Solutions

A closer look at mixtures.

Solutions and reactions in solution.
Reactions in Solution

Solubility

- Why Solids are Solid
- Making solutions
  - Molecular solvation
  - Ionic solvation — Dissociation
- Electrolyte solutions
  - Electrolyte & Non-Electrolyte solns
  - Electrolyte strength

- Reactions in solution
  - Double Displacement: \( AB + CD \rightleftharpoons AD + CB \)
  - Equilibrium
  - Precipitation Reactions

- Representing Aqueous Reactions
  - Molecular solutions
  - Ionic solutions
    - Molecular eqns
    - Complete Ionic eqns
    - Net Ionic eqns

Other Reaction Types

- Acid-Base Reactions
  - Neutralization; \( \text{H}_2\text{O} (l) \)
- Gas Evolution Reactions
  - \( \text{H}_2\text{S} (g), \text{CO}_2 (g), \text{NH}_3 (g), \text{NH}_4\text{OH}, \text{H}_2\text{CO}_3 \)

Oxidation & Reduction

- Single Displacement: \( A + BC \rightleftharpoons B + AC \)
  - How oxidation occurs
  - Oxidation Numbers
- Red-Ox Reactions
  - Half Reactions
  - Metal Activity
- Combustion Reactions
Solutions

- Solutions are homogenous mixtures.
- Mixtures can be liquids, gas, or solid.
- We’re going to discuss the structure of mixtures.
- How substances come into mixtures and how substances can be driven out of mixtures.
- How substances in mixtures interact.
  - ... and how that interaction facilitates chemical reaction between the mixtures components.

A **solution** is a homogenous mixture.

A **solvent** is the largest component of the mixture.

A **solute** is a smaller components of the mixture.
Why solids are solid.

- Intermolecular forces hold solids together.
  - It’s usually about plus being attracted to minus (electrostatic attraction).
  - **Molecular Solids** are held together by many types of intermolecular forces.
    - The quick story is molecules have a negative end and a positive end.
    - The negative end of one molecule sticks to the positive end of another.
    - We’ll discuss the rest in Chapter 11.
  - **Ionic Solids** are held together by one type of intermolecular force.
    - It’s a simpler story.
    - The cations stick to a bunch of anions.
    - Those anions stick to more cations.
    - The result is a big clump of particles.
Sugar dissolves in water.
- The molecules remain intact.
- Water molecules get in between sugar molecules.
- The result is a mixture of sugar and water.
- Mostly water.
Ionic Solids Dissolve in Water

- Salt dissolves in water.
- The ions separate.
- Water molecules get in between the ions.
- The result is a mixture of ions and water.
- Mostly water.
- Ions separating in solution is a process called dissociation.

\[ \text{NaCl}_\text{(s)} + \text{H}_2\text{O} \rightarrow \text{Na}^+\text{(aq)} + \text{Cl}^-\text{(aq)} \]

Dissociation of sodium chloride in water
Dissociation is often Reversible

- Dissolved ions in solution can find other dissolved ions.
- If the attraction between those ions is strong, they can re-associate.
- These dissolved ions form ion pairs.
- The ion pair is not a solid, it’s still dissolved in solution.
- Ions that dissociate and re-associate in solution are a kind of reversible reaction.

\[ \text{NaCl}_{(aq)} \rightleftharpoons \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)} \]

Dissociation of sodium chloride in water
Electrolytes & Acids in Solution

- Substances that dissociate in water are **electrolytes**.
- Those that do not dissociate in water are **non-electrolytes**.
- Electrolytic solutions contain dissociated ions.
- Substances that release $H^+$ are **acids**.
- Substances that accept $H^+$ are **bases**.
- **Equilibrium** is the state of a reversible reaction where the forward and reverse reactions are happening at the same rate.
- At equilibrium the ratio of products to reactants is constant.
- Different materials will have different product to reactant ratios.
- Electrolytic solutions conduct electricity.
- The more ions, the better it conducts.
- Electrical conductivity can be used to test the equilibrium ratio of dissociated ions to associated acids and electrolytes.
- Acids and electrolytes that favor the dissociated state are called **strong**.
- Acids and electrolytes that favor the associated state in water are called **weak**.

**Electrolytes:**
- eg: HCl, KNO$_3$, NaCl, CH$_3$COOH, HF

**Acids:**
- eg: HCl, CH$_3$COOH, HF, NH$_4^+$

**Bases:**
- eg: Cl$^-$, CH$_3$COO$^-$, F$^-$, NH$_3$

\[
\begin{align*}
\text{HCl (aq)} & \rightleftharpoons \text{H}^+ (aq) + \text{Cl}^- (aq) \\
\text{KNO}_3 (aq) & \rightleftharpoons \text{K}^+ (aq) + \text{NO}_3^- (aq) \\
\text{NH}_4^+ (aq) & \rightleftharpoons \text{H}^+ (aq) + \text{NH}_3 (aq) \\
\text{NaCl (aq)} & \rightleftharpoons \text{Na}^+ (aq) + \text{Cl}^- (aq) \\
\text{CH}_3\text{COOH (aq)} & \rightleftharpoons \text{CHCOO}^- (aq) + \text{H}^+ (aq) \\
\text{HF (aq)} & \rightleftharpoons \text{H}^+ (aq) + \text{F}^- (aq)
\end{align*}
\]
Electrolyte Strength

- **Strong Acids**
  - eg HCl (aq), NaCl, H₂SO₄

- **Strong Bases**
  - eg HOAc, HF (aq)

- **Soluble Ionic Salts**
  - eg HCl (aq), NaCl, H₂SO₄

- **Weak Acids**
  - eg HCl (aq) → H⁺ (aq) + Cl⁻ (aq)
  - 100 of 100 dissociate

- **Weak Bases**
  - eg HOAc, HF (aq)

- **Partially soluble Ionic Salts**
  - eg HCl (aq), NaCl, H₂SO₄

- **Molecular Substances**
  - eg Sugar, AgCl, NO₂

- **Insoluble Ionic Salts**
  - eg HCl (aq) → H⁺ (aq) + Cl⁻ (aq)
  - 4 of 100 molecules dissociate

- **Nonelectrolytes**
  - eg Sugar, AgCl, NO₂
Reactions in Solution

- Solubility
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  - Making solutions
    - Molecular solvation
    - Ionic solvation – Dissociation
  - Electrolyte solutions
    - Electrolyte & Non-Electrolyte solns
    - Electrolyte strength

- Other Reaction Types
  - Acid-Base Reactions
    - Neutralization; \( H_2O \) (l)
  - Gas Evolution Reactions
    - \( H_2S \) (g), \( CO_2 \) (g), \( NH_3 \) (g), \( NH_4OH \), \( H_2CO_3 \)

- Oxidation & Reduction
  - Single Displacement: \( A + BC \rightleftharpoons B + AC \)
    - How oxidation occurs
    - Oxidation Numbers
  - Red-Ox Reactions
    - Half Reactions
    - Metal Activity
  - Combustion Reactions

- Reactions in solution
  - Double Displacement: \( AB + CD \rightleftharpoons AD + CB \)
  - Equilibrium
  - Precipitation Reactions

- Representing Aqueous Reactions
  - Molecular solutions
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Reactions in Solution

- If you dissolve more than one electrolyte in solution, you get a mixture of ions.
- The ions bump into each other and apart again, trading partners and just bouncing around the solution.
- That’s not exciting.

\[
\text{KI}_{(aq)} \rightleftharpoons \text{K}^{+}_{(aq)} + \text{I}^{-}_{(aq)}
\]
\[
\text{NaNO}_3_{(aq)} \rightleftharpoons \text{Na}^{+}_{(aq)} + \text{NO}_3^{-1}_{(aq)}
\]

\[
\text{KI}_{(aq)} + \text{NaNO}_3_{(aq)} \rightleftharpoons \text{K}^{+}_{(aq)} + \text{I}^{-}_{(aq)} + \text{Na}^{+}_{(aq)} + \text{NO}_3^{-1}_{(aq)} \rightleftharpoons \text{NaI}_{(aq)} + \text{KNO}_3_{(aq)}
\]

- But those ions sometimes pair up to form things that are non-electrolytes.
- When they do an irreversible reaction occurs.

\[
\text{Pb(NO}_3^2_{(aq)} \rightleftharpoons \text{Pb}^{2+} (aq) + \text{NO}_3^{-1}_{(aq)}
\]
\[
\text{KI}_{(aq)} + \text{Pb(NO}_3^2_{(aq)} \rightleftharpoons \text{K}^{+}_{(aq)} + \text{I}^{-}_{(aq)} + \text{Pb}^{2+} (aq) + \text{NO}_3^{-1}_{(aq)}
\]
\[
\text{KI}_{(aq)} + \text{Pb(NO}_3^2_{(aq)} \rightleftharpoons \text{K}^{+}_{(aq)} + \text{I}^{-}_{(aq)} + \text{Pb}^{2+} (aq) + \text{NO}_3^{-1}_{(aq)} \rightarrow \text{PbI}_2 (s) \downarrow + \text{KNO}_3_{(aq)}
\]

- This removes dissociated ions from equilibrium. Which pulls more substrate ions into the dissociated state.
- And drives the reaction to complete formation of the non-electrolyte product.
- Possible non-electrolytes that can drive the reaction include:
  - insoluble solids (precipitates)
  - volatile gases (NH\text{3}, CO\text{2}, H\text{2}S)
  - water (H\text{2}O)
Double Displacement Reactions

- We call this class of reaction, where two electrolytes react in solution, a **double displacement reaction**.

  \[ AB + CD \rightarrow AD + CB \]

- It’s only a reaction if a product is a non-electrolyte.

  - \[ KI(s) + NaNO_3(s) \rightleftharpoons KNO_3(aq) + NaI(aq) \rightarrow \text{no reaction (write “N/R”)} \]
  - \[ KI(s) + Pb(NO_3)_2(s) \rightarrow PbI_2(s)↓ + KNO_3(aq) \rightarrow \text{a reaction because PbI}_2 \text{ is not soluble in water} \]

- When there is a reaction you can show it three different ways:

  - **Molecular Equation**: \[ KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s)↓ + KNO_3(aq) \]
  - **Complete Ionic Equation**: \[ K^+(aq) + I^-(aq) + Pb^{2+}(aq) + NO_3^{1-}(aq) \rightarrow PbI_2(s)↓ + K^+(aq) + NO_3^{1-}(aq) \]
  - **Net Ionic Equation**: \[ I^{1-}(aq) + Pb^{2+}(aq) \rightarrow PbI_2(s)↓ \]

- When there is no reaction you show it this way:

  \[ KI(s) + Pb(NO_3)_2(s) \rightarrow \text{N/R} \]

- How do you know if there’s a reaction? (non-electrolytes)

  If one of the following products form, you know a reaction occurred:
  (a) An insoluble solid (precipitate) (b) a Gas (c) Water
Solubility & Precipitation

- Different materials have different solubility properties.
- If an insoluble material forms in solution, it precipitates or falls out of solution.
Aqueous solutions of magnesium chloride and lead (II) acetate, are mixed, a bright yellow solid appears in the solution. What happened?

\[
\text{MgCl}_2(^{\text{aq}}) + \text{Pb(OAc)}_2(^{\text{aq}}) \rightarrow \text{Mg(OAc)}_2(^{\text{aq}}) + \text{PbCl}_2(^{\text{s}})
\]

\[
\text{Mg}^{2+}(^{\text{aq}}) + \text{Cl}^{-}(^{\text{aq}}) + \text{Pb}^{2+}(^{\text{aq}}) + \text{OAc}^{-}(^{\text{aq}}) \rightarrow \text{Mg}^{2+}(^{\text{aq}}) + \text{OAc}^{-}(^{\text{aq}}) + \text{PbCl}_2(^{\text{s}})
\]

\[
\text{Cl}^{-}(^{\text{aq}}) + \text{Pb}^{2+}(^{\text{aq}}) \rightarrow \text{PbCl}_2(^{\text{s}})
\]

\[
2 \text{Cl}^{-}(^{\text{aq}}) + \text{Pb}^{2+}(^{\text{aq}}) \rightarrow \text{PbCl}_2(^{\text{s}})
\]
### What forms a precipitate?

Check each step, **in order**.

<table>
<thead>
<tr>
<th><strong>Step 1</strong></th>
<th><strong>ANIONS</strong></th>
<th><strong>Soluble</strong></th>
<th><strong>Insoluble</strong></th>
</tr>
</thead>
<tbody>
<tr>
<td>Acetates</td>
<td>No precipitate</td>
<td>Insoluble forms precipitate</td>
<td></td>
</tr>
<tr>
<td>Nitrates</td>
<td>Always</td>
<td>Never</td>
<td></td>
</tr>
<tr>
<td>Ammonium</td>
<td>Always</td>
<td>Never</td>
<td></td>
</tr>
<tr>
<td>Alkali metal</td>
<td>Always</td>
<td>Never</td>
<td></td>
</tr>
<tr>
<td>Acids</td>
<td>(the ones we learned)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Carbonates</td>
<td>Never</td>
<td>Always</td>
<td></td>
</tr>
<tr>
<td>Phosphates</td>
<td>(PO₄²⁻)</td>
<td></td>
<td></td>
</tr>
<tr>
<td>Halogens</td>
<td>(Cl⁻, Br⁻, I⁻, F⁻)</td>
<td>Usually</td>
<td></td>
</tr>
<tr>
<td>Sulfates</td>
<td>(SO₄²⁻)</td>
<td>Usually</td>
<td></td>
</tr>
<tr>
<td>Sulfides</td>
<td>(S²⁻)</td>
<td>Except: Sr²⁺, Ba²⁺, Ca²⁺</td>
<td></td>
</tr>
<tr>
<td>Hydroxy Salts</td>
<td>(OH⁻)</td>
<td>Usually</td>
<td></td>
</tr>
</tbody>
</table>

If you remember 1-3 you’ll be good 85% of the time
If you remember 1-3 and 4 you’ll be good 95%
Remembering the exceptions isn’t that hard
— there’s only six ions that cause exceptions
and lead, mercury, and silver are the most commonly encountered ones.
Is it soluble?

- **Always:**
  - Acetates
  - Nitrates
  - Ammonium
  - Alkali metal
  - Acids
  - Carbonates
  - Phosphates

- **Usually:**
  - Halogens
  - Sulfates
  - Sulfides
  - Hydroxy Salts

- ✔️ KNO₃
- ✔️ (NH₄)₃P
- ✔️ MnCl₂
- ✔️ HClO₃
- ✔️ CuCH₃CO₂
- ✔️ Ca(OAc)₂
- ✔️ CaCO₃

- ❌ PbCl₂
Reactions in Solution

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Acid-Base Reactions

- Acids and bases have multiple definitions.
- For now:
  - An acid is any substance which dissociates to release H\(^+\) (aq).
  - A base is any substance which reacts with H\(^+\) (aq).


\[(\text{You will explore other definitions in Chem 220.})\]

- Acid-base reactions are reactions between an acid and a base.
- Neutralization reactions are irreversible reactions between an acid and a base.
- Neutralization reactions produce water.
- The irreversible production of water can drive equilibrium forward, the same as precipitate formation.

\[
\text{HCl} \text{ (aq)} + \text{NaOH} \text{ (aq)} \rightleftharpoons \text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)} + \text{Na}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)}
\]

\[
\text{HCl} \text{ (aq)} + \text{NaOH} \text{ (aq)} \rightarrow \text{H}_2\text{O} \text{ (l)} + \text{NaCl} \text{ (aq)}
\]

\[
\text{H}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)} + \text{Na}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)} \rightarrow \text{H}_2\text{O} \text{ (l)} + \text{Cl}^- \text{ (aq)} + \text{Na}^+ \text{ (aq)}
\]

\[
\text{H}^+ \text{ (aq)} + \text{OH}^- \text{ (aq)} \rightarrow \text{H}_2\text{O} \text{ (l)}
\]
Gas Formation Reactions

- Volatile gases like $\text{CO}_2 \,(\text{g})$, $\text{H}_2\text{S} \,(\text{g})$ and $\text{NH}_3 \,(\text{g})$ that form immediately bubble off.
- The gases escape, their formation is irreversible.
- Sometimes the double displacement reaction forms an unstable compound that decomposes into the gases. Example:

  \[
  \text{H}_2\text{CO}_3 \,(\text{aq}) \rightarrow \text{H}_2\text{O} \,(\text{l}) + \text{CO}_2 \,(\text{g}) \quad \text{NH}_4\text{OH} \,(\text{aq}) \rightarrow \text{H}_2\text{O} \,(\text{l}) + \text{NH}_3 \,(\text{g})
  \]

- If a double displacement reaction forms $\text{CO}_2 \,(\text{g})$, $\text{H}_2\text{S} \,(\text{g})$, or $\text{NH}_3 \,(\text{g})$ gases this irreversible reaction will drive equilibrium forward.
- If a double displacement reaction forms $\text{H}_2\text{CO}_3 \,(\text{aq})$ or $\text{NH}_4\text{Cl} \,(\text{aq})$ these decompose to gases and drive equilibrium forward.

Examples:

\[
\text{HCl} \,(\text{aq}) + \text{Na}_2\text{S} \,(\text{aq}) \rightleftharpoons \text{H}^+ \,(\text{aq}) + \text{Cl}^- \,(\text{aq}) + \text{Na}^+ \,(\text{aq}) + \text{S}^{2-} \,(\text{aq})
\]

\[
\text{HCl} \,(\text{aq}) + \text{Na}_2\text{S} \,(\text{aq}) \rightarrow \text{H}_2\text{S} \,(\text{g}) \uparrow + \text{NaCl} \,(\text{aq})
\]

\[
\text{H}^+ \,(\text{aq}) + \text{Cl}^- \,(\text{aq}) + \text{Na}^+ \,(\text{aq}) + \text{S}^{2-} \,(\text{aq}) \rightarrow \text{H}_2\text{S} \,(\text{g}) \uparrow + \text{Cl}^- \,(\text{aq}) + \text{Na}^+ \,(\text{aq})
\]

\[
\text{H}^+ \,(\text{aq}) + \text{S}^{2-} \,(\text{aq}) \rightarrow \text{H}_2\text{S} \,(\text{g}) \uparrow
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- Volatile gases like $\text{CO}_2 \ (\text{g})$, $\text{H}_2 \text{S} \ (\text{g})$ and $\text{NH}_3 \ (\text{g})$ that form immediately bubble off.
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- If a double displacement reaction forms $\text{CO}_2 \ (\text{g})$, $\text{H}_2\text{S} \ (\text{g})$, or $\text{NH}_3 \ (\text{g})$ gases this irreversible reaction will drive equilibrium forward.
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\[
\begin{align*}
\text{H}_2\text{SO}_4(\text{aq}) + \text{NaHCO}_3(\text{aq}) & \rightleftharpoons \text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Na}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) \\
\text{H}_2\text{SO}_4(\text{aq}) + \text{NaHCO}_3(\text{aq}) & \rightarrow \text{SO}_4^{2-}(\text{aq}) + \text{Na}^+(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq}) \\
\text{H}_2\text{SO}_4(\text{aq}) + \text{NaHCO}_3(\text{aq}) & \rightarrow \text{Na}_2\text{CO}_3(\text{aq}) + \text{H}_2\text{CO}_3(\text{aq}) \\
\text{H}_2\text{SO}_4(\text{aq}) + \text{NaHCO}_3(\text{aq}) & \rightarrow \text{Na}_2\text{SO}_4 \ (\text{aq}) + \text{H}_2\text{O} \ (\text{l}) + \text{CO}_2(\text{g}) \uparrow \\
\text{H}^+(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + \text{Na}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) & \rightarrow \text{SO}_4^{2-}(\text{aq}) + \text{Na}^+(\text{aq}) + \text{H}_2\text{O} \ (\text{l}) + \text{CO}_2(\text{g}) \uparrow \\
\text{H}^+(\text{aq}) + \text{HCO}_3^-(\text{aq}) & \rightarrow \text{H}_2\text{O} \ (\text{l}) + \text{CO}_2(\text{g}) \uparrow
\end{align*}
\]
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- Volatile gases like $\text{CO}_2\,(g)$, $\text{H}_2\text{S}\,(g)$ and $\text{NH}_3\,(g)$ that form immediately bubble off.
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- If a double displacement reaction forms $\text{CO}_2\,(g)$, $\text{H}_2\text{S}\,(g)$, or $\text{NH}_3\,(g)$ gases this irreversible reaction will drive equilibrium forward.
- If a double displacement reaction forms $\text{H}_2\text{CO}_3\,(aq)$ or $\text{NH}_4\text{Cl}\,(aq)$ these decompose to gases and drive equilibrium forward.

  $$\text{NaOH}_{(aq)} + \text{NH}_4\text{Cl}_{(aq)} \rightleftharpoons \text{Na}^+_{(aq)} + \text{OH}^-_{(aq)} + \text{NH}_4^+_{(aq)} + \text{Cl}^-_{(aq)}$$
  $$\text{NaOH}_{(aq)} + \text{NH}_4\text{Cl}_{(aq)} \rightarrow \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)} + \text{NH}_4\text{OH}_{(aq)}$$
  $$\text{NaOH}_{(aq)} + \text{NH}_4\text{Cl}_{(aq)} \rightarrow \text{NaCl} + \text{NH}_4\text{OH}_{(aq)}$$
  $$\text{NaOH}_{(aq)} + \text{NH}_4\text{Cl}_{(aq)} \rightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)} + \text{NH}_3(g)↑$$
  $$\text{Na}^+_{(aq)} + \text{OH}^-_{(aq)} + \text{NH}_4^+_{(aq)} + \text{Cl}^-_{(aq)} \rightarrow \text{Na}^+_{(aq)} + \text{Cl}^-_{(aq)} + \text{H}_2\text{O}_{(l)} + \text{NH}_3(g)↑$$
  $$\text{OH}^-_{(aq)} + \text{NH}_4^+_{(aq)} \rightarrow \text{H}_2\text{O}_{(l)} + \text{NH}_3(g)↑$$
Double Displacement Reactions

- If I mix two electrolytes (AB & CD), I can look at the two possible double displacement products (AD & CB) to predict if a reaction will occur.
- If either of the two products forms irreversibly, a reaction will occur.
  - Irreversible reactions include precipitation formation, neutralization and gas formation.
- For each pair of possible products below, did a reaction occur?

\[
\begin{align*}
AB + CD & \rightarrow AD + CB \\
\text{Fe(OAc)}_3 + \text{Mn(NO}_3)_2 & \rightarrow (\text{NH}_4)_2\text{SO}_4 + \text{H}_2\text{CO}_3 \\
\text{CO}_2 + \text{MnCl}_2 & \rightarrow \text{H}_2\text{SO}_4 + \text{Hg}_2(\text{NO}_3)_2 \\
\text{C}_2\text{H}_4\text{O} + \text{AgCl} & \rightarrow \text{HBrO}_3 + \text{H}_2\text{S} \\
\text{H}_2\text{O} + \text{KBr} & \rightarrow \text{NH}_4\text{OH} + \text{MgCO}_3
\end{align*}
\]
Predict the products...

NaCl + Mn(NO$_3$)$_3$ → N/R

NaCl + AgNO$_3$ → AgCl (s) ↓ + NaNO$_3$ (aq)

K$_2$CO$_3$ + Ca(NO$_3$)$_2$ → KNO$_3$ (aq) + CaCO$_3$ (s) ↓

K$_2$CO$_3$ + NaCl → N/R

K$_2$CO$_3$ + HBr → KBr$_{(aq)}$ + H$_2$O$_{(l)}$ + CO$_2$$_{(g)}$ ↑

FeCl$_3$ + Hg$_2$(OAc)$_2$ → Hg$_2$Cl$_2$ (s) ↓ + Fe(OAc)$_3$ (aq)

TiCl$_4$ + NH$_4$NO$_3$ → N/R

NH$_4$OH + H$_2$SO$_4$ → (NH$_4$)$_2$SO$_4$ (aq) + H$_2$O (l)
Reactions in Solution

- **Solubility**
  - Why Solids are Solid
  - Making solutions
    - Molecular solvation
    - Ionic solvation — Dissociation
  - Electrolyte solutions
    - Electrolyte & Non-Electrolyte solns
    - Electrolyte strength

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    - Molecular eqns
    - Complete Ionic eqns
    - Net Ionic eqns

- **Other Reaction Types**
  - Acid-Base Reactions
    - Neutralization; \( H_2O \) (l)
  - Gas Evolution Reactions
    - \( H_2S \) (g), \( CO_2 \) (g), \( NH_3 \) (g), \( NH_4OH \), \( H_2CO_3 \)

- **Oxidation & Reduction**
  - Single Displacement: \( A + BC \rightleftharpoons B + AC \)
    - How oxidation occurs
    - Oxidation Numbers
  - Red-Ox Reactions
    - Half Reactions
    - Metal Activity
  - Combustion Reactions
Oxidation & Reduction

- If an atom gains electrons, it’s said to be **reduced**.
  
  \[
  \text{Example: } \text{Fe}^{3+} \rightarrow \text{Fe}^{0} \\
  \text{Cl}^{0} \rightarrow \text{Cl}^{1-}
  \]

- If an atom loses electrons, it’s said to be **oxidized**.
  
  \[
  \text{Example Fe}^{0} \rightarrow \text{Fe}^{3+} \\
  \text{Cl}^{1-} \rightarrow \text{Cl}^{0}
  \]

- Chemical reactions where electrons are transferred from one atom to another are called **oxidation-reduction reactions**.
  
  \[
  \text{Example: } \text{Fe} + \text{HCl} \rightarrow \text{FeCl}_3 + \text{H}_2
  \]

- It can be tricky to figure out which atoms gained or lost electrons in a reaction.

- In the above reaction:
  
  - Iron was oxidized.
  - Chlorine neither gained nor lost electrons.
  - Hydrogen was reduced.

- To help us explore oxidation-reduction reactions we assign oxidation numbers to each atom in the solution.
Oxidation Numbers

- Every atom in solution has an oxidation number.
  - If the number goes up, the species has been oxidized.
  - If the number goes down it’s been reduced.
- Oxidation numbers can be positive or negative.
- The sum of the oxidation numbers in a molecule or ion equals its charge.
- Finding oxidation Numbers:
  - Elements in their natural state are always oxidation number 0.
    - Fe, Au, Ne, H₂, Cl₂, P₄, S₈ are all oxidation number 0.
  - Monatomic Ions have an oxidation number equal to their charge.
    - Na⁺ is 1, Mg²⁺ is 2, Ca²⁺ is 2, S²⁻ is -2, N³⁻ is -3
  - Elements in a compound or molecule…
    - Fluorine is the king. He is always oxidation number -1.
    - Hydrogen is the wild card. He’s usually:
      - +1 when bonded to non-metals
      - -1 when bonded to metals
    - Oxygen is next. Unless trumped by fluorine, oxygen is usually -2 (exception: in peroxides he’s -1)
    - Other elements get priority in order of their proximity to Fluorine:
      - elements in row 7A get -1, 6A get -2, 5A get -3
      - It’s like musical chairs, the last element get’s what ever is left over.
Chlorine’s oxidation number?

0   Cl₂
-1  NaCl (Na⁺ Cl¹⁻)
+1  ClO¹⁻
+5  HClO₃
+7  ClOF₅
+9  Cl₂O₉
Identifying Red-Ox Reaction

- When an atom's oxidation number goes up in a reaction, it’s been oxidized (lost electrons).
- When an atom's oxidation number goes down in a reaction, it’s been reduced (gained electrons).
- For underlined atom in each reaction below, determine if it’s been oxidized, reduced, or neither.

Iron rusting to Iron (III) oxide. Oxidized

\[2 \text{AgCl}_\text{(s)} + \text{H}_2(\text{g}) \rightarrow 2 \text{H}^+_(\text{aq}) + 2 \text{Ag}_\text{(s)} + 2 \text{Cl}^-_(\text{aq})\] Oxidized

\[\text{MnO}_4^-_(\text{aq}) + \text{I}^-_(\text{aq}) \rightarrow \text{Mn}^{2+}_\text{(aq)} + \text{I}_2(\text{s})\] Reduced

\[\text{Na}_3\text{PO}_4(\text{aq}) + \text{H}_2\text{SO}_4(\text{aq}) \rightarrow \text{H}_3\text{PO}_4(\text{aq}) + \text{Na}_2\text{SO}_4(\text{aq})\] Neither

Precipitating gold metal from gold ions in sea water. Reduced
Oxidation-Reduction Reactions

- Atoms that gain electrons (negative charges) are reduced.
- Atoms that loose electrons are oxidized.
- Electrons always end up somewhere. If something in the reaction is getting oxidized, something else is getting reduced.
- Red-ox processes are not an equilibrium processes — someone wins; someone looses; end of story. No trade-backs.
- You can drive equilibrium with red-ox processes, just like you drive it with other precipitation, gas formation or water formation.
- Metals can be oxidized by acids and salts (rust is an example).
- Metal oxidation often occurs by a single displacement mechanism.

\[ A + BC \rightarrow AC + B \]

\[ \text{Zn}^{(s)} + H\text{Br}^{(aq)} \rightarrow \text{ZnBr}_2^{(aq)} + H_2^{(g)} \]

\[ \text{Zn}^{(s)} + H_1^{+}\text{(aq)} + Br_1^{-}\text{(aq)} \rightarrow \text{Zn}^{2+}\text{(aq)} + Br_1^{-}\text{(aq)} + H_2^{(g)} \]

<table>
<thead>
<tr>
<th>oxidation number</th>
<th>0</th>
<th>+1</th>
<th>-1</th>
<th>+2</th>
<th>-1</th>
<th>0</th>
</tr>
</thead>
</table>

Zn is oxidized (0 goes to +2)
Hydrogen is Reduced (+1 goes to 0)
Bromine is neither.
Oxidation-Reduction Reactions

- Atoms that gain electrons (negative charges) are reduced.
- Atoms that loose electrons are oxidized.
- Electrons always end up somewhere. If something in the reaction is getting oxidized, something else is getting reduced.
- Red-ox processes are not an equilibrium processes — someone wins; someone looses; end of story. No trade-backs.
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- Metals can be oxidized by acids and salts (rust is an example).
- Metal oxidation often occurs by a single displacement mechanism.

\[ A + BC \rightarrow AC + B \]
\[ \text{Zn}_{(s)} + \text{HBr}_{(aq)} \rightarrow \text{ZnBr}_2(aq) + \text{H}_2(g) \]
\[ \text{Mn}_{(s)} + \text{Pb(NO}_3)_2(aq) \rightarrow \text{Mn(NO}_3)_2(aq) + \text{Pb}_{(s)} \]
\[ \text{Cu}_{(s)} + \text{Pb(NO}_3)_2(aq) \rightarrow \text{N/R} \]

- If the reaction will occur?
Oxidation-Reduction Half Reactions

- How do we know if the reaction happens? Look at the complete ionic equation.

  \[ \text{Mn}_\text{(s)} + \text{Pb(NO}_3\text{)}_2\text{(aq)} \rightarrow \text{Mn(NO}_3\text{)}_2\text{(aq)} + \text{Pb}_\text{(s)} \]  
  \[ \text{Mn}_\text{(s)} + \text{Pb}^{2+} + \text{NO}_3\text{ }^{1-} \rightarrow \text{Mn}^{2+} + \text{NO}_3\text{ }^{1-} + \text{Pb}_\text{(s)} \]

- Remove the spectator ions to see the net ionic equation.

  \[ \text{Mn}_\text{(s)} + \text{Pb}^{2+} \rightarrow \text{Mn}^{2+} + \text{Pb}_\text{(s)} \]

- There are two half reactions which make up the net ionic equation.

  \[ \text{Mn}_\text{(s)} \rightarrow \text{Mn}^{2+} + 2\text{e}^- \]  
  \[ \text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}_\text{(s)} \]

- The two half reactions show that we’re looking at a competition for electrons. It’s basically a tug of war.

- You can turn around one equation to compare them side to side. We need to decide who’s gonna win the fight over those two electrons.

  \[ \text{Mn}^{2+} + 2\text{e}^- \rightarrow \text{Mn}_\text{(s)} \]  
  \[ \text{Pb}^{2+} + 2\text{e}^- \rightarrow \text{Pb}_\text{(s)} \]

- We could look up numbers for whose is better at holding electrons, or we could just reference a list of “who beat’s who” — the activity series.
The Activity Series

<table>
<thead>
<tr>
<th>Metal</th>
<th>Oxidation Reaction</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>Li(s) → Li⁺(aq) + e⁻</td>
</tr>
<tr>
<td>Potassium</td>
<td>K(s) → K⁺(aq) + e⁻</td>
</tr>
<tr>
<td>Barium</td>
<td>Ba(s) → Ba²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca(s) → Ca²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na(s) → Na⁺(aq) + e⁻</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg(s) → Mg²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Aluminum</td>
<td>Al(s) → Al³⁺(aq) + 3e⁻</td>
</tr>
<tr>
<td>Manganese</td>
<td>Mn(s) → Mn²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn(s) → Zn²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr(s) → Cr³⁺(aq) + 3e⁻</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe(s) → Fe²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Cobalt</td>
<td>Co(s) → Co²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Nickel</td>
<td>Ni(s) → Ni²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn(s) → Sn²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb(s) → Pb²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Hydrogen</td>
<td>H₂(g) → 2H⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu(s) → Cu²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag(s) → Ag⁺(aq) + e⁻</td>
</tr>
<tr>
<td>Mercury</td>
<td>Hg(l) → Hg²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Platinum</td>
<td>Pt(s) → Pt²⁺(aq) + 2e⁻</td>
</tr>
<tr>
<td>Gold</td>
<td>Au(s) → Au³⁺(aq) + 3e⁻</td>
</tr>
</tbody>
</table>

- Which metal (oxidation zero) is more “active”?
- We look at the half reactions.
- An atom of an element in the activity series will displace an atom of an element below it from one of its compounds.
The Activity Series

Metal:
K
Ca
Na
Mg
Al
Zn
Fe
Ni
Sn
Pb
H
Cu
Ag
Au

Grouped Metals:
K Ca Na Mg
Ga Al Zn
Fe Co Ni
Sn Pb
H
Cu Ag Au

- Which metal (oxidation zero) is more “active”?
- An atom of an element in the activity series will displace an atom of an element below it from one of its compounds.
Metal (cation) Activity Series

- Metals
- Metalloids
- Nonmetals

1. Na
2. Cu
3. Fe
4. Mg
5. H
6. Pb
Oxidation & Reduction

- How do we know which metal gives up its electrons? Check “activity.” The more active ion is the one more likely to turn into a cation (give up its electrons).
- Which is more active (more likely to lose its electrons)?

Sodium or Iron?
Al or Co?
H₂ or Mg?
Hydrogen or Gold?
Sodium or Zinc?
Pb or Cu?
Nickel or Calcium?
Oxidation & Reduction

- How do we know which metal gives up its electrons? Check “activity.” The more active ion is the one more likely to turn into a cation (give up its electrons).

- Which reactions will occur?
  
  \[ A + BC \rightarrow AC + B \]

  \[
  \text{Na}_\text{(s)} + \text{FeBr}_3\text{(aq)} \rightarrow ?
  \]

  \[
  \text{Na} \text{ more active than Fe? Yes.}
  \]

  \[
  \text{Na}_\text{(s)} + \text{FeBr}_3\text{(aq)} \rightarrow \text{Fe}_\text{(s)} + \text{NaBr}_\text{(aq)}
  \]

  \[
  \text{Fe}_\text{(s)} + \text{Zn(ClO}_3\text{)}_2\text{(aq)} \rightarrow ?
  \]

  \[
  \text{Fe} \text{ more active than Zn? No.}
  \]

  \[
  \text{Fe}_\text{(s)} + \text{Zn(ClO}_3\text{)}_2\text{(aq)} \rightarrow \text{N/R}
  \]

  \[
  \text{Sn}_\text{(s)} + \text{HNO}_3\text{(aq)} \rightarrow ?
  \]

  \[
  \text{Sn} \text{ more active than H? Yes.}
  \]

  \[
  \text{Sn}_\text{(s)} + \text{HNO}_3\text{(aq)} \rightarrow \text{H}_2\text{(g)} \uparrow + \text{Sn(NO}_3\text{)}_4\text{(aq)}
  \]
Reactions in Solution

- **Solubility**
  - Why Solids are Solid
  - Making solutions
    - Molecular solvation
    - Ionic solvation — Dissociation
  - Electrolyte solutions
    - Electrolyte & Non-Electrolyte solns
    - Electrolyte strength

- **Reactions in solution**
  - Double Displacement: $AB + CD \rightleftharpoons AD + CB$
  - Equilibrium
  - Precipitation Reactions

- **Representing Aqueous Reactions**
  - Molecular solutions
  - Ionic solutions
    - Molecular eqns
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    - Net Ionic eqns

- **Other Reaction Types**
  - Acid-Base Reactions
    - Neutralization; H$_2$O (l)
  - Gas Evolution Reactions
    - H$_2$S (g), CO$_2$ (g), NH$_3$ (g), NH$_4$OH, H$_2$CO$_3$
  - Oxidation & Reduction
    - Single Displacement: $A + BC \rightleftharpoons B + AC$
    - How oxidation occurs
    - Oxidation Numbers
  - Red-Ox Reactions
    - Half Reactions
    - Metal Activity
  - Combustion Reactions
Combustion Reactions

- **Burning** something is causing it to combust.
- **Combustion reactions** are reacting any substance with oxygen to form the most stable binary compounds of its elements and oxygen.
- The most common products are CO\(_2\) and H\(_2\)O. Other common products are NO\(_2\) and P\(_2\)O\(_5\).
- Combustion reactions are red-ox reactions, in which oxygen is reduced.
- The driving force in combustion reactions is oxygens fierce demand for electrons. Harnessing that property of oxygen is what gave us the internal combustion engine and is at the heart of most of fuels humans use.

\[
X + O_2 \rightarrow H_2O + CO_2 + NO_2 + P_2O_5 + \ldots
\]
Reaction Types

Considering...

- **Kinetics** (what could be formed?)
  - Double Displacement
  - Single Displacement

- **Driving force** (will it happen?)
  - Precipitation Reactions
  - Acid Base Reactions
  - Gas Evolution Reactions
  - Reduction-Oxidation Reactions
    - Metal Activity
    - Combustion

... you can predict if two substances will react and what products it will likely produce.

\[
AB + CD \rightarrow AD + CB
\]

\[
A + BC \rightarrow B + AC
\]
Reactions in Solution

- **Solubility**
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Questions?