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

\[
2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}
\]

<table>
<thead>
<tr>
<th>#moles</th>
<th>Before reaction</th>
<th>After reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>(14)</td>
<td>10 (\text{H}_2) and 7 (\text{O}_2)</td>
<td>10 (\text{H}_2\text{O}) and 2 (\text{O}_2)</td>
</tr>
<tr>
<td>(7)</td>
<td>10 (\text{H}_2) and 7 (\text{O}_2)</td>
<td>10 (\text{H}_2\text{O}) and 2 (\text{O}_2)</td>
</tr>
<tr>
<td>Left:</td>
<td>0</td>
<td>2</td>
</tr>
</tbody>
</table>
Limiting Reactants

In the example below, the $\text{O}_2$ 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?

\[ 2N_2O + 3O_2 \rightarrow 4NO_2 \]

• If all \( N_2O \) used: 3 moles  \( \frac{3}{2}(3 \text{ moles}) \)
• But, don’t have 4.5 moles \( O_2 \) = 4.5 moles
• So: \( O_2 \) limiting: \( \frac{2}{3}(4) \) 4 \( \frac{4}{3}(4)=\frac{16}{3} \)

= \( \frac{8}{3} \)
Limiting reagents
A quick way to tell:

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

• \(2N_2O + 3O_2 \rightarrow 4NO_2\)
• 3 moles \(\rightarrow\) 4 moles ???
• \(3/2=1.5\) \(\rightarrow\) \(4/3=1.33\)
• 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 H₂ gas that can be produced and what quantity of the non-limiting reactant remains at the end?

• Equation: Zn(s) + 2HCl(aq) → ZnCl₂(aq) + H₂

• 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 $\text{N}_2\text{H}_4$ reacts with dinitrogen tetroxide $\text{N}_2\text{O}_4$ by this equation:

$$2 \text{N}_2\text{H}_4 + \text{N}_2\text{O}_4 \rightarrow 3 \text{N}_2 + 4 \text{H}_2\text{O}$$

When 3 mol $\text{N}_2\text{H}_4$ reacts with 2 mol $\text{N}_2\text{O}_4$, how many moles of $\text{N}_2$ are produced?

$$2\text{N}_2\text{H}_4 + \text{N}_2\text{O}_4 \rightarrow 3\text{N}_2 + 4\text{H}_2\text{O}$$

3 mole 2 mole ??

If $\text{N}_2\text{H}_4$: 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 = \frac{\text{actual amount}}{\text{theoretical amount}}

• Actual: what you actually got

• Theoretical: What you calculated you were going to get.
Percent yield example:

\[ \text{6UO}_3 + 8\text{BrF}_3 \rightarrow 6\text{UF}_4 + 4\text{Br}_2 + 9\text{O}_2 \]

\[
\begin{align*}
\text{6UO}_3 & \quad \text{8BrF}_3 & \quad \rightarrow & \quad 6\text{UF}_4 & \quad 4\text{Br}_2 & \quad 9\text{O}_2 \\
286 \text{ g mol}^{-1} & \quad 137 \text{ g mol}^{-1} & \quad & 314 \text{ g mol}^{-1} \\
357 \text{ g} & \quad 357 \text{ g} & \quad & 392.5 \text{ g} \\
1.25 \text{ mol} & \quad 2.61 \text{ mol} & \quad & 1.25(8/6)=1.67 & \quad 1.25(6/6)=1.25=392.5 \text{ g} \\
1.25 & \quad 1.25(8/6)=1.67 & \quad & 380 \text{ g} & \quad 380\text{ g}/392.5\text{ g} \times 100 = 96.8\% \\
\end{align*}
\]

If 380 g of UF\textsubscript{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 dissolves completely when dissolved in water.
- A weak electrolyte only dissolves partially when dissolved in water.
- A nonelectrolyte does not dissociate in water.

<table>
<thead>
<tr>
<th>Ionic</th>
<th>Strong Electrolyte</th>
<th>Weak Electrolyte</th>
<th>Nonelectrolyte</th>
</tr>
</thead>
<tbody>
<tr>
<td>Molecular</td>
<td>All</td>
<td>None</td>
<td>None</td>
</tr>
<tr>
<td></td>
<td>Strong acids (see Table 4.2)</td>
<td>Weak acids (H…)</td>
<td>Weak bases (NH₃)</td>
</tr>
<tr>
<td></td>
<td></td>
<td></td>
<td>All other compounds</td>
</tr>
</tbody>
</table>
Acids, definition

- Acid: Increases H⁺ concentration in solution
  - HCl → H⁺ + Cl⁻

- Base: Increases OH⁻ concentration in solution
  - NaOH → Na⁺ + OH⁻
Strong Electrolytes Are...

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

<table>
<thead>
<tr>
<th>Strong Acids</th>
<th>Strong Bases</th>
</tr>
</thead>
<tbody>
<tr>
<td>Hydrochloric, HCl</td>
<td>Group 1A metal hydroxides (LiOH, NaOH, KOH, RbOH, CsOH)</td>
</tr>
<tr>
<td>Hydrobromic, HBr</td>
<td>Heavy group 2A metal hydroxides [Ca(OH)₂, Sr(OH)₂, Ba(OH)₂]</td>
</tr>
<tr>
<td>Hydroiodic, HI</td>
<td></td>
</tr>
<tr>
<td>Chloric, HClO₃</td>
<td></td>
</tr>
<tr>
<td>Perchloric, HClO₄</td>
<td></td>
</tr>
<tr>
<td>Nitric, HNO₃</td>
<td></td>
</tr>
<tr>
<td>Sulfuric, H₂SO₄</td>
<td></td>
</tr>
</tbody>
</table>

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
- Hydrochloric, HCl
- Hydrobromic, HBr
- Hydroiodic, HI
- Chloric, HClO₃
- Perchloric, HClO₄
- Nitric, HNO₃
- Sulfuric, H₂SO₄

### Strong Bases
- Group 1A metal hydroxides (LiOH, NaOH, KOH, RbOH, CsOH)
- Heavy group 2A metal hydroxides [Ca(OH)₂, Sr(OH)₂, Ba(OH)₂]

**The strong bases**

**KNOW THEM!!!!**
Weak acids and bases

- Acids or bases that do not dissociate completely.
- $\text{HCH}_3\text{CO}_2 \rightarrow \text{H}^+ + \text{CH}_3\text{CO}_2^-$
- Mostly stays acetic acid.

Weak base:
$\text{NH}_3$ ammonia.
$\text{NH}_3 + \text{H}_2\text{O} \rightarrow \text{NH}_4^+ + \text{OH}^-$
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)

- HNO₃ is nitric acid (from the nitrate anion)
- HNO₂ is nitr**ous** acid (from the nitri**te** anion)
- HClO is hypochlor**ous** acid (from hypochlori**te**)
- Other examples:
  - H₂SO₄
  - HCN
  - HBrO₂
  - CH₃CO₂H
Types of reactions and their equations

- **Acid-base**
- \[ \text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O} \]
- But ions dissociate; to show that:
  - \[ \text{H}^+ + \text{Cl}^- + \text{Na}^+ + \text{OH}^- \rightarrow \text{Na}^+ + \text{Cl}^- + \text{H}_2\text{O} \]
- Called a detailed ionic equation. Now cross out everything that is the same on both sides:
  - \[ \text{H}^+ + \text{Cl}^- + \text{Na}^+ + \text{OH}^- \rightarrow \text{Na}^+ + \text{Cl}^- + \text{H}_2\text{O} \]
- Gives:
  - \[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \] A net ionic equation.
Types of reactions and their equations

- **Precipitation**, the formation of a product that is insoluble:

  \[
  \text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl(s)} + \text{NaNO}_3 \\
  \text{Ag}^+ + \text{NO}_3^- + \text{Na}^+ + \text{Cl}^- \rightarrow \text{AgCl(s)} + \text{NO}_3^- + \text{Na}^+ \\
  \text{Ag}^+ + \text{Cl}^- \rightarrow \text{AgCl(s)}
  \]

  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

• **Gas forming**, the formation of a product that is a gas:

\[
\text{NiCO}_3(s) + 2\text{HCl} \rightarrow \text{NiCl}_2 + \text{H}_2\text{O} + \text{CO}_2(g)
\]

\[
\text{NiCO}_3(s) + 2\text{H}^+ + 2\text{Cl}^2- \rightarrow \text{Ni}^{2+} + 2\text{Cl}^2- + \text{H}_2\text{O} + \text{CO}_2(g)
\]

\[
\text{NiCO}_3(s) + 2\text{H}^+ \rightarrow \text{Ni}^{2+}\text{H}_2\text{O} + \text{CO}_2(g)
\]

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.

Substance oxidized (loses electron) ➔ Substance reduced (gains electron)
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.

\[
\begin{align*}
\text{Na oxidation number} & \quad 0 \\
\text{Na}^+ \text{ oxidation number} & \quad +1
\end{align*}
\]
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.
  ➢ $\text{CCl}_4$, HCl, Cl o.n. -1
  ➢ $\text{ClO}_4^-$ Cl o.n. +7 (must be because O is always negative)
  ➢ HCOCICl 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.

\[
\text{CCl}_4 \quad \text{Cl o.n.} \ -1 \ -1(4) = -4 \quad \text{C o.n.} \ +4
\]
\[
\text{ClO}_4^- \quad \text{O:} \ -2(4) = -8 \quad \text{Cl:} \ +7 \quad (7-8=-1)
\]
Oxidation Numbers

<table>
<thead>
<tr>
<th>ClO$_2^-$</th>
<th>Mg$_3$P$_2$</th>
<th>SO$_4^{2-}$</th>
<th>MnO$_4$</th>
<th>BrF$_3$</th>
</tr>
</thead>
<tbody>
<tr>
<td>CaH$_2$</td>
<td>XeOF$_4$</td>
<td></td>
<td></td>
<td></td>
</tr>
<tr>
<td>CO$_3^{2-}$</td>
<td>NO$_3^-$</td>
<td>FeCl$_3$</td>
<td>SF$_6$</td>
<td>H$_2$S</td>
</tr>
<tr>
<td>H$_2$SO$_4$</td>
<td>CaH$_2$</td>
<td>BBr$_3$</td>
<td>SO$_3^-$</td>
<td>ClO$_2^-$</td>
</tr>
</tbody>
</table>
Oxidation reduction reactions

**Oxidation** is when an element loses electrons
results in an 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:**

\[ 2\text{Mg} + \text{CO}_2 \rightarrow 2\text{MgO} + \text{C} \]

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

• $\text{H}_2\text{SO}_4 + 2\text{NaOH} \rightarrow \text{Na}_2\text{SO}_4 + 2\text{H}_2\text{O}$
• 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?
  - \[ \frac{50 \text{ g}}{98 \text{ g/mol}} = 0.51 \text{ mol} \]
  - \[ \frac{0.51 \text{ mol}}{0.25 \text{ L}} = 2.0 \text{ M} \]

- What mass of NaOH is required to make 15 L of a 0.2 M solution.
  - \[ 0.2 \text{ M} \times 15 \text{ L} = 3 \text{ moles} \]
  - \[ 3 \text{ moles} \times (40 \text{ g/mol}) = 120 \text{ g} \]
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?
• \#\text{molesNeeded}: 1.0 \text{ L}(0.15\text{molL}^{-1}) = 0.15 \text{ mol}
• \text{Volume 6M HCl solution} = 0.15 \text{ mol}/(6.0 \text{ molL}^{-1}) = 0.025 \text{ 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:

\[ \text{Na} + \text{H}_2\text{O} \rightarrow \text{Na}^+ + \text{OH}^- + \frac{1}{2}\text{H}_2 \]

<table>
<thead>
<tr>
<th></th>
<th>Na</th>
<th>H₂O</th>
<th>→</th>
<th>Na⁺</th>
<th>OH⁻</th>
<th>1/2H₂</th>
</tr>
</thead>
<tbody>
<tr>
<td>0</td>
<td>+1</td>
<td>-2</td>
<td></td>
<td>+1</td>
<td>-2</td>
<td>+1</td>
</tr>
</tbody>
</table>