Chemical Reactions
- **Chemical reactions** describe processes involving chemical change.
- The chemical change involves rearranging matter.
- Converting one or more pure substances into new pure substances.
- **Reactants**
  - Substances combined in the reaction.
- **Products**
  - Substances produced in the reaction.

Chemical Equations
- **Chemical equations** are used to describe chemical reactions.
  - The chemical symbols for the **reactants** are shown on the left.
  - The chemical symbols for the **products** are shown on the right.
  - An arrow (→) is used to indicate that reactants are converting to products.
  - A plus sign (+) is used to separate individual reactants and products.

Example chemical equation:

```
2 H₂(g) + O₂(g) → 2 H₂O(l)
```

The equation accomplishes several things:

1. Demonstrates the fundamental law of conservation of matter:
   - **Atoms are neither created nor destroyed in chemical reactions, they are only rearranged.**
Stoichiometry

- Pronounced as *stoy-key-AHM-uh-tree*
- “The relationship between the quantities of reactants and products”

Balancing Chemical Equations

- For example, in the equation describing the formation of liquid water from hydrogen gas and oxygen gas:

  \[ 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) \]

  - There are four hydrogen atoms on both the left and right sides of the equation.
  - There are two oxygen atoms on both the left and right sides of the equation.
  - Therefore the equation is balanced.

Balancing Chemical Equations

- Other examples:
  
  \[ \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \]

  - Is it balanced? \( \text{NO}(\text{g}) + \text{O}(\text{g}) \rightarrow \text{NO}_2(\text{g}) \)
  - Is this OK?

  - Is it balanced? \( \text{NO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \)
  - Is this OK?

How to Balance a Chemical Equation?

- For example, in the equation describing the formation of liquid water from hydrogen gas and oxygen gas:

  \[ 2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2 \text{H}_2\text{O}(\text{l}) \]

  - First step: balance nitrogen atoms on both sides.
  - \( \text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}(\text{g}) + \text{H}_2\text{O}(\text{g}) \)

  - 2nd step: Then hydrogen atoms
    - \( 2\text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \)
    - But I will need to balance the N atoms again
    - \( 2\text{NH}_3(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \)

  - 3rd step: Balance oxygens
    - \( 2\text{NH}_3(\text{g}) + \frac{5}{2} \text{O}_2(\text{g}) \rightarrow 2\text{NO}(\text{g}) + 3\text{H}_2\text{O}(\text{g}) \)

  - 4th and final step: Just remove the fractions
    - \( 4\text{NH}_3(\text{g}) + 5 \text{O}_2(\text{g}) \rightarrow 4\text{NO}(\text{g}) + 6\text{H}_2\text{O}(\text{g}) \)

  - Now try:
    - \( \text{NO}(\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{NO}_2(\text{g}) \)
Balancing Chemical Equations

- \( \text{H}_2\text{S(aq)} + \text{I}_2\text{(aq)} \rightarrow \text{HI(aq)} + \text{S(s)} \)
- \( \text{Fe(s)} + \text{O}_2 \rightarrow \text{Fe}_2\text{O}_3 \)
- \( \text{KClO}_3\text{(s)} \rightarrow \text{KCl(s)} + \text{O}_2\text{(g)} \)
- \( \text{Mg} + \text{O}_2\text{(g)} \rightarrow \text{MgO} \)
- \( \text{Ca(OH)}_2\text{(s)} + \text{H}_3\text{PO}_4\text{(aq)} \rightarrow \text{Ca}_3\text{(PO}_4\text{)}_2\text{(s)} + \text{H}_2\text{O(l)} \)
- \( \text{P}_4\text{(s)} + \text{O}_2\text{(g)} \rightarrow \text{P}_4\text{O}_{10}\text{(s)} \)
- \( \text{Ba(ClO}_3\text{)}_2\text{(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow \text{HClO}_3\text{(aq)} + \text{BaSO}_4\text{(s)} \)
- \( \text{NH}_3\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{NO(g)} + \text{H}_2\text{O(g)} \)
- \( \text{C}_3\text{H}_8\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)} \)
- \( \text{C}_8\text{H}_{16}\text{(l)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(g)} \)

More Chemical Equations

- Balance the following equations
  - \( \text{SO}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{SO}_3\text{(g)} \)
  - \( \text{N}_2\text{(g)} + \text{O}_2\text{(g)} \rightarrow \text{N}_2\text{O(g)} \)
  - \( \text{C}_6\text{H}_{14}\text{(l)} + \text{O}_2\text{(g)} \rightarrow \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)} \)

Homework

Balancing Chemical Equations

- \( \text{H}_2\text{S(aq)} + \text{I}_2\text{(aq)} \rightarrow 2\text{HI(aq)} + \text{S(s)} \)
- \( 4\text{Fe(s)} + 3\text{O}_2 \rightarrow 2\text{Fe}_2\text{O}_3 \)
- \( 2\text{KClO}_3\text{(s)} \rightarrow 2\text{KCl(s)} + 3\text{O}_2\text{(g)} \)
- \( 2\text{Mg} + \text{O}_2\text{(g)} \rightarrow 2\text{MgO} \)
- \( 3\text{Ca(OH)}_2\text{(s)} + 2\text{H}_3\text{PO}_4\text{(aq)} \rightarrow \text{Ca}_3\text{(PO}_4\text{)}_2\text{(s)} + 6\text{H}_2\text{O(l)} \)
- \( \text{P}_4\text{(s)} + 5\text{O}_2\text{(g)} \rightarrow \text{P}_4\text{O}_{10}\text{(s)} \)
- \( \text{Ba(ClO}_3\text{)}_2\text{(aq)} + \text{H}_2\text{SO}_4\text{(aq)} \rightarrow 2\text{HClO}_3\text{(aq)} + \text{BaSO}_4\text{(s)} \)
- \( 4\text{NH}_3\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 4\text{NO(g)} + 6\text{H}_2\text{O(g)} \)
- \( \text{C}_2\text{H}_6\text{(g)} + 5\text{O}_2\text{(g)} \rightarrow 3\text{CO}_2\text{(g)} + 4\text{H}_2\text{O(g)} \)
- \( 2\text{C}_8\text{H}_{18}\text{(l)} + 25\text{O}_2\text{(g)} \rightarrow 16\text{CO}_2\text{(g)} + 18\text{H}_2\text{O(g)} \)

Types of Reactions

- Reaction Types
  - Combination
  - Decomposition
  - Double Replacement
  - Precipitation
  - Acid-Base
  - Single Replacement
  - Oxidation-Reduction
  - Gas Forming
### Combination Reactions

- **Combination reactions** have the form: 
  \[ A + B \rightarrow C \]
- Two or more reactants produce a single product

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

- **Other combination reactions:**
  1. \[ 2 \text{Na(s)} + \text{S(s)} \rightarrow \text{Na}_2\text{S(s)} \]
  2. \[ \text{SO}_3(g) + \text{H}_2\text{O(l)} \rightarrow \text{H}_2\text{SO}_4(aq) \]
  3. \[ 2\text{Na(s)} + \text{Cl}_2(g) \rightarrow 2\text{NaCl(s)} \]

### Decomposition Reactions

- **Decomposition reactions** have the form: 
  \[ A \rightarrow B + C \]
- Single reactant breaks down into two or more products

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

- **Further examples:**
  1. \[ 2 \text{HgO(s)} \rightarrow 2 \text{Hg(l)} + \text{O}_2(g) \]
     - This reaction was used by Joseph Priestley in the discovery of oxygen in 1774
  2. \[ \text{CaCO}_3(s) \rightarrow \text{CaO}(s) + \text{CO}_2(g) \]
     - This reaction is used industrially to produce both lime (CaO) and CO\(_2\) from limestone (CaCO\(_3\))
  3. **Air bags:**
     \[ 2 \text{NaN}_3(s) \rightarrow 2 \text{Na(s)} + 3 \text{N}_2(g) \]
Double Replacement

- **Double replacement** reactions are also called “metathesis” reactions or “partner swapping” reactions.
- They have the form $AX + BY \rightarrow BX + AY$.

$HF + NaOH \rightarrow NaF + H_2O$

- $+ = H$  
- $\bullet = O$  
- $\bigcirc = Na$  
- $\bigtriangleup = F$

Double Replacement

- **Double replacement** reactions often take place in water and are of two basic types:
  
  - **Acid base neutralization** reactions:
    
    $HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$
    
    (acid)          (base)          (salt)          (water)

    $H_2SO_4(aq) + 2 NaOH(aq) \rightarrow Na_2SO_4(aq) + H_2O(l)$

  - **Precipitation reactions** (solid forms):
    
    $Ba(NO_3)_2(aq) + Na_2SO_4(aq) \rightarrow BaSO_4(s) + 2 NaNO_3(aq)$
    
    - Here the barium sulfate **precipitates** out of solution.

Mixtures and Solutions

- **Solute**: the component of a solution that is dissolved in another substance.
- **Solvent**: the medium in which a solute dissolved to form a solution.
- **Solution**: a homogenous mixture in which the components are evenly distributed in each other.
- **Aqueous**: Any solution in which water is the solvent.
- **Precipitate**: Insoluble product of a reactant.
- **Electrolyte**: All ionic compounds that are soluble in water and conduct electricity.

- **Strong electrolyte**
- **Weak electrolyte**
- **Non-electrolyte**
Mixtures and Solutions

Figure 5.2

\[
CuCl_2(s) \rightarrow Cu^{2+}(aq) + 2Cl^-(aq)
\]

100% dissociation (strong electrolyte)

\[
CH_3COOH(aq) \rightarrow CH_3COO^-(aq) + H^+(aq)
\]

<5% ionized (weak electrolyte)

Solubility Rules

<table>
<thead>
<tr>
<th>Compounds Containing</th>
<th>Solubility</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Group IA (Na(^+), K(^+) and NH(_4^+))</td>
<td>Soluble</td>
<td></td>
</tr>
<tr>
<td>Nitrates</td>
<td>Soluble</td>
<td></td>
</tr>
<tr>
<td>Acetates (CH(_3)COO(^-))</td>
<td>Soluble</td>
<td></td>
</tr>
<tr>
<td>Perchlorates (ClO(_4^−))</td>
<td>Soluble</td>
<td></td>
</tr>
<tr>
<td>Chlorides, Bromides, Iodides</td>
<td>Soluble</td>
<td>Ag(^+), Pb(^{2+}), Hg(^+)</td>
</tr>
<tr>
<td>Fluorides</td>
<td>Soluble</td>
<td>Mg(^2+), Ca(^2+), Ag(^+), Pb(^{2+}), Sr(^{2+})</td>
</tr>
<tr>
<td>Sulfates</td>
<td>Soluble</td>
<td>Ag(^+), Ba(^{2+}), Sr(^{2+}), Pb(^{2+}), Hg(^+)</td>
</tr>
<tr>
<td>Sulfides</td>
<td>Insoluble</td>
<td>Group IA and NH(_4^+)</td>
</tr>
<tr>
<td>Carbonates</td>
<td>Insoluble</td>
<td>Group IA and NH(_4^+)</td>
</tr>
<tr>
<td>Phosphates</td>
<td>Insoluble</td>
<td>Group IA and NH(_4^+)</td>
</tr>
<tr>
<td>Hydroxides</td>
<td>Insoluble</td>
<td>Group IA and NH(_4^+)</td>
</tr>
</tbody>
</table>

Precipitation Reactions

- Precipitation reactions
  \[
  K_2SO_4(aq) + Pb(NO_3)_2(aq) \rightarrow
  \]

- Writing Equations
  - Molecular

- Total ionic

- Net ionic

Ionic Equations

- When an ion appears on both sides of a chemical equation it can be canceled out
  - "Spectator Ions"

- Net Ionic Equation

http://www.dlt.ncssm.edu/TIGER/Flash/moles/DoubleDisp_Reaction-Precipitation.html
### Ionic Equations

**Mg(ClO₄)₂(aq) + K₂CO₃(aq) → MgCO₃(s) + KClO₄(aq)**
- Total:
- Cancel:
- Net Ionic:

**Ba(NO₃)₂(aq) + Na₂SO₄(aq) → BaSO₄(s) + NaNO₃(aq)**
- Total:
- Cancel:
- Net Ionic:

### Acids and Bases

**Acid**
- a substance that, when dissolved in water, increases the concentration of hydrogen ions (H⁺) in solution

**Base**
- a substance that, when dissolved in water, increases the concentration of hydroxide ions (OH⁻) in solution

**Strong acid**
- Strong electrolyte

**Strong base**
**Acid-Base Reactions**

\[ \text{HNO}_3(aq) + \text{NaOH}(aq) \rightarrow \]

- **Total:**
- **Cancel:**
- **Net Ionic:**

**Acid-Base Reactions**

\[ \text{H}_2\text{SO}_4(aq) + \text{KOH}(aq) \rightarrow \]

- **Total:**
- **Cancel:**
- **Net Ionic:**

**Single Replacement**

- **Single replacement** reactions are also called substitution reactions.
  - They have the form \( \text{A} + \text{BX} \rightarrow \text{B} + \text{AX} \), where \( \text{A} \) and \( \text{B} \) are elements and \( \text{BX} \) and \( \text{AX} \) are compounds.

\[ \text{H}_2 + \text{CuO} \rightarrow \text{H}_2\text{O} + \text{Cu} \]

- **Oxidation**
- **Reduction**

**Single Replacement**

- Other **single replacement** reactions
  - \( 3 \text{C(s)} + 2 \text{Fe}_2\text{O}_3(s) \rightarrow 4 \text{Fe(s)} + 3 \text{CO}_2(g) \)
  - \( \text{Cu(NO}_3\)_2(aq) + \text{Zn(s)} \rightarrow \text{Zn(NO}_3\)_2(aq) + \text{Cu(s)} \)

- Single replacement reactions are also **oxidation-reduction** (REDOX) reactions.
  - **Oxidation**
  - **Reduction**

**OXIDATION NUMBER**
Redox Reactions and Electron Transfer

The **oxidation number** of an atom is the charge that atom would have if the compound was composed of ions.

**Oxidation of Mg (0) to Mg (II):**
2Mg (s) + O_2 (g) → 2MgO (s)

**Reduction of Fe (III) to Fe(0):**
Fe_2O_3 (s) + 3CO (g) → 2Fe (s) + 3 CO_2 (g)

**Reduction of Ag (I) to Ag and oxidation of Cu (0) to Cu (II):**
2Ag^+ (aq) + Cu(s) → 2 Ag (s) + Cu^{2+} (aq)

### Gas Forming Reactions

**Table 5.3** Gas-Forming Reactions

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metal carbonate or bicarbonate + acid</td>
<td>metal salt + CO_2(g) + H_2O(l)</td>
</tr>
<tr>
<td>Na_2CO_3(aq) + 2 HCl(aq) → 2NaCl(aq) + CO_2(g) + H_2O(l)</td>
<td></td>
</tr>
<tr>
<td>Metal sulfide + acid</td>
<td>metal salt + H_2S(g)</td>
</tr>
<tr>
<td>Na_2S(aq) + 2 HCl(aq) → 2NaCl(aq) + H_2S(g)</td>
<td></td>
</tr>
<tr>
<td>Metal sulfite + acid</td>
<td>metal salt + SO_2(g) + H_2O(l)</td>
</tr>
<tr>
<td>Na_2SO_3(aq) + 2 HCl(aq) → 2NaCl(aq) + SO_2(g) + H_2O(l)</td>
<td></td>
</tr>
<tr>
<td>Ammonium salt + strong base</td>
<td>metal salt + NH_3(g) + H_2O(l)</td>
</tr>
<tr>
<td>NH_4Cl(aq) + NaOH(aq) → NaCl(aq) + NH_3(g) + H_2O(l)</td>
<td></td>
</tr>
</tbody>
</table>

**Gas Forming Reactions**

- Reactions leading to the formation of an insoluble gas
- Examples

Na_2CO_3 (aq) + 2HCl (aq) → 2NaCl (aq) + CO_2 (g) + H_2O (l)