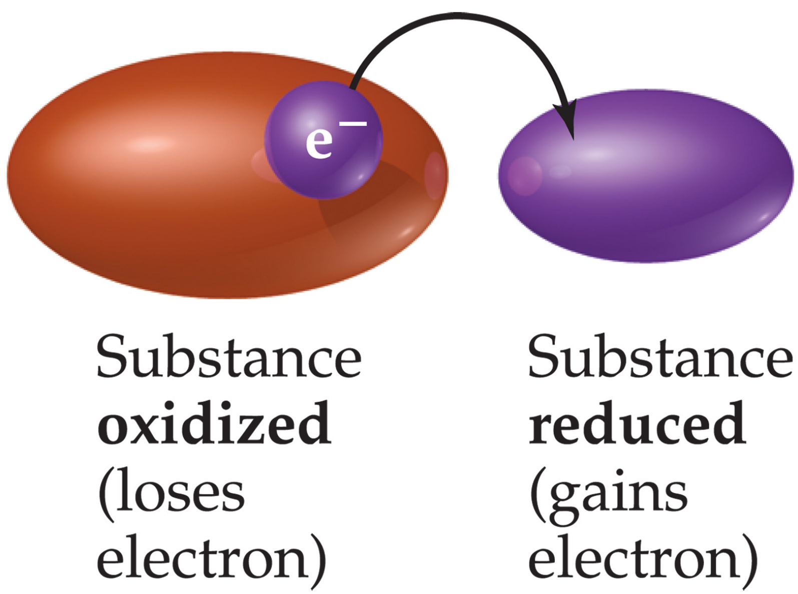
The carbon dioxide (CO_2) gas is mostly insoluble and bubbles out of solution.

Oxidation-Reduction Reactions

- An oxidation occurs when an atom or ion *loses* electrons.
- A reduction occurs when an atom or ion *gains* electrons.



Oxidation-Reduction Reactions



One cannot occur without the other.

Oxidation Numbers

To determine if an oxidation-reduction reaction has occurred, we assign an oxidation number to each element in a neutral compound or charged entity.

Book-keeping for electrons

Assigning Oxidation Numbers

- Elements in their elemental form have an oxidation number of 0.
- The oxidation number of a monatomic ion is the same as its charge.

Na oxidation number 0
Na^+ oxidation number +1

Assigning Oxidation Numbers

- Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions (when they are bound to other nonmetals).
  - Oxygen has an oxidation number of −2, except in the peroxide ion (O_2^{2-}) in which it has an oxidation number of −1.
    - CO_2, H_2O, CaO etc. O has -2 oxidation number
  - Hydrogen is −1 when bonded to a metal, +1 when bonded to a nonmetal.
    - NaH H has -1 oxidation number
    - HCl H has +1 oxidation number
    - CH_4 H has +1 oxidation number

Group 1A elements always oxidation number +1, group IIA always have +2 oxidation number.

Oxidation Numbers

- Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.
  - Fluorine always has an oxidation number of −1.
  - The other halogens have an oxidation number of −1 when the oxidation number is negative;
  - they can have positive oxidation numbers, however, most notably in oxyanions.
    - CCl_4, HCl, Cl o.n. -1
    - ClO_4^- Cl o.n. +7 (must be because O is always negative)
    - HCOCl Cl o.n. -1

Oxidation Numbers

- The sum of the oxidation numbers in a neutral compound is 0.
- The sum of the oxidation numbers in a polyatomic ion is the charge on the ion.

CCl_4 Cl o.n. -1 -1(4) = -4. C o.n. +4

ClO_4^- O: -2(4) = -8 Cl: +7 (7-8=-1)

Oxidation Numbers

ClO_2^-	Mg_3P_2	SO_4^{2-}	MnO_4	BrF_3
CaH_2	$XeOF_4$			
CO_3^{2-}	NO_3^-	$FeCl_3$	SF_6	H_2S
H_2SO_4	CaH_2	BBr_3	SO_3^-	ClO^-

Oxidation reduction reactions

Oxidation is when an element loses electrons
results in increase in oxidation number

Reduction is when an element gains electrons
results in a decrease in oxidation number

- A redox reaction is when elements gain or lose electrons during the process.
- Oxidation is always exactly balanced by reduction. The number of electrons lost in oxidation must equal the number of electrons gained in reduction
- **Example:**

$$2Mg + CO_2 \rightarrow 2MgO + C$$

Which element is reduced? This is called the oxidizing agent.

Which element is oxidized? This is called the reducing agent.

Solution stoichiometry

Reactions that happen in solution

Depend on Concentration:

moles reactant/volume of solution

- H_2SO_4 + $2NaOH$ → Na_2SO_4 + $2H_2O$
- 1 mole 2 moles 1 mole 2 moles

- Volume matters. Depends on moles/L Molarity M

Solution Stoichiometry

- Two important relationships:
- Mass/molar mass = # moles

- # moles = molarity * volume
- Allows you to measure out a volume and know the # of moles
- Also: mmol = mol/L*mL

Examples:

- Dissolve 50. g sulfuric acid in enough water to make 250 mL solution. What is the molarity?
- 50 g/98 $gmol^{-1}$= 0.51 .51 mole/0.25 L= 2.0 M

- What mass of NaOH is required to make 15 L of a 0.2 M solution.

- 0.2 mole/L(15 L) = 3 moles
- 3 moles*(40 $gmol^{-1}$) = 120 g.

Examples

- How many mL of a 6.0 M solution of HCl solution need to be added to water to make 1.0 L of a 0.15 M HCl solution?
- #molesNeeded: 1.0 L(0.15molL^{-1}) = 0.15 mol
- Volume 6M HCl solution = 0.15 mol/(6.0 molL^{-1}) = 0.025 L

Lecture 10, Redox reactions

- All chemical reactions can be divided into 2 categories:
  - Acid-base, “oxidation numbers” stay same
  - Redox, “oxidation numbers” change.

But what are oxidation numbers?

A convention for keeping track of electrons during a chemical reaction. Example:

$$\underset{0}{Na} + \underset{+1\ -2}{H_2O} \rightarrow \underset{+1}{Na^+} + \underset{-2\ +1}{OH^-} + \underset{0}{1/2H_2}$$