

## ALE 23. Balancing Redox Reactions

(Reference: Section 4.5 (pp. 158 – 166) and 21.1 Silberberg 5<sup>th</sup> edition)

How does one balance a reaction for both matter and charge?

**The Model**

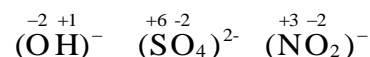
**Oxidation-reduction or Redox reactions** involve the transfer of one or more electrons from one chemical species to another. Redox reactions are involved in the corrosion of metals, the combustion of fuels, the generation of electricity from batteries and many biological processes including cellular respiration and photosynthesis. An understanding of redox chemistry is essential in the design of new kinds of batteries, increasing efficiency in fuel combustion, the prevention of corrosion, etc.

Recall from Chem 161 (Sec. 4.5 in Silberberg), the **oxidation number** of an atom is the “apparent” charge the atom would have if each of its bonding electrons were assigned to the more electronegative atom in each bond. Oxidation numbers are useful in determining the substance **oxidized** (**L.E.O.** = Loss of Electron(s) is **O**xidation) and the substance **reduced** (**G.E.R.** = Gain of Electron(s) is **R**eduction). Hence, the substance that is oxidized loses electrons and therefore serves as a **reducing agent** since it provides electrons to another atom thereby causing that atom to be reduced. The species being reduced serves as the **oxidizing agent** because it removes electrons from another substance, thereby causing that substance to be oxidized.

Rules to Assign Oxidation Numbers

Oxidation numbers are often written above the atomic symbol:  $\overset{+1}{\text{H}}\overset{+5}{\text{N}}\overset{-2}{\text{O}}_3$   $\overset{0}{\text{F}}_2$   $\overset{-4}{\text{C}}\overset{+1}{\text{H}}_4$   $\overset{+1}{\text{H}}_2\overset{-1}{\text{O}}_2$

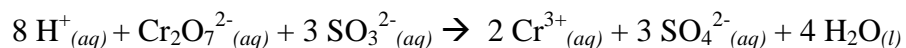
- 1) In **free elements** (*i.e.*, in the uncombined state), each atom has an oxidation number of zero.  
e.g. Al, S<sub>8</sub>, Na, Fe, H<sub>2</sub>, O<sub>2</sub>, N<sub>2</sub>, Cl<sub>2</sub> each have an O.N. = 0
- 2) Since **fluorine** is the most electronegative element, the oxidation number of fluorine in all of its compounds is -1.
- 3) The oxidation number of **oxygen** in most compounds (called “oxides”) is -2. In **peroxides** (e.g. H<sub>2</sub>O<sub>2</sub>), the oxidation number of oxygen is -1. In **superoxides** (e.g. O<sub>2</sub><sup>1-</sup>), the oxidation number of oxygen is -1/2.
- 4) The oxidation number of **hydrogen** is +1, except when it is bonded to metals in binary compounds (e.g. metallic hydrides: LiH, MgH<sub>2</sub>), in which its oxidation number is -1.
- 5) Group **IA and IIA metals** form compounds in which the metal atoms have oxidation numbers of +1 and +2, respectively.
- 6) The sum of the oxidation numbers of all the atoms in a species must be equal to the net charge of the species. e.g. The sum of the O.N.’s of a molecular substance is zero:  $\overset{+1}{\text{H}}_2\overset{-2}{\text{O}}$   $\overset{-4}{\text{C}}\overset{+1}{\text{H}}_4$   $\overset{+5}{\text{C}}\overset{-1}{\text{F}}_5$   $\overset{+4}{\text{N}}\overset{-2}{\text{O}}_2$   
The sum of the O.N.’s of the atoms in a polyatomic equals the charge of the ion:



## Key Questions

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1. The substances oxidized and reduced may not be obvious in the following redox reaction:



- Write the oxidation numbers above each atomic symbol on the left and right sides of the reaction above.
- Does the oxidation number of any atom increase? Is that species the oxidizing agent or the reducing agent?
- Does the oxidation number of any atom decrease? Is that species the oxidizing agent or the reducing agent?
- Which species is being oxidized? \_\_\_\_\_
- Which species is being reduced? \_\_\_\_\_

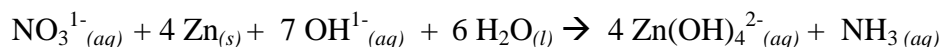
## Exercises

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2. Write the oxidation number above each symbol in the following compounds or ions.

- |                                   |                                   |                      |                     |
|-----------------------------------|-----------------------------------|----------------------|---------------------|
| a. KBr                            | b. BrF <sub>3</sub>               | c. HBrO <sub>3</sub> | d. CBr <sub>4</sub> |
| e. MnO <sub>4</sub> <sup>1-</sup> | f. Mn <sub>2</sub> O <sub>3</sub> | g. KMnO <sub>4</sub> |                     |

3. Assign oxidation numbers to each atom in the following reaction and then identify the oxidizing agent, reducing agent, the substance oxidized and the substance reduced.



Oxidizing agent: \_\_\_\_\_

Reducing agent: \_\_\_\_\_

Substance oxidized: \_\_\_\_\_

Substance reduced: \_\_\_\_\_



- 5 a. Why are the stoichiometric coefficients of the oxidation half reaction multiplied by 12 in [Step H of example 1](#)?
- b. True or False: An oxidation reaction *must* occur at the same time that a reduction reaction takes place. Justify your response.

### Exercise

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6. The permanganate ion,  $\text{MnO}_4^-$ , reacts with oxalate,  $\text{C}_2\text{O}_4^{2-}$ , to yield manganese (IV) oxide and carbonate,  $\text{CO}_3^{2-}$ . Use the half-reaction method to balance this reaction in acidic conditions. Show your work.

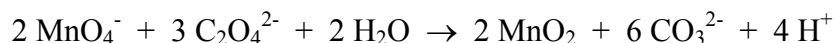
### The Model: Balancing a Redox Reaction in an Basic Solution

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The reaction you balanced in [exercise #6](#), above, between the permanganate ion and the oxalate ion occurs in basic solution—not in acidic solution! To balance a redox reaction in basic solution follow the steps used to balance an equation in acidic conditions (i.e. [steps A – J on page 3](#)) and then follow [steps K – L](#), below.

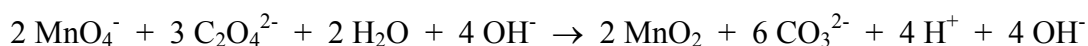
**Example 2.** In a basic solution, permanganate reacts with oxalate to yield manganese(IV) oxide and carbonate.

This is what you should have ended up with in [exercise #6](#), above, after completing [steps A – J on page 3](#):



To balance this equation in basic solution...

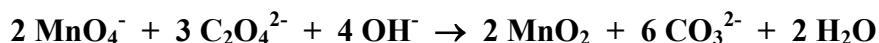
K. Make the solution basic by adding  $\text{OH}^-$  to each side:



L. Neutralize the  $\text{H}^+$  and  $\text{OH}^-$  to create water:



M. simplify algebraically, canceling any terms (e.g. water) that appear on both sides:



## Key Question

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7. Why in [Step K of Example 2](#) is “4 OH<sup>-</sup>” added to each side of the reaction? (Such a step did not take place in Example 1. What’s different about Example 2?)

## Exercises

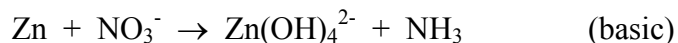
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Balance the following redox reactions by using the half-reaction method. *Show your work.*

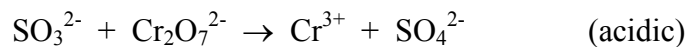
8. Many metals dissolve in hydrochloric acids. Copper does not. To dissolve copper, one must use an oxidizing acid, such as nitric acid. Gold, however, does not dissolve even in nitric acid. To dissolve the unreactive metal, one must place gold in *aqua regia*, which is a mixture of concentrated nitric acid and concentrated hydrochloric acid. The unbalanced chemical equation for that process is:



9. Zinc is a reactive metal that dissolves readily in hydrochloric or in nitric acid. But zinc also dissolves in basic solutions containing the nitrate ion. The unbalanced chemical equation for that process is:



10. A convenient way to volumetrically analyze solutions of sulfite (colorless) is to titrate them with an acidified standardized dichromate solution (orange).



The resulting chromium (III) sulfate solution is green. Suppose 25.00 mL of a sodium sulfite solution is transferred to an Erlenmeyer flask and acidified. It is then titrated with 0.0628 M  $\text{K}_2\text{Cr}_2\text{O}_7(aq)$ . When 22.38 mL of the dichromate solution has been added, the first presence of orange indicates “left over” dichromate and the endpoint has been reached. What is the molarity of the original  $\text{Na}_2\text{SO}_3(aq)$  solution? *Show your work.*