Oxidation-Reduction (REDOX) Reactions

Unit: Oxidation & Reduction

MA Curriculum Frameworks (2016): HS-PS1-10(MA)

Mastery Objective(s): (Students will be able to…)

- Assign oxidation numbers.
- Write and balance equations for simple REDOX reactions.

Success Criteria:

- Oxidation numbers agree with published/accepted values.
- Oxidation numbers add up to zero for compounds and to the charge for ions.
- Balanced REDOX reactions have the same number of each type of atom and the same number of electrons transferred on each side.

Tier 2 Vocabulary: reduce

Language Objectives:

- Explain electron transfer. Explain the charge that an ion gets when electrons are added or removed.

Notes:

oxidation-reduction reaction (REDOX reaction): a reaction in which one or more electrons are transferred from one atom to another.

In the 1700s, oxidation meant that an atom was combined with oxygen, and was therefore "oxidized". For example:

$$2\,Cu + O_2 \rightarrow 2\,CuO$$

In this reaction, oxygen is more electronegative than copper, so oxygen took electrons away from copper. This means that the copper (which was "oxidized" by oxygen) lost two electrons and ended up with a +2 charge. Oxygen gained two electrons and ended up with a −2 charge. As a result, the term "oxidation" has come to mean "losing electrons".

Also, in the 1700s, scientists found that if they heated the CuO (in which copper has a +2 charge), they ended up with copper metal (with a charge of zero), and the weight was reduced. Thus copper was said to be "reduced". As a result, the term "reduction" has come to mean "gaining electrons".

Use this space for summary and/or additional notes:

Big Ideas | Details | Unit: Oxidation & Reduction

oxidation: the loss of one or more electrons by an atom in a chemical reaction

reduction: the gain of one or more electrons by an atom in a chemical reaction.

Stupid Mnemonics: There are two popular mnemonics for remembering oxidation and reduction, one "Democratic" and one "Republican".

LEO the lion says 'GER' ("Democratic" mnemonic involving endangered species): LEO stands for "Loss of Electrons is Oxidation" and GER stands for "Gain of Electrons is Reduction"

OIL RIG ("Republican" mnemonic involving oil companies): OIL stands for "Oxidation Involves Loss (of electrons)", and RIG stands for "Reduction Involves Gain (of electrons)."

In a redox reaction, at least one element is oxidized, and at least one other element is reduced. *An element cannot be oxidized in a chemical reaction unless some other element is reduced, and vice-versa.* (After all, the electrons have to come from somewhere, and they have to go somewhere.)

All chemical reactions in which an element becomes part of a compound, or vice-versa, are redox reactions. This includes all single replacement reactions, combustion reactions, and many synthesis and decomposition reactions. However, chemists generally classify a reaction as a redox reaction only when most or all of the energy of the reaction comes from electron transfer.

Use this space for summary and/or additional notes:

Oxidation Numbers

An oxidation number is a measure of how "oxidized" an atom is. An element is neither oxidized nor reduced, so it has an oxidation number of zero.

- An element that has lost electrons (oxidized) gets a positive oxidation number, equal to the number of electrons it has lost.
- An element that has gained electrons (reduced) gets a negative oxidation number, equal to the number of electrons it has gained.

Therefore:

- when an element is oxidized, the oxidation number increases.
- When an element is reduced, the oxidation number is also reduced (decreases).

oxidation number (or "oxidation state"): the charge that an atom would have in a compound if all bonds were completely ionic and every atom in the compound or ion had a charge.

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Assigning Oxidation Numbers

- The oxidation number of a pure element is 0. (Even if it's diatomic.)
- The oxidation numbers in a compound add up to 0
- The oxidation number of an ion is its charge. (Oxidation numbers in a polyatomic ion add up to the charge of the polyatomic ion.)
- In a compound or polyatomic ion:
  - The most electronegative element (the last one in the formula) has a negative oxidation number equal to the number of electrons it would need to fill its valent shell.
  - All other atoms have positive oxidation numbers.
  - Fluorine is always −1.
  - Oxygen is always −2 except in the peroxide ion (O_2^{2-}) and in OF_2.
  - Hydrogen is always +1 except in metal hydrides (such as NaH).
  - Alkali (group 1) metals are always +1.
  - Alkaline Earth (group 2) metals are always +2.
  - Al is always +3, Zn is always +2, and Ag is always +1.
  - Calculate other elements from the above.

Sample Problem:

What are the oxidation numbers of each element in the compound Na_2HPO_4:

- Na_2HPO_4 is an ionic compound made of the ions Na^+ and HPO_4^{2-}.
- The Na^+ ion has a charge of +1, so the oxidation number of Na is +1.
- The HPO_4^{2-} ion has a charge of −2. This means the oxidation numbers of H, P, and O must add up to −2.
  - O = −2. There are 4 O atoms. (4)(−2) = −8
  - H is +1.
  - If the oxidation numbers for the O atoms add up to −8 and H is +1, then the oxidation number for P must be +5 so the total can add up to −2.

Use this space for summary and/or additional notes:

Balancing REDOX Reactions

To fully balance a redox reaction, you must balance:

- Atoms (as you would in a regular equation)
- Electrons lost/gained
- Total charge

Often, redox reactions are shown and balanced as net ionic equations. In this case, balancing them is often a simple matter of making sure that the same number of electrons are produced by the oxidation half-reaction and consumed by the reduction half-reaction.

For example, consider the unbalanced net ionic equation:

$$Al^0(s) + Zn^{2+}(aq) \rightarrow Al^{3+}(aq) + Zn^0(s)$$

In this reaction, Al is oxidized from Al^0 to Al^{+3}, and Zn is reduced from Zn^{+2} to Zn^0. The ***atoms*** appear balanced, but Zn^{2+} needs only 2 electrons to form Zn^0, but Al^0 produces 3 electrons when oxidized to Al^{3+}.

The two half-reactions are:

$$\text{Oxidation: } Al^0 \rightarrow Al^{3+} + 3\,e^-$$
$$\text{Reduction: } Zn^{2+} + 2\,e^- \rightarrow Zn^0$$

To balance the electrons, we need to multiply the first half-reaction by 2, and the second one by 3, giving:

$$2\,Al^0 \rightarrow 2\,Al^{3+} + 6\,e^-$$
$$3\,Zn^{2+} + 6\,e^- \rightarrow 3\,Zn^0$$

If we combine these and cancel the electrons (because we have the same number on both sides), we get the balanced net ionic equation:

$$2\,Al^0(s) + 3\,Zn^{2+}(aq) \rightarrow 2\,Al^{3+}(aq) + 3\,Zn^0(s)$$

REDOX reactions can get a lot more complicated, especially when acid-base reactions are also taking place and the water that the ions are dissolved in participates in the reaction. Balancing complex REDOX reactions is beyond the scope of this course, but is covered in AP® Chemistry.

Use this space for summary and/or additional notes:

Homework Problems

For each of the following compounds or ions, write the oxidation number of each element.

1. FeO

2. N_2

3. $KMnO_4$

4. NH_4OH

5. $(NH_4)_3PO_4$

6. $Cr_2O_7^{2-}$

Use this space for summary and/or additional notes:

Balance the following (redox) equations. A superscript "0" indicates a pure element.

7. $Zn^0 + Tl^+ \rightarrow Tl^0 + Zn^{2+}$

8. $Li^0 + Cr^{3+} \rightarrow Cr^0 + Li^+$

9. $K^0 + Mg^{2+} \rightarrow K^+ + Mg^0$

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