3: Chemical Reactions

1. Chemical Equations
2. Types of Chemical Equations
3. Balancing Chemical Equation
4. Aqueous Reactions
5. Acid-Base Reactions
6. Redox Reactions

### 3.1 Chemical Equations

**How Do Chemists Describe Chemical Reactions?**

**Chemical Equations:**

<table>
<thead>
<tr>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
</table>

What would be the equation describing how hydrogen burns with oxygen (a combustion reaction)?

\[ H_2 + O_2 \rightarrow H_2O \]

Is mass conserved in the above equation?

---

**The Physical States**

\[ 2H_2 (g) + O_2(g) \rightarrow 2H_2O(l) \]

Know Symbols for the States

<table>
<thead>
<tr>
<th>s</th>
<th>Solid</th>
</tr>
</thead>
<tbody>
<tr>
<td>g</td>
<td>gas, vapor</td>
</tr>
<tr>
<td>l</td>
<td>liquid</td>
</tr>
<tr>
<td>(aq)</td>
<td>aqueous</td>
</tr>
</tbody>
</table>

What is the difference between aqueous and liquid?

**Chemical Equations**

- Aqueous is a Solution Dissolved in Water
- Liquid is a Pure Substance

\[ NaCl(l) - Molten Sodium Chloride \]
\[ NaCl(aq) - Salt Water \]

What Other Information Can Be Contained in a Chemical Equation?

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**Reaction Conditions**

\[ 2H_2 (g) + O_2(g) \xrightarrow{\Delta Pt} 2H_2O(l) \]

As indicated, this reaction will not proceed at a reasonable rate without heat and a catalyst.

- A catalyst effects the way the reaction occurs. It is a chemical which is consumed in one step and reproduced during a subsequent step. So it is conserved during the course of the reaction.

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Types of Chemical Reactions

1. Synthesis
2. Decomposition Reactions
3. Single Displacement Reactions
4. Double Displacement Reactions
5. Combustion Reactions
Synthesis

\[ 2\text{KCl} + 3\text{O}_2 \rightarrow 2\text{KClO}_3 \]

Several Species Combine to Create a New Species

Decomposition Reactions

\[ 2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2 \]

One Species Decomposes into Multiple Species

(Different Than a Dissociation)

\[ \text{KClO}_3 \rightarrow \text{K}^+ + \text{ClO}_4^- \]

Single Displacement Reactions

\[ \text{Mg} + \text{CuCl}_2 \rightarrow \text{MgCl}_2 + \text{Cu} \]
(Metal Displacement)

\[ \text{BaCl}_2 + \text{F}_2 \rightarrow \text{BaF}_2 + \text{Cl}_2 \]
(Nonmetal Displacement)

One Metal or Nonmetal “Displaces” a Different Metal or Nonmetal in a Salt

Double Displacement Reactions (metathesis)

\[ \text{KCl} + \text{NaI} \rightarrow \text{KI} + \text{NaCl} \]

Two Salts Swap Ions

Combustion Reaction

\[ \text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} \]

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

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

“Burning” - Exothermic (Energy Releasing) Reaction

Balancing Chemical Equations

Basic Principles:

• Atoms Are Conserved
• Never Change the Subscript. That Changes the Identity of the Compound

\[ \text{H}_2 (\text{g}) + \text{O}_2(\text{g}) \rightarrow \text{H}_2\text{O}(\text{l}) \]
(Combustion of Hydrogen)

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

Represents the Formation of Hydrogen Peroxide, Not the Combustion of Hydrogen.
Balancing Chemical Equations

Basic Principles:

• Atoms Are Conserved
• Never Change the Subscript. That Changes the Identity of the Compound
• Balance by Changing the Stoichiometric Coefficient

2H₂(g) + O₂(g) → 2H₂O(l)

Stoichiometric Coefficients

(Comes from Greek: stoicheion - element & metron - measure)

Balancing Chemical Equations

Basic Techniques:

1. Identify Reactants and Products, write correct formulas
2. Write Unbalanced (Skeletal) Equation (reactants left of the arrow, products to the right), leave space for the stoichiometric coefficient in front of each compound
3. Balance Equation by Inspection (start with most complicated molecules first, least complicated last)
4. Check Work, make sure the same number of all atoms exist on both sides of the equation.

Balancing Equation Problem

Balance the Reaction for DRANO (Al/NaOH)

_Al(s) + _NaOH(s) + _H₂O(l) → _NaAl(OH)₄(aq) + _H₂(g)

Tip: Place a blank line in front of each species to indicate where you place the stoichiometric coefficient. This is the only number you can change while balancing the equation

Balancing Equation Problem

Balance Oxygen by Adding Water. (Do Not Add NaOH as Sodium is Already Balanced)

_Al(s) + _NaOH(s) + _H₂O(l) → _NaAl(OH)₄(aq) + _H₂(g)

Always Balance Elements Which Are “Pure” (Not in a Compound) Last

Balancing Double Displacement Reactions

Do Not Balance the Elements

Balance the Ions

Double Displacement Problem

Write the Balanced Equation for the Formation of Aluminum Chlorite & Ammonium Sulfate From Aluminum Sulfate and Ammonium Chlorite
Balancing Combustion Reactions

Balance Oxygen Last

What Is the Equation Describing How a Butane Lighter Works (Butane = C\textsubscript{4}H\textsubscript{10})

The Reaction Involves the Combustion of Butane With Oxygen (the Reactants) Forming Carbon Dioxide and Water (the Products)

Interactive Quizzes:

4A: Balancing Combustion Reactions
4B: Double Displacement Reactions
http://www.ualr.edu/rebelford/chem1402/q1402/chem1402QP.htm

Liquid Phase Solutions

**Solvent** - substance present in greatest amount.

**Solute** - substance(s) dissolved in solvent

Solute Can Be:
1. **Liquid** (alcohol in the wine)
2. **Solid** (salt in the sea water)
3. **Gas** (oxygen in your blood)

5.2 Aqueous Solvations

Covalent Compounds Do Not Ionize

Ionic Compounds Ionize

When a Solute Become Dissolved in a Solvent, It Becomes Solvated, or Hydrated When the Solvent Is Water

Aqueous Solutions

1. **Strong Electrolytes** – Strong Conductors of Electricity due to formation of a large number of Mobile Ions

2. **Weak Electrolytes** – Weak Conductors of Electricity due to formation of a few Mobile Ions

3. **NonElectrolytes** – Non Conductors of Electricity as they do not form Ions in aqueous solutions
Strong Electrolytes

**Ionic - Soluble Salts and Strong Bases**

NaCl(aq) → Na⁺(aq) + Cl⁻(aq)
NaOH(aq) → Na⁺(aq) + OH⁻(aq)

**Covalent - Strong Acids (protonate water)**

HCl(aq) + H₂O → H₃O⁺(aq) + Cl⁻(aq)
H₂SO₄(aq) + H₂O → H₃O⁺(aq) + HSO₄⁻(aq)

Weak Electrolytes

**Ionic - Slightly Soluble Salts**

CoCl₂(s) ⇌ Co⁺²(aq) + 2Cl⁻(aq)

**Covalent - Weak Acids & Amine Bases**

HF(aq) + H₂O ⇌ H₃O⁺(aq) + F⁻(aq)
NH₃(aq) + H₂O ⇌ NH₄⁺(aq) + OH⁻

Non Electrolytes

**Ionic - Insoluble Salts**

CoS(aq) ⇌ Co⁺²(aq) + S²⁻(aq)

**Covalent - Molecules which do not hydrolyze or protonate water**

C₁₂H₂₂O₁₁(s) + H₂O → C₁₂H₂₂O₁₁(aq)

3.6 Precipitation Reactions

(Metathesis Reactions
Double Displacement Reactions)

**Soluble Salts** - Have Anions and Cations Which Move Around in the Water Without Coming Together and Forming a Crystal Structure

**Insoluble Salts** - Have Anions and Cations Which Come Together in Water and Form a Solid (Crystal Structure), a **PRECIPITATE**

Predicting the Formation of Precipitates

Upon Mixing Two Soluble Salts, SWAP PARTNERS, and Determine If Resulting Salts Are Insoluble.

C₁A₁ + C₂A₂ → C₁A₂ + C₂A₁, where,

C₁ = Cation from 1st Salt  C₂ = Cation from 2nd Salt
A₁ = Anion from 1st Salt   A₂ = Anion from 2nd Salt

(Note, the subscripts in the actual salt formulas reflect the charge of the opposite ion)

Precipitation Reactions

The Big Question: Will an Aqueous Mixture of Two Soluble Salts Form a Precipitate (an Insoluble Salt)?

**STRATEGY:**

1. Identify the Ions Present
2. Apply Solubility Rules to Determine If the Resultant Salts Are Soluble or Insoluble

Note: Solubility Rules are based on the types of Ions the Salts are composed of.
Solubility Rules

I. Usually Soluble - Salts with
   a. cations: Group IA Cations and NH₄⁺
   b. anions: NO₃⁻, ClO₄⁻, ClO₃⁻, CH₃COO⁻

II. Usually Soluble - Salts with
   a. Cl⁻, Br⁻, I⁻ (Except those with Ag⁺,Hg₂⁺²⁺,Pb⁺²)
   b. F⁻ (Except those with Mg²⁺,Ca²⁺,Sr²⁺,Ba²⁺,Pb⁺²)
   c. SO₄²⁻ (Except those with Pb⁺², Ca²⁺, Sr⁺², Ba⁺²)

III. Insoluble - (Except With Above Ions), Salts With
   a. CO₃²⁻, PO₄³⁻, C₂O₄²⁻, CrO₄²⁻, S²⁻,
   b. OH⁻ (Except those with Ca²⁺, Sr⁺², Ba⁺²)

What is an Insoluble Salt?

• Salts Which Are Soluble to a Negligible Extent

Note: These rules are generalizations and suffice for most common salts (more advanced treatments use the concept of Equilibrium Constants)

Predicting and Writing Double Displacement Reactions

1. Identify Reactant Ions
2. Swap Partners and Identify Products
3. Write Skeletal Equation
4. Balance Skeletal Equation

Interactive Quizzes 5.1-5.5

Three Ways to Write Ionic Equations

1. General (Molecular Equations)
   - use ionic formulas, must include phase
2. Total Ionic Equations
   - show soluble ions as individual species
3. Net Ionic Equation
   - Do not include “Spectator Ions”, that is, ions which do not react in any way.

Ionic and Net Ionic Equations

Predict if a precipitate (ppt) will occur for mixtures of: Pb(NO₃)₂(aq) and KCl(aq)

Ionic and Net Ionic Equations

Predict if a precipitate (ppt) will occur for mixtures of: Mg(NO₃)₂(aq) and KI(aq)
3.7: Aqueous Acid & Base Reactions

**Auto-ionization**

\[ 2 \text{H}_2\text{O(aq)} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^- \]

\( \text{H}_3\text{O}^+ = \text{Hydronium Ion} \)

\( \text{OH}^- = \text{Hydroxide Ion} \)

---

**Acids – Proton Donors**

Acidus (Latin for “Sour”)

*Arrhenius Definition* - a Substance Which Increases \( \text{H}^+ \) When Dissolved in Water

\[ \text{HCl} \rightleftharpoons (\text{H}_2\text{O}) \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) \]

(Note, \( \text{H}^+ \) actually exists as \( \text{H}_3\text{O}^+ \), the hydronium ion)

*Bronstead Definition – A Proton Donor*

---

**Bases – Proton Acceptors**

Bases - *Alkali*, bitter and slippery to touch

Bases produce hydroxide (OH\(^-\)) when added to water

\[ \text{NaOH(s)} \rightleftharpoons (\text{H}_2\text{O}) \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \]

(soluble metal hydroxides)

\[ \text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^- \]

(amine type compound)

---

**Strong Acids**

Acids which completely ionize

Heavier Acid Halides

Larger Oxyacids

<table>
<thead>
<tr>
<th>Acid</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>HNO(_3)</td>
</tr>
<tr>
<td>HBr</td>
<td>(\text{H}_2\text{SO}_4)</td>
</tr>
<tr>
<td>HI</td>
<td>(\text{HClO}_4)</td>
</tr>
<tr>
<td></td>
<td>(\text{HClO}_3)</td>
</tr>
</tbody>
</table>

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**Strong Bases**

1. Hydroxides of
   a. Alkali metals (NaOH, KOH…)
   b. Heavier Alkaline Earths (Ca(OH)\(_2\)...)

<table>
<thead>
<tr>
<th>Base</th>
<th>Formula</th>
</tr>
</thead>
<tbody>
<tr>
<td>LiOH</td>
<td>Ca(OH)(_2)</td>
</tr>
<tr>
<td>NaOH</td>
<td>Sr(OH)(_2)</td>
</tr>
<tr>
<td>KOH</td>
<td>Ba(OH)(_2)</td>
</tr>
</tbody>
</table>

---

**Neutralization Reactions**

1. Strong Acid + Strong Base
2. Strong Acid + Weak Base
3. Weak Acid + Strong Base
4. Weak Acid + Weak Base
### Neutralization Reactions

1. **Strong Acid + Strong Base**
   
   \[ \text{HCl(aq)} + \text{NaOH(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(l)} \]
   
   Total Ionic Equation:
   
   \[ \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{H}_2\text{O(l)} \]
   
   Net Ionic Equation:
   
   \[ \text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{H}_2\text{O(l)} \]

2. **Strong Acid + Weak Base**
   
   \[ \text{HCl(aq)} + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+\text{Cl(aq)} \]
   
   Total Ionic Equation:
   
   \[ \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{Cl}^-(\text{aq}) \]
   
   Net Ionic Equation:
   
   \[ \text{H}^+(\text{aq}) + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) \]

3. **Weak Acid + Strong Base**
   
   \[ \text{HF(aq)} + \text{NaOH(aq)} \rightarrow \text{NaF(aq)} + \text{H}_2\text{O(l)} \]
   
   Total Ionic Equation:
   
   \[ \text{HF(aq)} + \text{Na}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightarrow \text{Na}^+(\text{aq}) + \text{F}^-(\text{aq}) + \text{H}_2\text{O(l)} \]
   
   Net Ionic Equation:
   
   \[ \text{HF(aq)} + \text{OH}^-(\text{aq}) \rightarrow \text{F}^-(\text{aq}) + \text{H}_2\text{O(l)} \]

4. **Weak Acid + Weak Base**
   
   \[ \text{HF(aq)} + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+\text{F(aq)} \]
   
   Total Ionic Equation:
   
   \[ \text{HF(aq)} + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{F}^-(\text{aq}) \]
   
   Net Ionic Equation:
   
   \[ \text{HF(aq)} + \text{NH}_3(\text{aq}) \rightarrow \text{NH}_4^+(\text{aq}) + \text{F}^-(\text{aq}) \]

(Interactive Quiz 5f:)

### 3.9 Oxidation - Reduction Reactions

*(Redox Reactions)*

Involve the Transfer of Electrons Between Different Reactants

Historically:

Oxidation Meant the Formation of Oxides and Reduction Meant the Removal of Oxygen From the Oxide

#### Redox Terminology

**Oxidation** - Loss of Electrons

**Reduction** - Gain of Electrons

Note Oxidation and Reduction Must Occur Concurrently!

**Oxidant** - Reactant Which Oxidizes (gets Reduced)

**Reductant** - Reactant Which Reduces (gets Oxidized)
**Redox Example**

Formation of Iron(III) Oxide

Oxidation of Iron into an oxide

\[ 4Fe + 3O_2 \rightarrow 2Fe_2O_3 \]

Fe \[ \rightarrow \] Fe^{3+} + 3e^- (oxidation)

O_2 + 4e^- \[ \rightarrow \] 2O^2- (reduction)

**Types of Redox Reactions**

**Name Some Types of Redox Reactions**

- Formation of Ionic Compound from a Metal and a Nonmetal
- Corrosion of Metals (Rusting)
- Combustion Reactions
- Electric Batteries

**Identification of Redox Rxns**

- Assign each element an Oxidation Number (State). If any element changes its oxidation state during a reaction, it is a redox reaction

**Zero Oxidation States** indicate element is neutral

**[+] Oxidation States** indicate element is “electron poor”

**[-] Oxidation States** indicate element is “electron rich”

**Oxidation Numbers**

1. Oxd # = 0 for pure elements
2. Oxd # = charge of monatomic ion
3. Oxd # of F = -1 in compounds with other elements
4. Oxd # of Cl, Br & I = -1 in compounds except with Oxygen & Fluorine
5. Oxd # of H= +1, except for Metal Hydrides (-1)
6. Oxd # of O = -2, except with fluorides, Peroxides (-1) and Superoxides (-1/2)
7. The Sum of the oxd # ’s of all elements in a compound = 0, and = the charge of a polyatomic ion

**Oxidation Number Problems**

Identify the oxidation number/state of all atoms in the following species

1. SO_4^{2-}
2. H_2O_2
3. NaH
4. HPO_3^{2-}
5. KMnO_4
6. K_2MnO_4
7. K_3MnO_4
8. MnO_2

**Can You Have Non-integer Oxidation States?**

Fe_3O_4

YES

O = 2

Fe = 8/3
Oxidation States & Nomenclature

Look at Oxyanions and Oxyacids

-ate & -ic refer to higher oxidation states of nonoxygen

-ite & -ous refer to lower oxidation states of nonoxygen

-Identify oxidation state of chlorine in the 4 oxy-chlorides

Which is nitrous and which is nitric oxide, NO or N₂O?

Single Displacement Reactions

Metals exist in 2 forms

1. Elemental (M)
2. Cationic (M⁺ⁿ)

What happens when Magnesium is placed in aqueous Copper(II) Chloride

Mg(s) + Cu²⁺(aq) --> Cu(s) + Mg²⁺(aq)

The Magnesium is more “Active” and has “Displaced” the Copper from the Salt

Single Displacement Reactions

What Happens When Magnesium Is Placed in Aqueous Sodium Chloride?

Nothing

So the Sodium Is More “Active” Than the Magnesium.

How Do We Predict If a Metal Will Displace Another From a Salt?

Predicting Single Displacement Reactions

Determine if:
1. Mg will dissolve in CrCl₃(aq)
2. Ag will dissolve in Fe(NO₃)₂
3. Mg will dissolve in KCl

From Table 4.3 of handouts, Explain why gold and Platinum are often called the “Noble Metals”