From Balancing Redox Reactions to Determining Change of Oxidation Numbers

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From Balancing Redox Reactions to Determining Change of Oxidation Numbers

By Pong Kau Yuen and Cheng Man Diana Lau

Redox reaction is a core concept in teaching and learning chemistry. This article explores a new method for balancing organic redox reactions that requires the balancing of both atoms and charges. The $H^+$, $O$, $H_2O$, and $e^-$ are used as balanced vehicles in two half reactions. A non-oxidation number approach can be applied to both molecular and ionic equations. The article also provides standard operating procedures and examples. The number of transferred electrons is first determined by balancing a half redox reaction; consequently, the change of oxidation numbers can be calculated. The mathematical equation of $Te^\text{e} = n \Delta ON$ is established, and the change of oxidation numbers ($\Delta ON$) can be counted by the number of transferred electrons ($Te^\text{e}$) and the number of atoms with oxidation numbers change ($n$). By using this mathematical equation as a new approach, students can conveniently calculate the change of mean oxidation numbers for an assigned atom in a half redox reaction.

Redox reaction is a core concept in teaching and learning chemistry (Goodstein, 1970). The oxidation number method has been widely used for balancing redox reactions, and the determination of the oxidation number is a required step in the process (Herndon, 1997; Chang & Goldsby, 2013; Tro, 2014; Generalic & Vladislavic, 2018). However, the act of balancing organic redox reactions often poses problems for students (Lockwood, 1961; Jensen, 2009). In this article, we discuss a new method, the proton method, that can be applied to both molecular and ionic chemical equations. The method can be used for not only balancing redox equations but also determining the number of transferred electrons and the change of oxidation numbers.

Proton method: Procedures for balancing half reactions

This section shows how the proton method is used to balance both atoms and charges. $H^+$, $O$, $H_2O$, and $e^-$ are used as vehicles in two half reactions.

**Step 1.** Balance atoms.
Step 1.1. Balance all other atoms except $H$ and $O$.
Step 1.2. Balance hydrogen atoms with $H^+$.
Step 1.3. Balance oxygen atoms with $O$.
Step 1.4. Add two $H^+$ atoms for each $O$ atom.
Step 1.5. Convert two $H^+$ and one $O$ to one $H_2O$ molecule.

**Step 2.** Balance charges.
Step 2.1. Count the number of charges on both sides.
Step 2.2. Add electrons to make charges on both sides equivalent.

**Step 3.** Convert $H^+$ to $OH^-$ (optional step when working in a basic medium).
Step 3.1. Add one $OH^-$ for each $H^+$.
Step 3.2. Convert one $OH^-$ and one $H^+$ to one $H_2O$ molecule.
Step 3.3. Simplify an overall equation.

Examples for balancing molecular half reactions

**Example 1**
Convert the equation $\text{CH}_3\text{CH}_2\text{OH} \rightarrow \text{CH}_3\text{COOH}$ to $\text{C}_2\text{H}_6\text{O} \rightarrow \text{C}_2\text{H}_4\text{O}_2$.

**Step 1.** Balance atoms.

$$\begin{align*}
\text{C}_2\text{H}_6\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 \\
\text{C}_2\text{H}_6\text{O} + \text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 2\text{H}^+ \\
\text{C}_2\text{H}_6\text{O} + \text{O} + 2\text{H}^+ & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 2\text{H}^+ + 2\text{H}^+ \\
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 4\text{H}^+.
\end{align*}$$

**Step 2.** Balance charges.

$$\begin{align*}
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 4\text{H}^+ \\
(0) & \rightarrow (+4) \\
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 4\text{H}^+ + 4e^-
\end{align*}$$

**Example 2**
Convert the equation $\text{CH}_3\text{CH}_2\text{CH}_3 \rightarrow \text{CO}_2$ to $\text{C}_3\text{H}_8 \rightarrow \text{CO}_2$.

**Step 1.** Balance atoms.

$$\text{C}_3\text{H}_8 \rightarrow \text{CO}_2$$

$$\text{C}_3\text{H}_8 \rightarrow 3\text{CO}_2$$
C₂H₄ + 6O → 3CO₂ + 8H⁺
C₂H₆ + 6O + 12H⁺ → 3CO₂ + 8H⁺ + 12H²⁺
C₂H₄ + 6H₂O → 3CO₂ + 20H

Step 2. Balance charges.
C₂H₄ + 6H₂O → 3CO₂ + 20H⁻
(0)
(+20)
C₂H₄ + 6H₂O → 3CO₂ + 20H⁻ + 20e⁻

Examples for balancing ionic half reactions

Example 3
Given an ionic equation Cr₂O₇⁻² → Cr³⁺ (in acidic medium):

Step 1. Balance atoms.
Cr₂O₇⁻² → Cr³⁺
CrO₂⁻³ → 2Cr³⁺
CrO₂⁻³ → 2Cr³⁺ + 7O
Cr₂O₇⁻² + 14H⁺ → 2Cr³⁺ + 7O + 14H⁺
Cr₂O₇⁻² + 14H⁺ → 2Cr³⁺ + 7H₂O

Step 2. Balance charges.
Cr₂O₇⁻² + 14H⁺ → 2Cr³⁺ + 7H₂O
(+12)
(+6)
Cr₂O₇⁻² + 14H⁺ + 6e⁻ → 2Cr³⁺ + 7H₂O

Example 4
Given an ionic equation MnO₄⁻ → Mn₂O₄ (in basic medium):

Step 1. Balance atoms.
MnO₄⁻ → Mn₂O₄
MnO₄⁻ → Mn₂O₄ + 2O
MnO₄⁻ + 4H⁺ → MnO₂ + 2O + 4H⁺
MnO₄⁻ + 4H⁺ → MnO₂ + 2H₂O

Step 2. Balance charges.
MnO₄⁻ + 4H⁺ → MnO₂ + 2H₂O
(+3)
(0)
MnO₄⁻ + 4H⁺ + 3e⁻ → MnO₂ + 2H₂O

Step 3. Convert H⁺ to OH⁻ (in basic medium).
MnO₄⁻ + 4H⁺ + 3e⁻ → MnO₂ + 2H₂O
MnO₄⁻ + 4H⁺ + 3e⁻ + 4OH⁻ → MnO₂ + 2H₂O + 4OH⁻
MnO₄⁻ + 4H⁺ + 3e⁻ + 4OH⁻ → MnO₂ + 2H₂O + 4OH⁻
MnO₄⁻ + 2H₂O + 3e⁻ → MnO₂ + 4OH⁻

Determination of Te⁻ in a balanced half redox reaction

First, one needs to balance atoms and charges, then the number of transferred electrons can be determined accordingly. In Example 1, the reaction of C₂H₆O + H₂O → C₂H₄O₂ + 4H⁺ + 4e⁻, a loss of four electrons identifies an oxidation reaction. In Example 3, the reaction of Cr₂O₇⁻² + 14H⁺ + 6e⁻ → 2Cr³⁺ + 7H₂O, a gain of six electrons identifies a reduction reaction. An oxidation reaction can be represented by Te⁻ > 0 (positive value), and a reduction reaction can be represented by Te⁻ < 0 (negative value). These half-reaction examples are quantified and defined in Table 1.

| Table 1 |
|---|---|---|
| **Example of half redox reaction** | Te⁻ | Type |
| **C₂H₆O + H₂O → C₂H₄O₂ + 4H⁺ + 4e⁻** | +4 | loss | oxidation |
| **C₂H₆ + 6H₂O → 3CO₂ + 20H⁺ + 20e⁻** | +20 | loss | oxidation |
| **Cr₂O₇⁻² + 14H⁺ + 6e⁻ → 2Cr³⁺ + 7H₂O** | -6 | gain | reduction |
| **MnO₄⁻ + 2H₂O + 3e⁻ → MnO₂ + 4OH⁻** | -3 | gain | reduction |

Relationship among Te⁻, n, and ΔON in balanced half redox reactions

Based on Examples 1 through 4, the values of Te⁻, n, and ΔON are shown in Table 2. In Table 2, Te⁻, n, and ΔON denote the number of transferred electrons, the number of atoms with oxidation number change, and the change of mean oxidation numbers, respectively. ΔON can be counted by the difference between the oxidation number of an atom on the product side (ONf) and the oxidation number of an atom on the reactant side (ONi). An oxidation reaction can be represented by ΔON > 0 (positive value), and a reduction reaction can be represented by ΔON < 0 (negative value).

In a balanced half reaction, the relationship among Te⁻, n, and ΔON can be deduced, and the mathematical equation is shown as follows:
Te⁻ = n ΔON

| Table 2 |
|---|---|---|---|
| **Example of half redox reaction** | **Atom with oxidation numbers change** | Te⁻ | n | ΔON |
| **C₂H₆O + H₂O → C₂H₄O₂ + 4H⁺ + 4e⁻** | C | +4 | 2 | +2 |
| **C₂H₆ + 6H₂O → 3CO₂ + 20H⁺ + 20e⁻** | C | +20 | 3 | +20 + 3 |
| **Cr₂O₇⁻² + 14H⁺ + 6e⁻ → 2Cr³⁺ + 7H₂O** | Cr | -6 | 2 | -3 |
| **MnO₄⁻ + 2H₂O + 3e⁻ → MnO₂ + 4OH⁻** | Mn | -3 | 1 | -3 |
Proton method: Procedures for balancing overall redox reactions

An overall number of transferred electrons is counted by the least common multiple (LCM) of two half redox reactions. The procedures for balancing an overall reaction are as follows:

Step 1. Divide into two half reactions.
Step 2. Balance all atoms in two half reactions.
Step 3. Balance charges of two half reactions by adding electrons.
Step 4. Make electrons of two half reactions equivalent.
Step 5. Combine and simplify two half reactions.
Step 6. Convert H\(^+\) to OH\(^-\) (optional step when working in basic medium).

Examples for balancing overall redox reactions

Example 5

Convert the ionic chemical equation \(\text{CH}_3\text{CH}_2\text{OH} + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{CH}_3\text{COOH} + \text{Cr}^{3+}\) (in acidic medium) to \(\text{C}_2\text{H}_6\text{O} + \text{Cr}_2\text{O}_7^{2-} \rightarrow \text{C}_2\text{H}_4\text{O}_2 + \text{Cr}^{3+}\).

Step 1. Divide into two half reactions:
\[
\begin{align*}
\text{C}_2\text{H}_6\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 \\
\text{Cr}_2\text{O}_7^{2-} & \rightarrow \text{Cr}^{3+} 
\end{align*}
\]

Step 2. Balance all atoms in two half reactions.
\[
\begin{align*}
\text{C}_2\text{H}_6\text{O} + \text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 2\text{H}^+ \\
\text{C}_2\text{H}_6\text{O} + 2\text{H}^+ & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 4\text{H}^+ \\
\text{Cr}_2\text{O}_7^{2-} & \rightarrow 2\text{Cr}^{3+} + 7\text{O} \\
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ & \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O} \\
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- & \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\end{align*}
\]

Step 3. Balance charges of two half reactions by adding electrons.
\[
\begin{align*}
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \\
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- & \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\end{align*}
\]

Step 4. Make electrons of two half reactions equivalent (LCM = 12).
\[
\begin{align*}
\text{C}_2\text{H}_6\text{O} + \text{H}_2\text{O} & \rightarrow \text{C}_2\text{H}_4\text{O}_2 + 4\text{H}^+ + 4\text{e}^- \\
\text{Cr}_2\text{O}_7^{2-} + 14\text{H}^+ + 6\text{e}^- & \rightarrow 2\text{Cr}^{3+} + 7\text{H}_2\text{O}
\end{align*}
\]

Step 5. Combine and simplify two half reactions.
\[
\begin{align*}
3\text{C}_2\text{H}_6\text{O} + 3\text{H}_2\text{O} + 2\text{Cr}_2\text{O}_7^{2-} + 28\text{H}^+ + 12\text{e}^- & \rightarrow 3\text{C}_2\text{H}_4\text{O}_2 + 12\text{H}^+ + 4\text{Cr}^{3+} + 14\text{H}_2\text{O} + 12\text{e}^- \\
3\text{C}_2\text{H}_6\text{O} + 2\text{Cr}_2\text{O}_7^{2-} + 16\text{H}^+ & \rightarrow 3\text{C}_2\text{H}_4\text{O}_2 + 4\text{Cr}^{3+} + 11\text{H}_2\text{O}
\end{align*}
\]

Example 6

Convert molecular chemical equation \(\text{CH}_3\text{CH}_2\text{CH}_3 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}\) to \(\text{C}_3\text{H}_8 + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}\).

Step 1. Divide into two half reactions.
\[
\begin{align*}
\text{C}_3\text{H}_8 & \rightarrow \text{CO}_2 \\
\text{O}_2 & \rightarrow \text{H}_2\text{O}
\end{align*}
\]

Step 2. Balance all atoms in two half reactions.
\[
\begin{align*}
\text{C}_3\text{H}_8 & \rightarrow 3\text{CO}_2 \\
\text{C}_3\text{H}_8 + 6\text{O} + 12\text{H}^+ & \rightarrow 3\text{CO}_2 + 8\text{H}^+ + 12\text{e}^- \\
\text{C}_3\text{H}_8 + 6\text{H}_2\text{O} & \rightarrow 3\text{CO}_2 + 20\text{H}^+
\end{align*}
\]

Step 3. Balance charges of two half reactions by adding electrons.
\[
\begin{align*}
\text{O}_2 + 4\text{H}^+ + 4\text{e}^- & \rightarrow 2\text{H}_2\text{O}
\end{align*}
\]

Step 4. Make electrons of the two half reactions equivalent (LCM = 20).
\[
\begin{align*}
(\text{C}_3\text{H}_8 + 6\text{H}_2\text{O} & \rightarrow 3\text{CO}_2 + 20\text{H}^+ + 20\text{e}^-) \times 1 \\
(\text{O}_2 + 4\text{H}^+ + 4\text{e}^- & \rightarrow 2\text{H}_2\text{O}) \times 5
\end{align*}
\]

Step 5. Combine and simplify two half reactions.
\[
\begin{align*}
\text{C}_3\text{H}_8 + 6\text{H}_2\text{O} + 5\text{O}_2 + 20\text{H}^+ + 20\text{e}^- & \rightarrow 3\text{CO}_2 + 20\text{H}^+ + 10\text{H}_2\text{O} + 20\text{e}^- \\
\text{C}_3\text{H}_8 + 5\text{O}_2 & \rightarrow 3\text{CO}_2 + 4\text{H}_2\text{O}
\end{align*}
\]

Example 7

Convert ionic chemical equation \(\text{CH}_3\text{CHO} + \text{Cu}^{2+} \rightarrow \text{CH}_3\text{COO}^- + \text{Cu}^{2+}\) (in basic medium) to \(\text{C}_2\text{H}_4\text{O} + \text{Cu}^{2+} \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{Cu}^{2+}\).

Step 1. Divide into two half reactions.
\[
\begin{align*}
\text{C}_2\text{H}_4\text{O} & \rightarrow \text{C}_2\text{H}_3\text{O}_2^- \\
\text{Cu}^{2+} & \rightarrow \text{Cu}^{2+}
\end{align*}
\]

Step 2. Balance all atoms in two half reactions.
\[
\begin{align*}
\text{C}_2\text{H}_4\text{O} + \text{O} & \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + \text{H}^+ \\
\text{Cu}^{2+} & \rightarrow \text{Cu}^{2+} + \text{H}_2\text{O}
\end{align*}
\]

Step 5. Combine and simplify two half reactions.
\[
\begin{align*}
3\text{C}_2\text{H}_4\text{O} + 3\text{H}_2\text{O} + 2\text{Cr}_2\text{O}_7^{2-} + 28\text{H}^+ + 12\text{e}^- & \rightarrow 3\text{C}_2\text{H}_3\text{O}_2^- + 12\text{H}^+ + 4\text{Cr}^{3+} + 14\text{H}_2\text{O} + 12\text{e}^- \\
3\text{C}_2\text{H}_4\text{O} + 2\text{Cr}_2\text{O}_7^{2-} + 16\text{H}^+ & \rightarrow 3\text{C}_2\text{H}_3\text{O}_2^- + 4\text{Cr}^{3+} + 11\text{H}_2\text{O}
\end{align*}
\]
Step 3. Add electrons to balance charges of half reactions.
\[ \text{C}_2\text{H}_4\text{O} + \text{H}_2\text{O} \rightarrow \text{C}_2\text{H}_3\text{O}_2^- + 3\text{H}^+ + 2\text{e}^- \\
2\text{Cu}^{2+} + \text{H}_2\text{O} + 2\text{e}^- \rightarrow \text{Cu}_2\text{O} + 2\text{H}^+ \]

Step 4. Make electrons of two half reactions equivalent (LCM = 2).
\[
\begin{align*}
(C_2H_4O + H_2O & \rightarrow C_2H_3O_2^- + 3H^+ + 2e^- ) \times 1 \\
(2Cu^{2+} + H_2O + 2e^- & \rightarrow Cu_2O + 2H^+) \times 1
\end{align*}
\]

Step 5. Combine and simplify into two half reactions.
\[
\begin{align*}
C_2H_4O + H_2O + 2Cu^{2+} + \text{H}_2O + 2e^- & \rightarrow C_2H_3O_2^- + 3H^+ + Cu_2O + 2H^+ \\
C_2H_4O + 2Cu^{2+} + 2H_2O & \rightarrow C_2H_3O_2^- + Cu_2O + 5H^-
\end{align*}
\]

Step 6. Convert H\(^+\) to OH\(^-\).
\[
\begin{align*}
C_2H_4O + 2Cu^{2+} + 2H_2O + 5OH^- & \rightarrow C_2H_3O_2^- + Cu_2O + 3H_2O \\
C_2H_4O + 2Cu^{2+} + 5OH^- & \rightarrow C_2H_3O_2^- + Cu_2O + 3H_2O \\
C_2H_4O + 2Cu^{2+} + 5OH^- & \rightarrow C_2H_3O_2^- + Cu_2O + 3H_2O
\end{align*}
\]

**Calculation of \(\Delta ON\) in balanced half redox reactions**

By using the oxidation method, the oxidation numbers of atoms are first assigned, then the change of oxidation numbers and the number of transferred electrons are counted, and finally the given equation is balanced. The proton method works in a reverse direction. By balancing an equation, the number of transferred electrons can be determined first, then the change of oxidation numbers can be calculated. Here are two half reactions from Example 7 that demonstrate calculations:

**Example 7a**

Given the balanced half redox reaction: Reduction
\[2Cu^{2+} + H_2O + 2e^- \rightarrow Cu_2O + 2H^+ \]
\[
\begin{align*}
ncu &= 2; \ Te^- = -2 \\
\Delta ON &= \frac{Te^-}{n} \\
\Delta ON Cu &= \frac{-2}{2} = -1
\end{align*}
\]

In \(2Cu^{2+} + H_2O + 2e^- \rightarrow Cu_2O + 2H^+\), there are two copper atoms gaining two electrons in the half reduction reaction. The change of mean copper oxidation numbers (\(\Delta ON Cu\)) from \(Cu^{2+}\) (ON) to \(Cu_2O\) (ON) equals -1.

**Example 7b**

Given the balanced half redox reaction: Oxidation
\[C_2H_4O + H_2O \rightarrow C_2H_3O_2^- + 3H^+ + 2e^- \]
\[
\begin{align*}
nC &= 2; \ Te^- = +2 \\
\Delta ON &= \frac{Te^-}{n} \\
\Delta ON C &= \frac{+2}{2} = +1
\end{align*}
\]

In the balanced half equation of \(C_2H_4O + H_2O \rightarrow C_2H_3O_2^- + 3H^+ + 2e^-\), two carbon atoms carry the oxidation numbers change and lose two electrons. The calculated change of mean carbon oxidation numbers (\(\Delta ON C\)) equals +1.

**Example 8**

Determine the change of mean carbon oxidation numbers (\(\Delta ON C\)) for a redox couple of \(C_6H_5CH_3\) and \(C_6H_5COOH\).

**Solution**

**Step 1.** Balance atoms.
\[
\begin{align*}
C_6H_5CH_3 & \rightarrow C_6H_5COOH \\
C_6H_5 & \rightarrow C_6H_5O_2 \\
C_6H_5 + 2O & \rightarrow C_6H_5O_2 + 2H^+ \\
C_6H_5 + 2O + 4H^- & \rightarrow C_6H_5O_2 + 2H^+ + 4H^- \\
C_6H_5 + 2H_2O & \rightarrow C_6H_5O_2 + 6H^+
\end{align*}
\]

**Step 2.** Balance charges.
\[
\begin{align*}
C_6H_5 + 2H_2O & \rightarrow C_6H_5O_2 + 6H^+
\end{align*}
\]

**Step 3.** Determine change of mean carbon oxidation numbers.
\[
\begin{align*}
C_6H_5 & \rightarrow C_6H_5O_2 + 6H^+ + 6e^- \\
nC &= 7; \ Te^- = +6 \\
\Delta ON &= \frac{Te^-}{n} \\
\Delta ON C &= \frac{+6}{7} = + \frac{6}{7}
\end{align*}
\]

**Example 9**

Given a half redox reaction \(C_6H_{12}O_6 \rightarrow CO_2\), determine the change of mean carbon oxidation numbers (\(\Delta ON C\)).

**Solution**

**Step 1.** Balance atoms.
\[
\begin{align*}
C_6H_{12}O_6 & \rightarrow CO_2 \\
C_6H_{12}O_6 & \rightarrow 6CO_2 \\
C_6H_{12}O_6 + 6O & \rightarrow 6CO_2 + 12H^- \\
C_6H_{12}O_6 + 6O + 12H^- & \rightarrow 6CO_2 + 12H^- + 12H^- \\
C_6H_{12}O_6 + 6H_2O & \rightarrow 6CO_2 + 24H^- \\
\end{align*}
\]

**Step 2.** Balance charges.
\[
\begin{align*}
C_6H_{12}O_6 + 6H_2O & \rightarrow 6CO_2 + 24H^- \\
C_6H_{12}O_6 + 6H_2O & \rightarrow 6CO_2 + 24H^- + 24e-
\end{align*}
\]
Step 3. Determine change of mean carbon oxidation numbers.

\[
\text{C}_6\text{H}_{12}\text{O}_6 + 6\text{H}_2\text{O} \rightarrow 6\text{CO}_2 + 24\text{H}^+ + 24\text{e}^-
\]

\[
\Delta \text{ON} = \frac{\text{Te}^- - n}{n} \Delta \text{ONc} = \frac{+24}{6} = +4
\]

Conclusion

By using a non-oxidation number approach, the proton method can be used to balance atoms and charges in redox reactions for both molecular and ionic chemical equations. Regarding a balanced half redox equation, the mathematic equation of \(\text{Te}^- = n \Delta \text{ON}\) is established. This offers a new approach to count the change of mean oxidation numbers for an assigned atom in a balanced half reaction.

References


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