IONIC COMPOUNDS

Compounds in Aqueous Solution

Many reactions involve ionic compounds, especially reactions in water — aqueous solutions.

KMnO₄ in water

K⁺(aq) + MnO₄⁻(aq)
Copper chloride is added to water. Interactions between water and the Cu$^{2+}$ and Cl$^{-}$ ions allow the solid to dissolve.

The ions are now sheathed in water molecules.

An Ionic Compound, CuCl$_2$, in Water

Positive H atoms attracted to negative ion
Negative O atom attracted to positive ion

Water can interact with positive cations and negative anions in aqueous solution.
How do we know ions are present in aqueous solutions?

The solutions **conduct electricity**!

They are called **ELECTROLYTES**.

HCl, MgCl$_2$, and NaCl are **strong electrolytes**. They dissociate completely (or nearly so) into ions.
Aqueous Solutions

HCl, MgCl\(_2\), and NaCl are **strong electrolytes**. They dissociate completely (or nearly so) into ions.
Aqueous Solutions

Acetic acid ionizes only to a small extent, so it is a weak electrolyte.

\[ \text{CH}_3\text{CO}_2\text{H}(aq) \rightarrow \text{CH}_3\text{CO}_2^-(aq) + \text{H}^+(aq) \]
Aqueous Solutions

Acetic acid ionizes only to a small extent, so it is a weak electrolyte.

\[
\text{CH}_3\text{CO}_2\text{H}(aq) \quad \rightarrow \quad \text{CH}_3\text{CO}_2^-(aq) + \text{H}^+(aq)
\]
Aqueous Solutions

Some compounds dissolve in water but do not conduct electricity. They are called nonelectrolytes.

Examples include:
- sugar
- ethanol
- ethylene glycol
# Water Solubility of Ionic Compounds

## Soluble Compounds

<table>
<thead>
<tr>
<th>Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>Almost all salts of Na⁺, K⁺, NH₄⁺</td>
<td>Halides of Ag⁺, Hg₂⁺, Pb⁺²</td>
</tr>
<tr>
<td>Salts of nitrate, NO₃⁻, chlorate, ClO₃⁻, perchlorate, ClO₄⁻, acetate, CH₃CO₂⁻</td>
<td></td>
</tr>
<tr>
<td>Almost all salts of Cl⁻, Br⁻, I⁻</td>
<td>Fluorides of Mg⁺², Ca⁺², Sr⁺², Ba⁺²⁺, Pb⁺²⁺</td>
</tr>
<tr>
<td>Compounds containing F⁻</td>
<td>Sulfates of Mg⁺², Ca⁺², Sr⁺², Ba⁺²⁺, Pb⁺²⁺</td>
</tr>
<tr>
<td>Salts of sulfate, SO₄²⁻</td>
<td></td>
</tr>
</tbody>
</table>

## Insoluble Compounds

<table>
<thead>
<tr>
<th>Compounds</th>
<th>Exceptions</th>
</tr>
</thead>
<tbody>
<tr>
<td>All salts of carbonate, CO₃²⁻, phosphate, PO₄³⁻, oxalate, CO₂⁺², chromate, CrO₄²⁻</td>
<td>Salts of NH₄⁺ and the alkali metal cations</td>
</tr>
<tr>
<td>Most metal sulfides, S²⁻</td>
<td></td>
</tr>
<tr>
<td>Most metal hydroxides and oxides</td>
<td></td>
</tr>
</tbody>
</table>

*Screen 5.4 & Figure 5.1*
Water Solubility of Ionic Compounds

Common minerals are often formed with anions that lead to insolubility:

- sulfide
- fluoride
- carbonate
- oxide

Azurite, a copper carbonate

Iron pyrite, a sulfide

Orpiment, arsenic sulfide
ACIDS

An acid ------> H⁺ in water

Some **strong** acids are

- HCl, hydrochloric
- H₂SO₄, sulfuric
- HClO₄, perchloric
- HNO₃, nitric
An acid $\rightarrow$ $H^+$ in water

HCl(aq) $\rightarrow$ $H^+(aq)$ + $Cl^-(aq)$
The Nature of Acids

HCl → Cl⁻ + H₂O → H⁺ + H₂O → H₃O⁺

hydronium ion
Weak Acids

WEAK ACIDS = weak electrolytes

CH$_3$CO$_2$H
acetic acid

H$_2$CO$_3$
carbonic acid

H$_3$PO$_4$
phosphoric acid

HF
hydrofluoric acid

Acetic acid

phosphoric acid

hydrofluoric acid
Nonmetal oxides can be acids

\[ \text{CO}_2(aq) + \text{H}_2\text{O}(liq) \rightarrow \text{H}_2\text{CO}_3(aq) \]

\[ \text{SO}_3(aq) + \text{H}_2\text{O}(liq) \rightarrow \text{H}_2\text{SO}_4(aq) \]

and can come from burning coal and oil.
**BASES**

see Screen 5.9 and Table 5.2

Base $\rightarrow$ OH$^-$ in water

NaOH(aq) $\rightarrow$ Na$^+$ (aq) + OH$^-$ (aq)

NaOH is a strong base
Ammonia, NH₃
An Important Base

NH₃(aq) + H₂O(ℓ) ⇌ NH₄⁺(aq) + OH⁻(aq)

ammonia, base weak electrolyte < 100% ionized
water

ammonium ion hydroxide ion

04m08an2.mov
Metal oxides are bases

\[
\text{CaO(s)} + \text{H}_2\text{O(liq)} \rightarrow \text{Ca(OH)}_2(\text{aq})
\]

CaO in water. Indicator shows solution is basic.
Table 5.2  • Common Acids and Bases

<table>
<thead>
<tr>
<th>Strong Acids (Strong Electrolytes)</th>
<th>Strong Bases (Strong Electrolytes)</th>
</tr>
</thead>
<tbody>
<tr>
<td>HCl</td>
<td>Hydrochloric acid</td>
</tr>
<tr>
<td>HBr</td>
<td>Hydrobromic acid</td>
</tr>
<tr>
<td>HI</td>
<td>Hydroiodic acid</td>
</tr>
<tr>
<td>HNO₃</td>
<td>Nitric acid</td>
</tr>
<tr>
<td>HClO₄</td>
<td>Perchloric acid</td>
</tr>
<tr>
<td>H₂SO₄</td>
<td>Sulfuric acid</td>
</tr>
</tbody>
</table>

<table>
<thead>
<tr>
<th>Weak Acids (Weak Electrolytes)</th>
<th>Weak Base (Weak Electrolyte)</th>
</tr>
</thead>
<tbody>
<tr>
<td>H₃PO₄</td>
<td>Phosphoric acid</td>
</tr>
<tr>
<td>H₂CO₃</td>
<td>Carbonic acid</td>
</tr>
<tr>
<td>CH₃CO₂H</td>
<td>Acetic acid</td>
</tr>
<tr>
<td>H₂C₂O₄</td>
<td>Oxalic acid</td>
</tr>
<tr>
<td>C₄H₆O₆</td>
<td>Tartaric acid</td>
</tr>
<tr>
<td>C₆H₈O₇</td>
<td>Citric acid</td>
</tr>
<tr>
<td>C₉H₈O₄</td>
<td>Aspirin</td>
</tr>
</tbody>
</table>

*These are representative of hundreds of weak acids.
Net Ionic Equations

\[ \text{Mg(s)} + 2 \text{HCl(aq)} \rightarrow \text{H}_2(g) + \text{MgCl}_2(\text{aq}) \]

We really should write

\[ \text{Mg(s)} + 2 \text{H}^+(\text{aq}) + 2 \text{Cl}^-\text{(aq)} \rightarrow \text{H}_2(g) + \text{Mg}^{2+}(\text{aq}) + 2 \text{Cl}^-\text{(aq)} \]

The two Cl\(^-\) ions are **SPECTATOR IONS** — they do not participate. Could have used NO\(_3^-\).
Net Ionic Equations

Mg(s) + 2 HCl(aq) → H₂(g) + MgCl₂(aq)

Mg(s) + 2 H⁺(aq) + 2 Cl⁻(aq) → H₂(g) + Mg²⁺(aq) + 2 Cl⁻(aq)

We leave the spectator ions out —
Mg(s) + 2 H⁺(aq) → H₂(g) + Mg²⁺(aq)

to give the **NET IONIC EQUATION**
Chemical Reactions in Water
Sections 5.2 & 5.4-5.6—CD-ROM Ch. 5

We will look at EXCHANGE REACTIONS

\[
\text{AX} + \text{BY} \rightarrow \text{AY} + \text{BX}
\]

The anions exchange places between cations.

\[
Pb(NO_3)_2(aq) + 2 \text{KI}(aq) \rightarrow PbI_2(s) + 2 \text{KNO}_3(aq)
\]
The “driving force” is the formation of an insoluble compound — a precipitate.

\[
Pb(NO_3)_2(aq) + 2 KI(aq) \rightarrow 2 KNO_3(aq) + PbI_2(s)
\]

Net ionic equation

\[
Pb^{2+}(aq) + 2 I^-(aq) \rightarrow PbI_2(s)
\]
Acid–Base Reactions

• The “driving force” is the formation of water.
  \[
  \text{NaOH(aq)} + \text{HCl(aq)} \rightarrow \text{NaCl(aq)} + \text{H}_2\text{O(liq)}
  \]

• Net ionic equation
  \[
  \text{OH}^-\text{(aq)} + \text{H}^+\text{(aq)} \rightarrow \text{H}_2\text{O(liq)}
  \]

• This applies to ALL reactions of strong acids and bases.
Acid-Base Reactions

CCR, page 162

HCl (acid) + NaOH (base) → NaCl (salt) + H₂O

H⁺(aq) + Cl⁻(aq) + Na⁺(aq) + OH⁻(aq)
Acid-Base Reactions

- A-B reactions are sometimes called **NEUTRALIZATIONS** because the solution is neither acidic nor basic at the end.

- The other product of the A-B reaction is a **SALT**, MX.

$$HX + MOH \rightarrow MX + H_2O$$

$M^{n+}$ comes from **base** & $X^{n-}$ comes from **acid**

This is one way to make compounds!
Gas-Forming Reactions

This is primarily the chemistry of metal carbonates.

CO₂ and water ---> H₂CO₃

H₂CO₃(aq) + Ca²⁺ --->

2 H⁺(aq) + CaCO₃(s) (limestone)

Adding acid reverses this reaction.

MCO₃ + acid ---> CO₂ + salt
Gas-Forming Reactions

CaCO₃(s) + H₂SO₄(aq) --->

2 CaSO₄(s) + H₂CO₃(aq)

Carbonic acid is unstable and forms CO₂ & H₂O

H₂CO₃(aq) ---> CO₂ + water

(Antacid tablet has citric acid + NaHCO₃)
**Table 5.3 • Gas-Forming Reactions**

<table>
<thead>
<tr>
<th>Reaction Description</th>
<th>Chemical Equation</th>
</tr>
</thead>
<tbody>
<tr>
<td>Metal carbonate or bicarbonate + acid → metal salt + CO₂(g) + H₂O(ℓ)</td>
<td>Na₂CO₃(aq) + 2 HCl(aq) → 2 NaCl(aq) + CO₂(g) + H₂O(ℓ)</td>
</tr>
<tr>
<td>Metal sulfide + acid → metal salt + H₂S(g)</td>
<td>Na₂S(aq) + 2 HCl(aq) → 2 NaCl(aq) + H₂S(g)</td>
</tr>
<tr>
<td>Metal sulfite + acid → metal salt + SO₂(g) + H₂O(ℓ)</td>
<td>Na₂SO₃(aq) + 2 HCl(aq) → 2 NaCl(aq) + SO₂(g) + H₂O(ℓ)</td>
</tr>
<tr>
<td>Ammonium salt + strong base → metal salt + NH₃(g) + H₂O(ℓ)</td>
<td>NH₄Cl(aq) + NaOH(aq) → NaCl(aq) + NH₃(g) + H₂O(ℓ)</td>
</tr>
</tbody>
</table>
Quantitative Aspects of Reactions in Solution
Sections 5.8-5.10
In solution we need to define the:

- **SOLVENT**
  the component whose physical state is preserved when solution forms

- **SOLUTE**
  the other solution component
Concentration of Solute

The amount of solute in a solution is given by its concentration.

Molarity (M) = \( \frac{\text{moles solute}}{\text{liters of solution}} \)

Concentration (M) = [ …]
1.0 L of water was used to make 1.0 L of solution. Notice the water left over.
PROBLEM: Dissolve 5.00 g of NiCl$_2$•6 H$_2$O in enough water to make 250 mL of solution. Calculate molarity.

**Step 1:** Calculate moles of NiCl$_2$•6H$_2$O

\[
5.00 \text{ g} \times \frac{1 \text{ mol}}{237.7 \text{ g}} = 0.0210 \text{ mol}
\]

**Step 2:** Calculate molarity

\[
\frac{0.0210 \text{ mol}}{0.250 \text{ L}} = 0.0841 \text{ M}
\]

\[[\text{NiCl}_2\cdot6 \text{ H}_2\text{O}] = 0.0841 \text{ M}\]
The Nature of a CuCl\textsubscript{2} Solution Ion Concentrations

CuCl\textsubscript{2}(aq) \rightarrow Cu^{2+}(aq) + 2 \text{Cl}^{-}(aq)

If \([\text{CuCl}_2] = 0.30\ \text{M},\) then
\([\text{Cu}^{2+}] = 0.30\ \text{M}\)
\([\text{Cl}^{-}] = 2 \times 0.30\ \text{M}\)
What mass of oxalic acid, H$_2$C$_2$O$_4$, is required to make 250. mL of a 0.0500 M solution?

Because

\[
\text{Conc (M)} = \frac{\text{moles}}{\text{volume}} = \frac{\text{mol}}{V}
\]

this means that

\[
\text{moles} = M \cdot V
\]
What mass of oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$, is required to make 250. mL of a 0.0500 M solution?

$$\text{moles} = M \cdot V$$

**Step 1:** Calculate moles of acid required.

$$(0.0500 \text{ mol/L})(0.250 \text{ L}) = 0.0125 \text{ mol}$$

**Step 2:** Calculate mass of acid required.

$$(0.0125 \text{ mol})(90.00 \text{ g/mol}) = 1.13 \text{ g}$$
Preparing Solutions

- Weigh out a solid solute and dissolve in a given quantity of solvent.
- Dilute a concentrated solution to give one that is less concentrated.
PROBLEM: You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

Add water to the 3.0 M solution to lower its concentration to 0.50 M

Dilute the solution!
PROBLEM: You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do? 

H₂O

But how much water do we add?

3.0 M NaOH
Concentrated

0.50 M NaOH
Dilute
PROBLEM: You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

How much water is added?
The important point is that --->

moles of NaOH in ORIGINAL solution = moles of NaOH in FINAL solution
**Problem:** You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

Amount of NaOH in original solution =

\[ M \cdot V = \]

\[ (3.0 \text{ mol/L})(0.050 \text{ L}) = 0.15 \text{ mol NaOH} \]

Amount of NaOH in final solution must also =

0.15 mol NaOH

Volume of final solution =

\[ (0.15 \text{ mol NaOH})(1 \text{ L/0.50 mol}) = 0.30 \text{ L} \]

or 300 mL
PROBLEM: You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What do you do?

Conclusion: add 250 mL of water to 50.0 mL of 3.0 M NaOH to make 300 mL of 0.50 M NaOH.
Preparing Solutions by Dilution

A shortcut

\[ C_{\text{initial}} \cdot V_{\text{initial}} = C_{\text{final}} \cdot V_{\text{final}} \]
**pH, a Concentration Scale**

pH: a way to express acidity -- the concentration of H\(^+\) in solution.

- **Low pH**: high [H\(^+\)]
- **High pH**: low [H\(^+\)]

- **Acidic solution**: pH < 7
- **Neutral**: pH = 7
- **Basic solution**: pH > 7
The pH Scale

\[ \text{pH} = \log \left( \frac{1}{[H^+]} \right) = -\log [H^+] \]

In a neutral solution,

\[ [H^+] = [OH^-] = 1.00 \times 10^{-7} \text{ M at 25} \degree \text{C} \]

\[ \text{pH} = -\log [H^+] = -\log (1.00 \times 10^{-7}) \]

\[ = -(-7) \]

\[ = 7 \]

See CD Screen 5.17 for a tutorial
If the [H$^+$] of soda is 1.6 x 10$^{-3}$ M, the pH is ____?

Because pH = - log [H$^+$]

then

pH = - log (1.6 x 10$^{-3}$)

pH = - (-2.80)

pH = 2.80
If the pH of Coke is 3.12, it is ______________.

Because pH = - log [H⁺] then

\[ \log [H^+] = - \text{pH} \]

Take antilog and get

\[ [H^+] = 10^{-\text{pH}} \]

\[ [H^+] = 10^{-3.12} = 7.6 \times 10^{-4} \text{ M} \]
• Zinc reacts with acids to produce $H_2$ gas.
• Have 10.0 g of Zn
• What volume of 2.50 M HCl is needed to convert the Zn completely?
GENERAL PLAN FOR STOICHIOMETRY CALCULATIONS

- Mass zinc
- Moles zinc
- Stoichiometric factor
- Moles zinc
- Mass HCl
- Moles HCl
- Volume HCl
Zinc reacts with acids to produce H₂ gas. If you have 10.0 g of Zn, what volume of 2.50 M HCl is needed to convert the Zn completely?

Step 1: Write the balanced equation
Zn(s) + 2 HCl(aq) → ZnCl₂(aq) + H₂(g)

Step 2: Calculate amount of Zn

\[
10.0 \text{ g Zn} \cdot \frac{1.00 \text{ mol Zn}}{65.39 \text{ g Zn}} = 0.153 \text{ mol Zn}
\]

Step 3: Use the stoichiometric factor
Zinc reacts with acids to produce $H_2$ gas. If you have 10.0 g of Zn, what volume of 2.50 M $\text{HCl}$ is needed to convert the Zn completely?

**Step 3:** Use the stoichiometric factor

$\text{0.153 mol Zn} \times \frac{2 \text{ mol HCl}}{1 \text{ mol Zn}} = 0.306 \text{ mol HCl}$

**Step 4:** Calculate volume of HCl req’d

$\text{0.306 mol HCl} \times \frac{1.00 \text{ L}}{2.50 \text{ mol}} = 0.122 \text{ L HCl}$
ACID-BASE REACTIONS
Titrations

\[ \text{H}_2\text{C}_2\text{O}_4(\text{aq}) + 2 \text{NaOH(} \text{aq} \text{)} \rightarrow \text{Na}_2\text{C}_2\text{O}_4(\text{aq}) + 2 \text{H}_2\text{O(} \text{liq} \text{)} \]

Carry out this reaction using a TITRATION.

Oxalic acid, \( \text{H}_2\text{C}_2\text{O}_4 \)
Setup for titrating an acid with a base

(a) 50-mL buret containing aqueous NaOH of accurately known concentration

(b) A solution of NaOH is added slowly to the sample being analyzed. The sample is mixed.

(c) When the amount of NaOH added from the buret exactly equals the amount of H⁺ supplied by the acid being analyzed, the dye (indicator) changes color.

CCR, page 186
1. Add solution from the buret.
2. Reagent (base) reacts with compound (acid) in solution in the flask.
3. Indicator shows when exact stoichiometric reaction has occurred.
4. Net ionic equation
   \[ \text{H}^+ + \text{OH}^- \rightarrow \text{H}_2\text{O} \]
5. At equivalence point
   moles H$^+$ = moles OH$^-$
LAB PROBLEM #1: Standardize a solution of NaOH — i.e., accurately determine its concentration.

1.065 g of H$_2$C$_2$O$_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point. What is the concentration of the NaOH?
1.065 g of $\text{H}_2\text{C}_2\text{O}_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point. What is the concentration of the NaOH?

**Step 1:** Calculate amount of $\text{H}_2\text{C}_2\text{O}_4$

$$
1.065 \text{ g} \cdot \frac{1 \text{ mol}}{90.04 \text{ g}} = 0.0118 \text{ mol}
$$

**Step 2:** Calculate amount of NaOH req’d

$$
0.0118 \text{ mol acid} \cdot \frac{2 \text{ mol NaOH}}{1 \text{ mol acid}} = 0.0236 \text{ mol NaOH}
$$
1.065 g of H$_2$C$_2$O$_4$ (oxalic acid) requires 35.62 mL of NaOH for titration to an equivalence point. What is the concentration of the NaOH?

**Step 1:** Calculate amount of H$_2$C$_2$O$_4$

\[ = 0.0118 \text{ mol acid} \]

**Step 2:** Calculate amount of NaOH req’d

\[ = 0.0236 \text{ mol NaOH} \]

**Step 3:** Calculate concentration of NaOH

\[
\frac{0.0236 \text{ mol NaOH}}{0.03562 \text{ L}} = 0.663 \text{ M}
\]

\[ [\text{NaOH}] = 0.663 \text{ M} \]
LAB PROBLEM #2:
Use standardized NaOH to determine the amount of an acid in an unknown.

Apples contain malic acid, $\text{C}_4\text{H}_6\text{O}_5$.

$\text{C}_4\text{H}_6\text{O}_5(\text{aq}) + 2 \text{NaOH(aq)} \rightarrow \text{Na}_2\text{C}_4\text{H}_4\text{O}_5(\text{aq}) + 2 \text{H}_2\text{O(liq)}$

76.80 g of apple requires 34.56 mL of 0.663 M NaOH for titration. What is weight % of malic acid?
76.80 g of apple requires 34.56 mL of 0.663 M NaOH for titration. What is weight % of malic acid?

**Step 1:** Calculate amount of NaOH used.
\[ C \cdot V = (0.663 \text{ M})(0.03456 \text{ L}) \]
\[ = 0.0229 \text{ mol NaOH} \]

**Step 2:** Calculate amount of acid titrated.
\[ 0.0229 \text{ mol NaOH} \cdot \frac{1 \text{ mol acid}}{2 \text{ mol NaOH}} \]
\[ = 0.0115 \text{ mol acid} \]
76.80 g of apple requires 34.56 mL of 0.663 M NaOH for titration. What is weight % of malic acid?

**Step 1:** Calculate amount of NaOH used.

\[ 0.0115 \text{ mol acid} \times \frac{134 \text{ g}}{\text{mol}} = 1.54 \text{ g} \]

**Step 2:** Calculate amount of acid titrated

\[ 0.0229 \text{ mol NaOH} = 0.0115 \text{ mol acid} \]

**Step 3:** Calculate mass of acid titrated.
76.80 g of apple requires 34.56 mL of 0.663 M NaOH for titration. What is weight % of malic acid?

**Step 1:** Calculate amount of NaOH used.
\[ \text{= 0.0229 mol NaOH} \]

**Step 2:** Calculate amount of acid titrated.
\[ \text{= 0.0115 mol acid} \]

**Step 3:** Calculate mass of acid titrated.
\[ \text{= 1.54 g acid} \]

**Step 4:** Calculate % malic acid.
\[ \frac{1.54 \text{ g}}{76.80 \text{ g}} \times 100\% = 2.01\% \]
**Oxidation-Reduction Reactions**

*Section 5.7*

Thermite reaction

Fe$_2$O$_3$(s) + 2 Al(s)

---->

2 Fe(s) + Al$_2$O$_3$(s)
EXCHANGE: Precipitation Reactions

EXCHANGE
Gas-Forming Reactions

EXCHANGE
Acid-Base Reactions

REDOX REACTIONS
REDOX REACTIONS

Oxidation—

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

\[ \text{Mg}(\text{s}) + 2 \text{HCl}(\text{aq}) \rightarrow \text{MgCl}_2(\text{aq}) + \text{H}_2(\text{g}) \]

All corrosion reactions are oxidations.

\[ 2 \text{Al}(\text{s}) + 3 \text{Cu}^{2+}(\text{aq}) \rightarrow 2 \text{Al}^{3+}(\text{aq}) + 3 \text{Cu}(\text{s}) \]

Reduction—

\[ \text{Fe}_2\text{O}_3(\text{s}) + 2 \text{Al}(\text{s}) \rightarrow 2 \text{Fe}(\text{s}) + \text{Al}_2\text{O}_3(\text{s}) \]
Cu(s) + 2 Ag⁺(aq)  
---> Cu²⁺(aq) + 2 Ag(s)

In all reactions if something has been oxidized then something has also been reduced
Why Study Redox Reactions

- Manufacturing metals
- Batteries
- Fuels
- Corrosion
Redox reactions are characterized by ELECTRON TRANSFER between an electron donor and electron acceptor. Transfer leads to—

1. increase in oxidation number of some element = OXIDATION

2. decrease in oxidation number of some element = REDUCTION
**OXIDATION NUMBERS**

The electric charge an element APPEARS to have when electrons are counted by some arbitrary rules:

1. Each atom in free element has ox. no. = 0.
   - Zn
   - O₂
   - I₂
   - S₈

2. In simple ions, ox. no. = charge on ion.
   - -1 for Cl⁻
   - +2 for Mg²⁺
3. O has ox. no. = -2
(except in peroxides: in H₂O₂, O = -1)

4. Ox. no. of H = +1
(except when H is associated with a metal as in NaH where it is -1)

5. Algebraic sum of oxidation numbers
   = 0 for a compound
   = overall charge for an ion
Oxidation Numbers

NH₃  N =
ClO⁻  Cl =
H₃PO₄  P =
MnO₄⁻  Mn =
Cr₂O₇²⁻  Cr =
C₃H₈  C =

Oxidation number of F in HF?
Corrosion of aluminum

\[ 2 \text{Al(s)} + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu(s)} \]

\[ \text{Al(s)} \rightarrow \text{Al}^{3+}(aq) + 3 \text{e}^- \]

- Ox. no. of Al increases as e\(^-\) are donated by the metal.
- Therefore, Al is OXIDIZED
- Al is the REDUCING AGENT in this balanced half-reaction.
Recognizing a Redox Reaction

Corrosion of aluminum

\[ 2 \text{Al}(s) + 3 \text{Cu}^{2+}(aq) \rightarrow 2 \text{Al}^{3+}(aq) + 3 \text{Cu}(s) \]

\[ \text{Cu}^{2+}(aq) + 2 \text{e}^- \rightarrow \text{Cu}(s) \]

- Ox. no. of Cu decreases as e\(^-\) are accepted by the ion.
- Therefore, Cu is REDUCED
- Cu is the OXIDIZING AGENT in this balanced half-reaction.
Notice that the 2 half-reactions add up to give the overall reaction —if we use 2 mol of Al and 3 mol of Cu\(^{2+}\).

\[
\begin{align*}
2 \text{ Al(s)} & \rightarrow 2 \text{ Al}^{3+}(aq) + 6 \text{ e}^- \\
3 \text{ Cu}^{2+}(aq) + 6 \text{ e}^- & \rightarrow 3 \text{ Cu(s)}
\end{align*}
\]

-----------------------------------------------------------

\[
2 \text{ Al(s)} + 3 \text{ Cu}^{2+}(aq) \rightarrow 2 \text{ Al}^{3+}(aq) + 3 \text{ Cu(s)}
\]

Final eqn. is balanced for mass and charge.
Metals (Cu) are reducing agents

HNO₃ is an oxidizing agent

Cu + HNO₃ → Cu²⁺ + NO₂

2 K + 2 H₂O → 2 KOH + H₂
# Recognizing a Redox Reaction

See Table 5.4

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>Oxidation</th>
<th>Reduction</th>
</tr>
</thead>
<tbody>
<tr>
<td>In terms of oxygen</td>
<td>gain</td>
<td>loss</td>
</tr>
<tr>
<td>In terms of halogen</td>
<td>gain</td>
<td>loss</td>
</tr>
<tr>
<td>In terms of electrons</td>
<td>loss</td>
<td>gain</td>
</tr>
</tbody>
</table>
Examples of Redox Reactions

Metal + halogen

\[ 2 \text{Al} + 3 \text{Br}_2 \rightarrow \text{Al}_2\text{Br}_6 \]
Examples of Redox Reactions

Metal (Mg) + Oxygen

Nonmetal (P) + Oxygen

Metal (Mg) + Oxygen
Examples of Redox Reactions

Metal + acid
Mg + HCl
Mg = reducing agent
H⁺ = oxidizing agent

Metal + acid
Cu + HNO₃
Cu = reducing agent
HNO₃ = oxidizing agent