Oxidation-Reduction
An Introduction

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Oxidation-Reduction Reactions

In an oxidation-reduction (Redox) reaction, electrons are transferred from one species to another.

For example, in a single replacement reaction

$$Cu_{(s)} + 2\ AgNO_{3\,(aq)} \rightarrow 2\ Ag_{(s)} + Cu(NO_3)_{2\,(aq)}$$

The Cu atoms lose electrons to form Cu^{2+} in the $Cu(NO_3)_2$ and the Ag^{+} gains electrons to form metallic Ag

Oxidation-Reduction Reactions

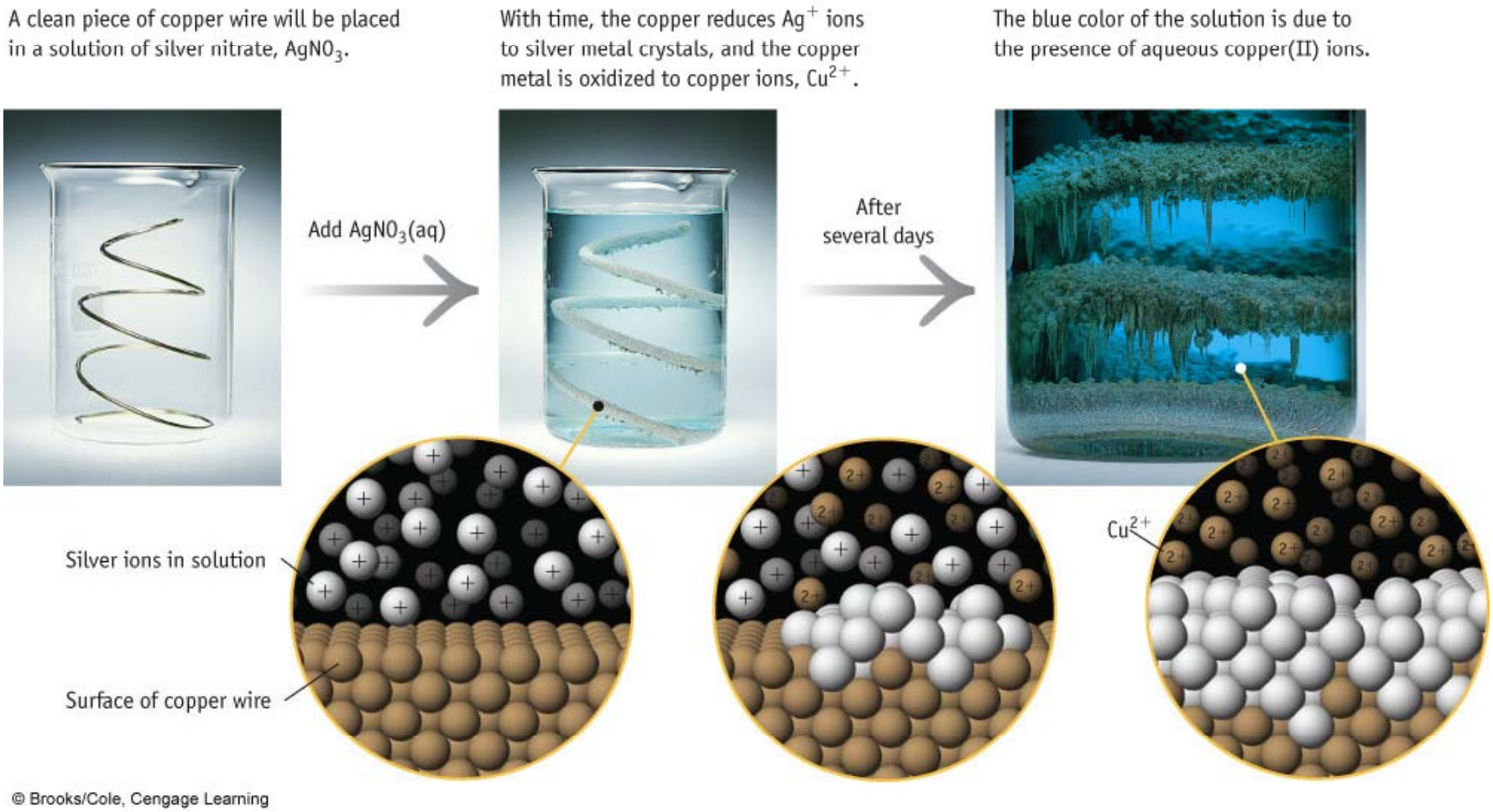
- This can be more easily observed by writing the net ionic equation for the reaction:

$$Cu_{(s)} + 2\,Ag^{+}_{(aq)} \rightarrow 2\,Ag_{(s)} + Cu^{2+}_{(aq)}$$

- The metallic Cu atoms are uncombined, so they are considered to have an oxidation number of zero.
- The combined Ag atoms are in a +1 oxidation state.
- Each Cu atom will lose 2 electrons to 2 Ag^{+} ions
- The resulting Ag atoms are considered to have an oxidation number of zero

Oxidation-Reduction Reactions

$$Cu_{(s)} + 2\,Ag^{+}_{(aq)} \rightarrow 2\,Ag_{(s)} + Cu^{2+}_{(aq)}$$



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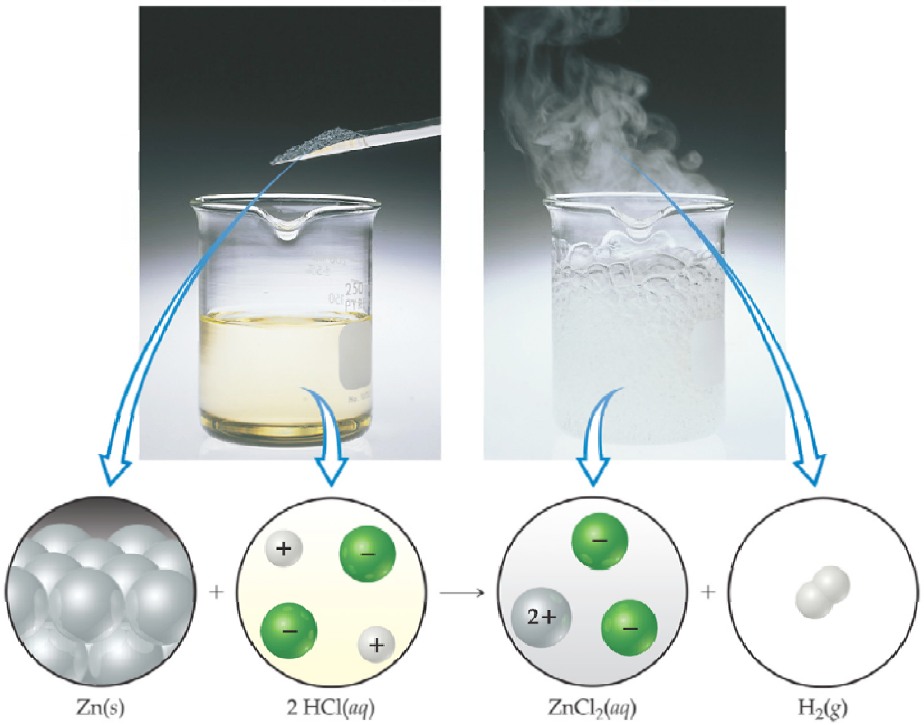
Oxidation Numbers

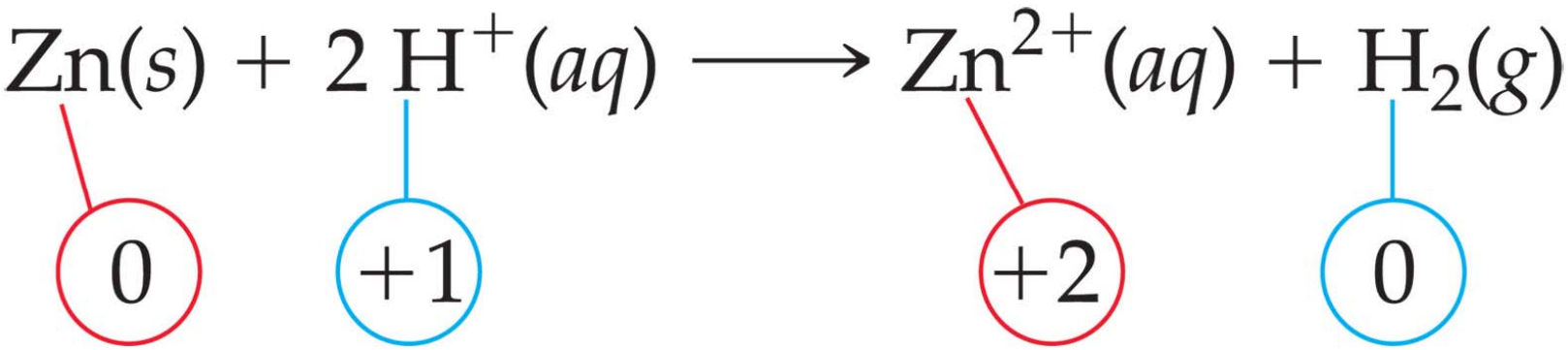
$$Zn(s) + 2\,H^{+}(aq) \longrightarrow Zn^{2+}(aq) + H_2(g)$$

0 +1 +2 0



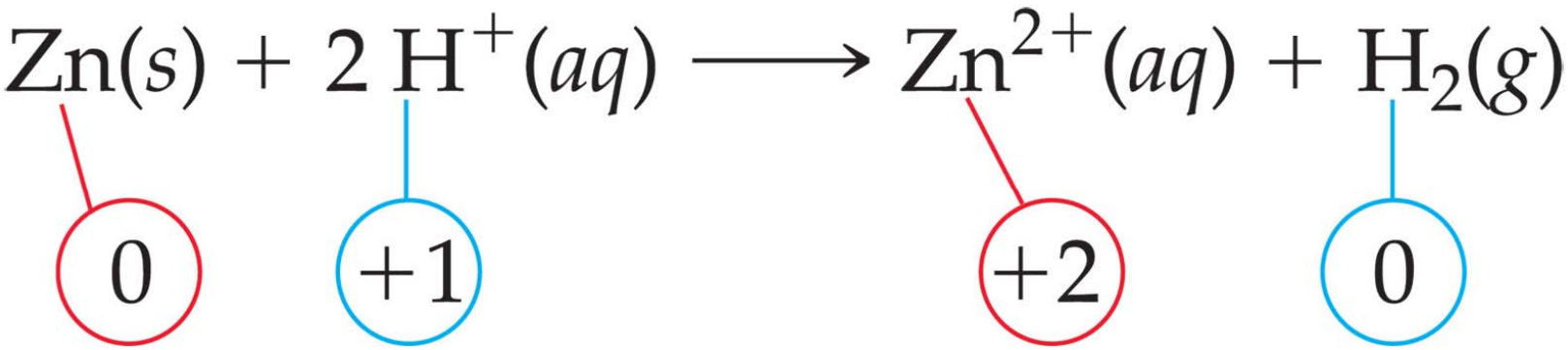
In order to keep track of what loses electrons and what gains them, we list the oxidation numbers of each element.

Oxidation and Reduction



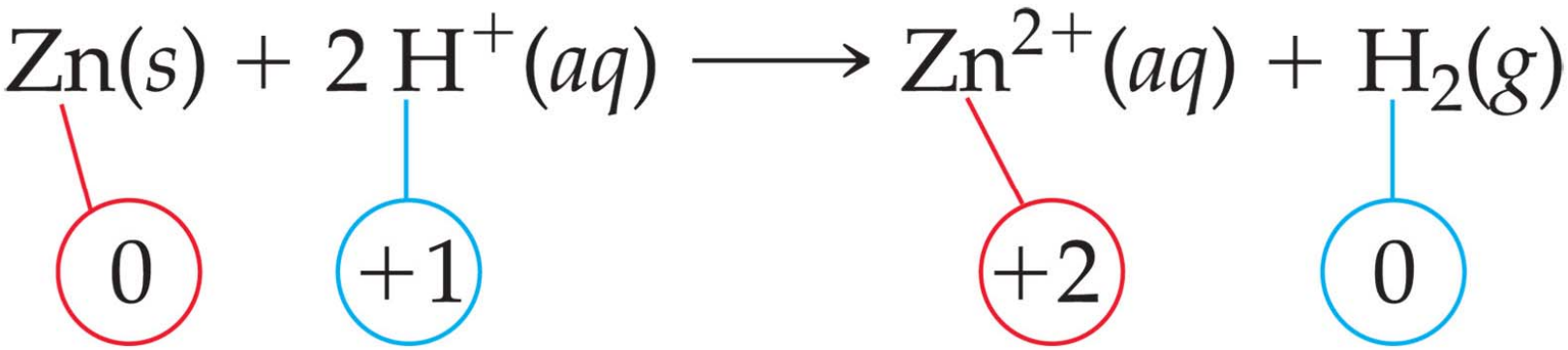
- **A species is oxidized when it loses electrons.**
  - **Here, zinc loses two electrons to go from neutral zinc metal to the Zn^{2+} ion.**
  - **Oxidation is observed as an increase in oxidation number**

Oxidation and Reduction



- **A species is reduced when it gains electrons.**
  - **Here, each of the H^+ gains an electron and they combine to form H_2.**
  - **Reduction is observed as an decrease in oxidation number**

Oxidation and Reduction



- **The species that contains the element that is reduced is the oxidizing agent.**
  - **H^+ oxidizes Zn by taking electrons from it.**
- **The species that contains the element that is oxidized is the reducing agent.**
  - **Zn reduces H^+ by giving it electrons.**

Assigning Oxidation Numbers

1. **Elements in their elemental form have an oxidation number of 0.**
2. **The oxidation number of a monatomic ion is the same as its charge.**

Assigning Oxidation Numbers

3. **The oxidation number of metals depends on their position in the periodic table**
   - **Group IA elements are +1**
   - **Group IIA elements are +2**
   - **Group IIIA elements are +3**
   - **Group IVA metals are usually +2 or +4**
   - **Group VA metals are usually +3 or +5**

Assigning Oxidation Numbers

4. **Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.**
   - **Oxygen always has an oxidation number of −2, except in the peroxide ion in which it has an oxidation number of −1.**
   - **Hydrogen is always −1 when bonded to a metal**
   - **Hydrogen is +1 when bonded to a nonmetal.**

Assigning Oxidation Numbers

4. **Nonmetals (continued).**
   - **Fluorine always has an oxidation number of –1.**
   - **The halogens (Cl, Br, and I)have an oxidation number of –1 when they are negative**
   - **The halogens (Cl, Br, and I) will have positive oxidation numbers in oxyanions (ClO^-, ClO_2^-, ClO_3^-, etc.)**

Assigning Oxidation Numbers

5. The sum of the oxidation numbers in a neutral compound is 0.
6. The sum of the oxidation numbers in a polyatomic ion is the charge on the ion.