Chapter 4: Chemical Reactions

Quiz 1 Scores

![Bar chart showing quiz scores](image)
Announcements:

- Thurs Oct 6 quiz (#2) will cover chapter 3
  - Bring student ID or we cannot accept your quiz!
  - No notes, no calculators
  - You need to know your name, PID, and section #
Types of Reactions

- Oxidation-Reduction
- Precipitation
- Acid-Base

Problem: balancing equations

Balance the following reaction by placing the appropriate numbers in front of each reactant and product. Indicate the states [for example, (aq), (s), etc.] of each reactant and product:

\[ \text{C} \text{C}_{12} \text{H}_{22} \text{O}_{11} + \text{H}_2\text{SO}_4 \rightarrow \text{C} + \text{H}_2\text{SO}_4 + \text{H}_2\text{O} \]
Solution: balancing equations

\[ C_{12}H_{22}O_{11(s)} + H_2SO_4(l) \rightarrow 12 C(s) + H_2SO_4(aq) + 11 H_2O(l) \]

**How do you know the states?**

You need to memorize the states of the elements at normal (room temperature and atmospheric pressure conditions (e.g., C is a solid, He is a gas, Br\(_2\) is a liquid).

For compounds, you will be given their state if needed.

Problem: balancing difficult redox equations using half-reactions

Balance the following reaction by placing the appropriate numbers in front of each reactant and product:

\[ Fe^{2+} + MnO_4^- + H^+ \rightarrow Fe^{3+} + Mn^{2+} + H_2O \]

*This is an example of a REDOX reaction.*

*Many REDOX reactions are difficult to balance by inspection…*
Solution: balancing difficult redox equations using half-reactions

(1) Assign oxidation states to all elements:

\[ \text{Fe}^{2+} + \text{MnO}_4^- + \text{H}^+ \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + \text{H}_2\text{O} \]

Rules for Assigning Oxidation States

1. For single elements, oxidation state = charge on the ion:
2. In single-element compounds, oxidation state = 0

**In compounds, assign using the following rules:**

1. Oxidation states must sum to equal the charge on the molecule.
2. H is always +1 when combined with nonmetals, -1 when combined with metals
3. O is usually -2 (Exception: Peroxide O$_2^{2-}$)
4. Halogens are -1 unless combined with O or other halogens (the halogen on top in the periodic table wins--F is always -1).
Solution: balancing difficult redox equations using half-reactions

(1) Assign oxidation states to all elements:

\[ \text{Fe}^{2+} + \text{MnO}_4^- + \text{H}^+ \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + \text{H}_2\text{O} \]

(2) Assign who gets oxidized, who gets reduced

- Oxidation: oxidation state becomes more positive
- Reduction: oxidation state becomes more negative

Fe\textsuperscript{2+} loses 1 electron

Fe\textsuperscript{2+} + MnO\textsubscript{4}^- + H\textsuperscript{+} → Fe\textsuperscript{3+} + Mn\textsuperscript{2+} + H\textsubscript{2}O

Mn gains 5 electrons
(2) Assign who gets oxidized, who gets reduced
- Oxidation: oxidation state becomes more positive
- Reduction: oxidation state becomes more negative

\[
\begin{align*}
\text{Fe}^{2+} & \text{ loses 1 electron} \\
\text{MnO}_4^- & \text{ gains 5 electrons}
\end{align*}
\]

\[\text{Fe}^{2+} + \text{MnO}_4^- + \text{H}^+ \rightarrow \text{Fe}^{3+} + \text{Mn}^{2+} + \text{H}_2\text{O}\]

MnO\text{4}^- \text{ is reduced in the reaction; it is the oxidizing agent}

Fe\text{2+} \text{ is oxidized in the reaction; it is the reducing agent}

Solution: balancing difficult redox equations using half-reactions

(3) Split into half-reactions, balance atoms being reduced or oxidized

\[
\begin{align*}
\text{Oxidation half-reaction:} & \quad \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} \\
\text{Reduction half-reaction:} & \quad \text{MnO}_4^- \rightarrow \text{Mn}^{2+}
\end{align*}
\]
What is a half-reaction?

Example: Electrolysis of water (water splitting)

\[ \text{H}_2\text{O} \rightarrow \text{H}_2 + \frac{1}{2} \text{O}_2 \]

Oxidation half-reaction: \( \text{H}_2\text{O} \rightarrow \frac{1}{2} \text{O}_2 + 2\text{H}^+ + 2e^- \)
Reduction half-reaction: \( 2\text{H}_2\text{O} + 2e^- \rightarrow \text{H}_2 + 2\text{OH}^- \)

\[ \frac{1}{2}\text{H}_2\text{O} \rightarrow \text{H}_2 + \frac{1}{2} \text{O}_2 + 2\text{H}_2\text{O} \]

Solution: balancing difficult redox equations using half-reactions

(4) Balance oxidation numbers with electrons

Oxidation half-reaction: \( \text{Fe}^{2+} \rightarrow \text{Fe}^{3+} + e^- \)
Reduction half-reaction: \( \text{MnO}_4^- + 5e^- \rightarrow \text{Mn}^{2+} \)
Solution: balancing difficult redox equations using half-reactions

(5) Multiply one of the half-reactions to get the same number of electrons in both equations

Oxidation half-reaction: \(5 \text{Fe}^{2+} \rightarrow 5 \text{Fe}^{3+} + 5e^-\)

Reduction half-reaction: \(\text{MnO}_4^- + 5e^- \rightarrow \text{Mn}^{2+}\)

(6) Add half-reactions together

Oxidation half-reaction: \(5 \text{Fe}^{2+} \rightarrow 5 \text{Fe}^{3+} + 5e^-\)

Reduction half-reaction: \(\text{MnO}_4^- + 5e^- \rightarrow \text{Mn}^{2+}\)

\[5 \text{Fe}^{2+} + \text{MnO}_4^- + 5e^- \rightarrow 5 \text{Fe}^{3+} + 5e^- + \text{Mn}^{2+}\]
Solution: balancing difficult redox equations using half-reactions

(7) Cancel electrons, balance charges with H⁺

Oxidation half-reaction: \[ 5 \text{Fe}^{2+} \rightarrow 5 \text{Fe}^{3+} + 5 \text{e}^- \]
Reduction half-reaction: \[ \text{MnO}_4^- + 5 \text{e}^- \rightarrow \text{Mn}^{2+} \]

\[ 8 \text{H}^+ + 5 \text{Fe}^{2+} + \text{MnO}_4^- \rightarrow 5 \text{Fe}^{3+} + \text{Mn}^{2+} \]

9+ on this side 17+ on this side

(8) Balance H⁺ with H₂O

\[ 8 \text{H}^+ + 5 \text{Fe}^{2+} + \text{MnO}_4^- \rightarrow 5 \text{Fe}^{3+} + \text{Mn}^{2+} + 4 \text{H}_2\text{O} \]
Solution: balancing difficult redox equations using half-reactions

(9) Simplify, check to see if charge and atoms balance

\[ 8H^+ + 5Fe^{2+} + MnO_4^- \rightarrow 5Fe^{3+} + Mn^{2+} + 4H_2O \]

<table>
<thead>
<tr>
<th>Left side</th>
<th>Right side</th>
</tr>
</thead>
<tbody>
<tr>
<td>Number of H: 8</td>
<td>4x2 = 8</td>
</tr>
<tr>
<td>Number of Fe: 5</td>
<td>5</td>
</tr>
<tr>
<td>Number of Mn: 1</td>
<td>1</td>
</tr>
<tr>
<td>Number of O: 4</td>
<td>4</td>
</tr>
<tr>
<td>Total charge: 8 + 5(2) - 1 = 17</td>
<td>5(3) + 2 = 17</td>
</tr>
</tbody>
</table>

Aqueous solubility rules (Table 4.1)

<table>
<thead>
<tr>
<th>Soluble</th>
<th>Insoluble</th>
</tr>
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<tbody>
<tr>
<td>group 1: Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺</td>
<td>Carbonates (CO₃²⁻)</td>
</tr>
<tr>
<td>Ammonium salts: NH₄⁺</td>
<td>Chromates (CrO₄³⁻)</td>
</tr>
<tr>
<td></td>
<td>Oxalates (C₂O₄²⁻)</td>
</tr>
<tr>
<td></td>
<td>Phosphates (PO₄³⁻)</td>
</tr>
<tr>
<td></td>
<td>EXCEPT Group 1 and NH₄⁺</td>
</tr>
<tr>
<td>The Halides Cl⁻, Br⁻, I⁻</td>
<td>Sulfides (S²⁻)</td>
</tr>
<tr>
<td>EXCEPT Ag⁺, Cu⁺, Hg₂⁺, Pb²⁺</td>
<td>EXCEPT Group 1, Group 2, and NH₄⁺</td>
</tr>
<tr>
<td>All fluorides (F⁻) except Pb⁻ and Group 2</td>
<td>Hydroxides (OH⁻)</td>
</tr>
<tr>
<td></td>
<td>EXCEPT Group 1, Group 2**, and NH₄⁺</td>
</tr>
<tr>
<td>Nitrates (NO₃⁻)</td>
<td>Oxides (O²⁻)</td>
</tr>
<tr>
<td>Acetates (CH₃COO⁻)</td>
<td>EXCEPT Group 1, Group 2</td>
</tr>
<tr>
<td>Perchlorates (ClO₄⁻)</td>
<td></td>
</tr>
<tr>
<td>Sulfates (SO₄²⁻)</td>
<td></td>
</tr>
<tr>
<td>EXCEPT Ca²⁺, Sr²⁺, Ba²⁺, Pb²⁺, and Ag⁺⁺⁺⁺</td>
<td></td>
</tr>
<tr>
<td>*PbCl₂ slightly soluble</td>
<td>**Mg(OH)₂ very slightly soluble</td>
</tr>
<tr>
<td>***Ag₂SO₄ slightly soluble</td>
<td>Ca(OH)₂, Sr(OH)₂ slightly soluble</td>
</tr>
</tbody>
</table>

Green compounds are not in Table 4.1 but you should know them
Problem: precipitation and quantitative analysis

150.0 mL of a contaminated water sample contains Cd\(^{2+}\) ions at a concentration of 5.353 \times 10^{-3} \text{ M}. If 50.0 mL of a solution containing excess Na\(_2\)S is added to the solution, the precipitate that is formed will weigh:

- a) 0.155 g
- b) 6.28 \times 10^{-2} \text{ g}
- c) 0.116 g
- d) 6.28 \times 10^{-3} \text{ g}
- e) none of the above

Solution: precipitation and quantitative analysis

Balanced equation:

\[
\text{Cd}^{2+}(aq) + 2 \text{Na}^+(aq) + S^{2-}(aq) \rightarrow \text{CdS}(s) + 2 \text{Na}^+(aq)
\]

Net ionic equation:

\[
\text{Cd}^{2+}(aq) + S^{2-}(aq) \rightarrow \text{CdS}(s)
\]
How many moles of Cd$^{2+}$ are in that sample?

\[
\begin{array}{c|c}
5.353 \times 10^{-3} \text{ mol} & 0.150 \text{ L solution} \\
\hline
\text{L} & \\
\hline
\end{array}
\]

\[= 8.03 \times 10^{-4} \text{ mol Cd}^{2+}\]

How many grams of CdS are in that sample?

Molecular weight of CdS:

\[= 112.411 + 32.066 = 144.477 \text{ g/mol CdS}\]

So the number of grams of CdS\(_{(s)}\) produced is

\[
\begin{array}{c|c|c|c}
5.353 \times 10^{-3} \text{ mol Cd}^{2+} & 0.150 \text{ L solution} & 1 \text{ mol CdS} & 144.477 \text{ g CdS} \\
\hline
\text{L} & 1 \text{ mol CdS} & \text{mol CdS} & \\
\hline
\end{array}
\]

\[= 0.116 \text{ g CdS}\]

Since there is excess S\(^2\), you didn’t need to know how much of it was there…
Problem: balancing difficult redox equations using half-reactions

Balance the following reaction by placing the appropriate numbers in front of each reactant and product. Identify the oxidant and the reductant:

\[
\text{NO}_2 + \text{H}_2\text{O} \rightarrow \text{NO} + \text{HNO}_3
\]

This is another example of a REDOX reaction that is difficult to balance by inspection….
## Sections

<table>
<thead>
<tr>
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<th>TIME</th>
<th>LOCATION</th>
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<td>M 3-3:50 pm</td>
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</tr>
<tr>
<td>D09</td>
<td>F 4-4:50 pm</td>
<td>WLH 2115</td>
</tr>
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</table>

## Extras